**FORM ONE**

**Introduction to chemistry**

Chemistry is a branch of Science. Science is basically the study of living and nonliving things. The branch of science that study living things is called Biology. The branch of science that study non-living things is called Physical Science. Physical Science is made up of:

(i)Physics- the study of matter in relation to energy (ii)Chemistry- the study of composition of matter.

Chemistry is thus defined as the branch of science that deals with the structure composition, properties and behavior of matter.

Basic Chemistry involves studying:

**(a)States/phases of matter**

Matter is anything that has weight/**mass** and occupies space/**volume.** Naturally, there are basically **three** states of matter.

(i) **Solid**-e.g. soil, sand, copper metal ,bucket, ice.

(ii)**Liquid-**e.g water, Petrol, ethanol/alcohol, Mercury(liquid metal). (iii)**gas-** e.g. Oxygen, Nitrogen ,Water vapour.

A solid is made up of particles which are very closely packed. It thus has a definite/fixed shape and fixed/definite volume /occupies definite space. It has a very high density.

A liquid is made up of particles which have some degree of freedom. It thus has no definite/fixed shape. It takes the shape of the container it is put. A liquid has fixed/definite volume/occupies definite space.

A gas is made up of particles free from each other. It thus has no definite/fixed shape. It takes the shape of the container it is put. It has no fixed/definite volume/occupies every space in a container.

**(b) Separation of mixture**

A mixture is a combination of two or more substances that can be separated by physical means. Simple methods of separating mixtures at basic chemistry level include

(i)**Sorting/picking-**this involve physically picking one pure substance from a mixture with another/other. e. g. sorting maize from maize beans mixture. (ii)**Decantation-**this involve pouring out a liquid from a solid that has settled /sinking solid in it. e. g. Decanting water form sand **.**

(iii)**Filtration**-this involves sieving /passing particles of a mixture through a filter containing small holes that allow smaller particle to pass through but do not allow bigger particle to pass through.

**(iv)Skimming-**this involve scooping floating particles. e.g. cream from milk

1. **Metals and non-metals**

Metals are shiny, ductile(able to form wires),malleable(able to form sheet) and coil without breaking. e.g. Iron, gold, silver, copper. Mercury is the only **liquid metal** known.

Non-metals are dull, not ductile(do not form wires), not malleable(do not form sheet) and break on coiling/brittle. e.g. Charcoal, Sulphur , plastics.

1. **Conductors and non-conductors**

A conductor is a solid that allow electric current to pass through. A non-conductor is a solid that do not allow electric current to pass through.

All metals conduct electricity. All non-metals do not conduct electricity except carbon **graphite**.

**(e)Drugs**

A drug is a natural or synthetic/man-made substance that when taken changes/alter the body functioning. A natural or synthetic/man-made substance that when taken changes/alter the abnormal body functioning to normal is called **medicine**. Medicines are thus drugs intended to correct abnormal body functions. . Medicines should therefore be taken on **prescription** and **dosage**.

A prescription is a medical instruction to a patient/sick on the correct type of medicine to take and period/time between one intake to the other . A dosage is the correct quantity of drug required to alter the abnormal body function back to normal. This is called **treatment**.

It is the professional work of qualified doctors/pharmacists to administer correct prescription and dosage of drugs/medicine to the sick.

Prescription and dosage of drugs/medicine to the sick use medical language.

**Example**

1. **2 x 4** ; means “**2**” tablets for **solid** drugs/spoon fulls for **liquid** drugs taken “**4**” times for a duration of one day/24 hours and then repeated and continued until all the drug given is finished.
2. **1 x 2** ; means “**1**” tablets for **solid** drugs/spoon fulls for **liquid** drugs taken “**2**” times for a duration of one day/24 hours and then repeated and continued until all the drug given is finished.

Some drugs need minimal prescription and thus are available without pharmacist/ doctor’s prescription. They are called **O**ver **T**he **C**ounter(OTC) drugs. OTC drugs used to treat mild headaches, stomach upsets, common cold include:

(i) painkillers (ii) anti acids (iii) cold/flu drugs.

All medicines require correct intake dosage. When a prescription dosage is not followed, this is called drug **misuse/abuse.**

Some drugs are used for other purposes other than that intended. This is called **drug abuse**.

Drug abuse is when a drug is intentionally used to alter the normal functioning of the body. The intentional abnormal function of the drug is to make the victim have false feeling of well being.

The victim lack both mental and physical coordination.

Some drugs that induce a false feeling of well being are illegal. They include heroin, cocaine, bhang, mandrax and morphine.

Some abused drugs which are not illegal include: miraa, alcohol, tobacco, sleeping pills.

**The role of chemistry in society**

**(a)**Chemistry is used in the following:

(i)Washing/cleaning with soap:

Washing/cleaning is a chemical process that involve jnteraction of water,soap and dirt so as to remove the dirt from a garment.

(ii)Understanding chemicals of life

Living thing grow, respire and feed. The formation and growth of cells involve chemical processes in living things using carbohydrates, proteins and vitamins.

(iii)Baking:

Adding baking powder to dough and then heating in an oven involves interactions that require understanding of chemistry.

(iv)Medicine:

Discovery, test ,prescription and dosage of drugs to be used for medicinal purposes require advanced understanding of chemistry

(v)Fractional distillation of crude oil:

Crude oil is fractional distilled to useful portions like petrol,diesel,kerosene by applying chemistry.

(vi)Manufacture of synthetic compounds/substances

Large amounts of plastics, glass, fertilizers, insecticides, soaps, cements, are manufactured worldwide. Advanced understanding of the chemical processes involved is a requirement.

(vii)Diagnosis/test for abnormal body functions.

If the body is not functioning normally,it is said to be sick/ill.Laboaratory test are done to diagnose the illness/sickness.

**(b)**The following career fields require Chemistry as one of subject areas of advanced/specialized study:

(i)Chemical engineering/chemical engineer

(ii)Veterinary medicine/Veterinary doctor

(iii)Medicine/Medical doctor/pharmacist/nurse

(iv)Beauty/Beautician

(v)Teaching/Chemistry teacher.

**The School Chemistry Laboratory**

Chemistry is studied mainly in a science room called a school chemistry **laboratory**.

The room is better ventilated than normal classroom. It has electricity, gas and water **taps**.

A school chemistry laboratory has a qualified professional whose called Laboratory technician/assistant.

All students user in a school chemistry laboratory must consult the Laboratory technician/assistant for all their laboratory work.

A school chemistry laboratory has chemicals and apparatus.

A chemical is a substance whose composition is known. All chemical are thus labeled as they are.

This is because whereas physically a substance may appear similar, chemically they may be different.

All Chemicals which are not labeled should never be use.

Some chemicals are toxic/poisonous, explosive, corrosive, caustic, irritants, flammable, oxidizing, carcinogenic, or radioactive.

Care should always be taken when handling any chemical which have any of the above characteristic properties.

Common school chemistry laboratory chemicals include:

(i)distilled water

(ii)Concentrated mineral acid which are very corrosive(on contact with skin they cause painful open wounds)

(iii)Concentrated alkali/bases which are caustic(on contact with skin they cause painful blisters) (iv)Very many types of salts

The following safety guideline rules should be followed by chemistry laboratory users:

(i)Enter the laboratory with permission in an orderly manner without rushing/pushing/scrabbling.

(ii)Do not try unauthorized experiments. They may produce flammable, explosive or toxic substances that affect your health.

(iii)Do not taste any chemical in the laboratory. They may be poisonous.

(iv)Waft gas fumes to your nose with your palm.Do not inhale/smell gases directly. They may be highly poisonous/toxic.

(v)Boil substances with mouth of the test tube facing away from others and yourself. Boiling liquids spurt out portions of the hot liquid. Products of heating solids may be a highly poisonous/toxic gas.

(vi)Wash with lots of water any skin contact with chemicals immediately.Report immediately to teacher/laboratory technician any irritation, cut, burn, bruise or feelings arising from laboratory work.

(vii)Read and follow safety instruction.All experiments that evolve/produce poisonous gases should be done in the open or in a fume chamber.

(viii)Clean your laboratory work station after use.Wash your hand before leaving the chemistry laboratory.

(ix)In case of fire, remain calm, switch of the source of fuel-gas tap. Leave the laboratory through the emergency door. Use fire extinguishers near the chemistry laboratory to put of medium fires. Leave strong fires wholly to professional fire fighters.

(x)Do not carry unauthorized item from a chemistry laboratory.

An apparator /apparatus are scientific tools/equipment used in performing scientific experiments. The conventional apparator used in performing a scientific experiments is called **standard** apparator/apparatus. If the conventional standard apparator/apparatus is not available, an **improvised** apparator/apparatus may be used in performing a scientific experiments. An improvised apparator/apparatus is one used in performing a scientific experiment **for** a standard apparator/apparatus. Most standard apparatus in a school chemistry laboratory are made of **glass** because:

(i)Glass is transparent and thus reactions /interactions inside are clearly visible from outside

(ii)Glass is comparatively cheaper which reduces cost of equipping the school chemistry laboratory

(iii)glass is comparatively easy to clean/wash after use.

(iv)glass is comparatively unreactive to many chemicals.

Apparatus are designed for the purpose they are intended in a school chemistry laboratory:

**(a)Apparatus for measuring volume**

1. Measuring cylinder

Measuring cylinders are apparatus used to measure volume of liquid/ solutions.

They are calibrated/ graduated to measure any volume required to the maximum. Measuring cylinders are named according to the maximum calibrated/graduated volume e.g.

“10ml” measuring cylinder is can hold maximum calibrated/graduated volume of “10mililitres” /“10 cubic centimetres” “50ml” measuring cylinder is can hold maximum calibrated/graduated volume of “50mililitres” /“50 cubic centimetres”

“250ml” measuring cylinder is can hold maximum calibrated/graduated volume of “250mililitres” /“250 cubic centimetres” “1000ml” measuring cylinder is can hold maximum calibrated/graduated volume of “1000mililitres” /“1000 cubic centimetres”

2.Burette

Burette is a long and narrow/thin apparatus used to measure small accurate and exact volumes of a liquid solution. It must be clamped first on a stand before being used. It has a tap to run out the required amount out. They are calibrated/ graduated to run out small volume required to the maximum 50ml/50cm3.

The maximum 50ml/50cm3 calibration/ graduation reading is at the **bottom** .This ensure the amount run **out** from a tap **below** can be determined directly from **burette reading** before and after during volumetric analysis. **Burettes are expensive and care should be taken when using them.**

1. (i) Pipette

Pipette is a long and narrow/thin apparatus that widens at the middle used to measure and transfer small very accurate/exact volumes of a liquid solution.

It is open on either ends.

The maximum 25ml/25cm3 calibration/ graduation mark is a visible **ring** on one thin end.

To fill a pipette to this mark, the user must suck up a liquid solution upto a level above the mark then adjust to the mark using a finger.

This require practice.

(ii) Pipette filler

Pipette filler is used to suck in a liquid solution into a pipette instead of using the mouth. It has a suck, adjust and eject button for ensuring the exact volume is attained. This requires practice.

1. Volumetric flask.

A volumetric flask is thin /narrow but widens at the base/bottom. It is used to measure very accurate/exact volumes of a liquid solution. The maximum calibration / graduation mark is a visible **ring.**

Volumetric flasks are named according to the maximum calibrated/graduated volume e.g.

“250ml” volumetric flask has a calibrated/graduated mark at exact volume of “250mililitres” /“250centimetres”

“1l” volumetric flask has a calibrated/graduated mark at exact volume of “one litre” /“1000 cubic centimetres”

“2l” volumetric flask has a calibrated/graduated mark at exact volume of “two litres” /“2000 cubic centimetres”

5. Dropper/teat pipette

A dropper/teat pipette is a long thin/narrow glass/rubber apparatus that has a flexible rubber head.

A dropper/teat pipette is used to measure very small amount/ drops of liquid solution by pressing the flexible rubber head. The number of drops needed are counted by pressing the rubber gently at a time

**(b)Apparatus for measuring mass**

1. Beam balance

A beam balance has a pan where a substance of unknown mass is placed. The scales on the opposite end are adjusted to “balance” with the mass of the unknown substance. The mass from a beam balance is in **grams**.

2. Electronic/electric balance.

An electronic/electric balance has a pan where a substance of unknown mass is placed. The mass of the unknown substance in **grams** is available immediately on the screen.

**(c)Apparatus for measuring temperature**

A thermometer has alcohol or mercury trapped in a bulb with a thin enclosed outlet for the alcohol/mercury in the bulb.

If temperature rises in the bulb, the alchohol /mercury expand along the thin narrow enclosed outlet.

The higher the temperature, the more the expansion.

Outside, a calibration /graduation correspond to this expansion and thus changes in temperature.

A thermometer therefore determines the temperature when the bulb is fully dipped in to the substance being tested. To determine the temperature of solid is thus very difficult.

**(d)Apparatus for measuring time**

The stop watch/clock is the standard apparatus for measuring time. Time is measured using hours, minutes and second.

Common school stop watch/clock has start, stop and reset button for determining time for a chemical reaction.This require practice.

**(e) Apparatus for scooping**

1. Spatula

A spatula is used to **scoop** solids which do not require accurate measurement. Both ends of the spatula can be used at a time.

A solid scooped to the **brim** is “one spatula end full” A solid scooped to **half** **brim** is “half spatula end full”.

2. Deflagrating spoon

A deflagrating spoon is used to **scoop** solids which do not require accurate measurement mainly for heating. Unlike a spatula, a deflagrating spoon is longer.

**(f) Apparatus for putting liquids/solid for heating.**

1. Test tube.

A test tube is a narrow/thin glass apparatus open on one side. The end of the opening is commonly called the “the mouth of the test tube”.

2. Boiling/ignition tube.

A boiling/ignition tube is a wide glass apparatus than a test tube open on one side. The end of the opening is commonly called the “the mouth of the boiling/ignition tube”.

3. Beaker.

Beaker is a wide calibrated/graduated lipped glass/plastic apparatus used for transferring liquid solution which do not normally require very accurate measurements

Beakers are named according to the maximum calibrated/graduated volume they can hold e.g.

“250ml” beaker has a maximum calibrated/graduated volume of “250mililitres” /“250 cubic centimetres”

“1l” beaker has a maximum calibrated/graduated volume of “one litre” /“1000 cubic centimetres”

“5 l” beaker has a maximum calibrated/graduated volume of “two litres” /“2000 cubic centimetres”

4. Conical flask.

A conical flask is a moderately narrow glass apparatus with a wide base and no calibration/graduation. Conical flasks thus carry/hold exact volumes of liquids that have been measured using other apparatus. It can also be put some solids. The narrow mouth ensures no spirage.

Conical flasks are named according to the maximum volume they can hold e.g.

“250ml” Conical flasks hold a maximum volume of “250mililitres” /“250 cubic centimetres”

“500ml” Conical flasks hold a maximum volume of “500ml” /“1000 cubic centimetres”

5. Round bottomed flask

A round bottomed flask is a moderately narrow glass apparatus with a wide round base and no calibration/graduation. Round bottomed flask thus carry/hold exact volumes of liquids that have been measured using other apparatus. The narrow/thin mouth prevents spirage. The flask can also hold (weighed) solids. A round bottomed flask must be held/ clamped when in use because of its wide narrow base.

6. Flat bottomed flask

A flat bottomed flask is a moderately narrow glass apparatus with a wide round base with a small flat bottom. It has no calibration/graduation.

Flat bottomed flask thus carry/hold exact volumes of liquids that have been measured using other apparatus. The narrow/thin mouth prevents spirage. They can also hold (weighed) solids. A flat bottomed flask must be held/ clamped when in use because it’s flat narrow base is not stable.

**(g) Apparatus for holding unstable apparatus( during heating).**

1. Tripod stand

A tripod stand is a three legged metallic apparatus which unstable apparatus are placed on (during heating).Beakers. conical flasks,round bottomed flask and flat bottomed flasks are placed on top of tripod stand (during heating).

2. Wire gauze/mesh

Wire gauze/mesh is a metallic/iron plate of wires crossings. It is placed on top of a tripod stand:

1. ensure even distribution of heat to prevent cracking glass apparatus
2. hold smaller apparatus that cannot reach the edges of tripod stand 3 Clamp stand

A clamp stand is a metallic apparatus which tightly hold apparatus at their “neck” firmly.

A clamp stand has a wide metallic base that ensures maximum stability. The height and position of clamping is variable. This require practice

4.Test tube holder

A test tube holder is a hand held metallic apparatus which tightly hold test/boiling/ignition tube at their “neck” firmly on the other end.

Some test tube holders have wooden handle that prevent heat conduction to the hand during heating.

5. Pair of tong.

A pair of tong is a scissor-like hand held metallic apparatus which tightly hold firmly a small solid sample on the other end.

6.Gas jar

A gas jar is a long wide glass apparatus with a wide base.

It is open on one end. It is used to collect/put gases.

This requires practice.

**(h) Apparatus for holding/directing liquid solutions/funnels ( to avoid spirage).** 1. Filter funnel

A filter funnel is a wide mouthed (mainly plastic) apparatus that narrow drastically at the bottom to a long extension.

When the long extension is placed on top of another apparatus, a liquid solution can safely be directed through the wide mouth of the filter funnel into the apparatus without spirage.

Filter funnel is also used to place a filter paper during filtration.

2. Thistle funnel

A thistle funnel is a wide mouthed glass apparatus that narrow drastically at the bottom to a very long extension.

The long extension is usually drilled through a stopper/cork.

A liquid solution can thus be directed into a stoppered container without spirage

3. Dropping funnel

A dropping funnel is a wide mouthed glass apparatus with a tap that narrow drastically at the bottom to a very long extension.

The long extension is usually drilled through a stopper/cork.

A liquid solution can thus be directed into a stoppered container without spirage at the rate determined by adjusting the tap.

4. Separating funnel

A separating funnel is a wide mouthed glass apparatus with a tap at the bottom narrow extension.

A liquid solution can thus be directed into a separating funnel without spirage. It can also safely be removed from the funnel by opening the tap. It is used to separate two or more liquid solution mixtures that form layers/immiscibles. This requires practice.

**(h) Apparatus for heating/Burners**

1. Candle, spirit burner, kerosene stove, charcoal burner/jiko are some apparatus that can be used for heating.

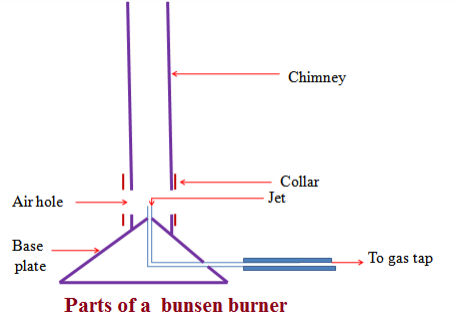
Any flammable fuel when put in a container and ignited can produce some heat.

2.Bunsen burner

The Bunsen burner is the **standard** apparatus for heating in a Chemistry school laboratory.

It was discovered by the German Scientist Robert Wilhelm Bunsen in1854.

(**a**)Diagram of a Bunsen burner



A Bunsen burner uses butane/laboratory gas as the fuel. The butane/laboratory gas is highly flammable and thus usually stored safely in a secure chamber outside Chemistry school laboratory. It is tapped and distributed into the laboratory through gas pipes.

The gas pipes end at the gas tap on a chemistry laboratory bench .If opened the gas tap releases butane/laboratory gas. Butane/laboratory gas has a characteristic odour/smell that alerts leakages/open gas tap.

The Bunsen burner is fixed to the gas tap using a strong rubber tube.

The Bunsen burner is made up of the following parts:

(i)base plate –to ensure the burner can stand on its own

(ii)Jet-a hole through which laboratory gas enters the burner

(iii)Collar/sleeve-adjustable circular metal attached to the main chimney/burell with a side hole/entry. It controls the amount of air entering used during burning.

(iv)Air hole- a hole/entry formed when the collar side hole is in line with chimney side hole. If the collar side hole is **not** in line with chimney side hole, the air hole is said to be “closed” If the collar side hole is **in line** with chimney side hole, the air hole is said to be “open”

(v)Chimney- tall round metallic rod attached to the base plate.

(**b**)**Procedure for lighting/igniting a Bunsen burner** 1. Adjust the collar to ensure the air holes are closed.

2. Connect the burner to the gas tap using a rubber tubing. Ensure the rubber tubing has no side leaks. 3. Turn on the gas tap.

1. Ignite the top of the chimney using a lighted match stick/gas lighter/wooden splint.
2. Do not delay excessively procedure (iv) from (iii) to prevent highly flammable laboratory gas from escaping/leaking.

(**c**)**Bunsen burner flames**

A Bunsen burner produces two types of flames depending on the amount of air entering through the air holes.

If the air holes are **fully open**, a **non luminous** flame is produced. If the air holes are **fully closed**, a **luminous flame** is produced. If the air air holes are **partially** open/ closed, a **hybrid** of non luminous and luminous flames is produced.

**Characteristic differences between luminous and non-luminous flame**

|  |  |
| --- | --- |
| **Luminous flame** | **Non-luminous flame** |
| **1.** Produced when the air holes are fully/completely **closed**. | **1.** Produced when the air holes are fully/completely **open**. |
| **2.** when the air holes are fully/ completely closed there is **incomplete** burning/ combustion of the laboratory gas | **2.**when the air holes are fully/ completely open there is **complete** burning/ combustion of the laboratory gas |
| **3.** Incomplete burning/ combustion of the laboratory gas produces fine unburnt carbon particles which make the flame **sooty/smoky** | **3.** complete burning/ combustion of the laboratory gas does not produce carbon particles. This make the flame non**sooty /non- smoky.** |
| **4.** Some carbon particles become white hot and emit light.This flame is thus bright **yellow** in colour producing **light.** This makes luminous flame useful for **lighting** | **4.** Is mainly **blue** in colour and is **hotter** than luminous flame. This makes nonluminous flame useful for **heating** |
| **5.** Is **larger**, **quiet** and **wavy**/easily swayed by wind | **5.**Is **smaller**, **noisy** and **steady** |
| Luminous flame has **three** main regions:  (i)the top yellow region where there is incomplete combustion/burning  (ii)the region of unburnt gas below the yellow region where the gas does not burn  (iii) blue region on the sides of region of unburnt gas where there is complete burning | Non-luminous flame has **four** main regions:  (i)the top colourless region   1. blue region just below where there is complete burning.It is the hottest region 2. green region surrounded by the blue region where there is complete burning   (ii)the region of unburnt gas at the innermost surrounded by green and blue regions. No burning takes place here |

Scientific apparatus are drawn:

(i)using a proportional **two** dimension(**2D**) cross-sections. Three dimensions (3D) are not recommended.

(ii)straight edges of the apparatus on a scientific diagram should be drawn using ruler.

(iii)curved edges of the apparatus on a scientific diagram should be drawn using free hand.

(iv)The bench, tripod or clamp to support apparatus which cannot stand on their own should be shown.

CLASSIFICATION OF SUBSTANCES

Substances are either pure or impure. A pure substance is one which contains only one substance.

An impure substance is one which contains two or more substances. A pure substance is made up of a pure solid, pure liquid or pure gas.

A mixture is a combination of two or more pure substances which can be separated by physical means. The three states of matter in nature appear mainly as mixtures of one with the other.

Common mixtures include:

**(a)Solutions/solid-liquid dissolved mixture Experiment:**

To make a solution of copper(II)sulphate(VI)/Potassium manganate(VII) /sodium chloride Procedure

Put about 100 cm3 of water in three separate beakers. Separately place a half spatula end full of copper(II)sulphate(VI) ,Potassium manganate(VII) and sodium chloride crystals to each beaker. Stir for about two minutes.

Observation

Copper(II)sulphate(VI) crystals dissolve to form a blue solution

Potassium manganate(VII) crystals dissolve to form a purple solution

Sodium chloride crystals dissolve to form a colourless solution

Explanation

Some solids, liquids and gases dissolve in some other liquids.

A substance/liquid in which another substance dissolves is called solvent.

A substance /solid /gas which dissolves in a solvent is called solute.

When a solute dissolves in a solvent it forms a uniform mixture called **solution**. A solute dissolved in water as the solvent exists in another state of matter called **aqueous state.**Water is refered as the **universal solvent** because it dissolves many solutes. A solute that dissolves in a solvent is said to be **soluble**. Soluble particles uniformly spread between the particles of water/solvent and cannot be seen.

**Solute + Solvent -> solution**

**Solute + Water -> Aqueous solution of solute**

The solute dissolved in water gives the **name** of the solution

e. g.

1. Sodium chloride solution is a solution formed after dissolving sodium chloride crystals/solid in water. Sodium chloride exists in aqueous state after dissolving.

Sodium chloride + Water -> Sodium chloride solution

NaCl(**s**) + (aq) -> NaCl(**aq**)

1. Ammonia solution is a solution formed after dissolving ammonia gas in water.

Ammonia exists in aqueous state after dissolving.

Ammonia gas + Water -> Aqueous ammonia

NH3(**g**) + (aq) -> NH3(**aq**)

1. Copper (II)sulphate(VI) solution is a solution formed after dissolving Copper(II) sulphate (VI) crystals/solid in water. Copper (II)sulphate(VI) exist in aqueous state after dissolving.

Copper (II)sulphate(VI) + Water -> Copper (II)sulphate(VI) solution CuSO4(**s**) + (aq) -> CuSO4 (**aq**)

1. Potassium manganate(VII) solution is a solution formed after dissolving Potassium manganate(VII) crystals/solid in water.

Potassium manganate(VII)exist in aqueous state after dissolving.

Potassium manganate(VII) + Water -> Potassium manganate(VII) solution KMnO4(**s**) + (aq) -> KMnO4 (**aq**)

**(b)Suspension/ precipitates/solid-liquid mixture which do not dissolve**

**Experiment: To make soil,flour and Lead(II)Iodide suspension/precipitate** Procedure

Put about 100 cm3 of water in three separate beakers. Separately place a half spatula end full of soil ,maize and lead(II)Iodide to each beaker. Stir for about two minutes.

Observation

Some soil , maize and lead(II)Iodide float in the water

A brown suspension/precipitate/particles suspended in water containing soil

A white suspension/precipitate/particles suspended in water containing flour A yellow suspension/precipitate/particles suspended in water containing Lead(II)iodide.

Some soil , maize and lead(II)Iodide settle at the bottom after some time.

Explanation

Some solid substances do not dissolve in a liquid. They are said to be **insoluble** in the solvent .When an insoluble solid is put in liquid:

* 1. some particles remain **suspended/floating** in the liquid to form a **suspension** /**precipitate**.
  2. some particles **sink/settle** to the bottom to form **sediments** after being allowed to stand .

An **insoluble** solid acquire the colour of the suspension/precipitate .e.g .

1.A white suspension /precipitate has some fine white particles suspended /floating in the liquid. **Not** “white solution”

2.A blue suspension /precipitate has some fine **blue** particles suspended /floating in the liquid.

3.A green suspension /precipitate has some fine **green** particles suspended /floating in the liquid.

4.A **brown** suspension /precipitate has some fine **brown** particles suspended /floating in the liquid.

4.A yellow suspension /precipitate has some fine yellow particles suspended /floating in the liquid.

(c) (i) Miscibles /Liquid-liquid mixtures

**To form water-ethanol and Kerosene-turpentine miscibles** Procedure

(i)Measure 50cm3 of ethanol into 100cm3 beaker. Measure 50cm3 of water. Place the water into the beaker containing ethanol. Swirl for about one minute. (ii)Measure 50cm3 of kerosene into 100cm3 beaker. Measure 50cm3 of turpentine oil. Place the turpentine oil into the beaker containing kerosene. Swirl for about one minute.

Observation

Two liquids do not form layers.

Ethanol and water form a uniform mixture.

Kerosene and turpentine oil form uniform mixture

Explanation

Ethanol is miscible in Water. Kerosene is miscible in turpentine oil. Miscible mixture form uniform mixture. They do not form layers. The particles of one liquid are smaller than the particles of the other. The smaller particles occupy the spaces between the bigger particles.

(ii) Immiscibles /Liquid-liquid mixtures

**To form water-turpentine oil and Kerosene-water miscibles**

Procedure

(i)Measure 50cm3 of water into 100cm3 beaker. Measure 50cm3 of turpentine oil.

Place the oil into the beaker containing water. Swirl for about one minute.

(ii) Measure 50cm3 of water into 100cm3 beaker. Measure 50cm3 of kerosene.

Place the kerosene into the beaker containing water. Swirl for about one minute.

Observation

Two liquids form layers.

Turpentine and water do not form a uniform mixture.

Water and kerosene do not form uniform mixture

Explanation

Kerosene is immiscible in Water. Water is immiscible in turpentine oil. Immiscible mixtures do not form uniform mixtures. They form layers. The size of the particles of one liquid is almost equal to the particles of the other. The particles of one liquid cannot occupy the spaces between the particles of the other. The heavier particles settle at the bottom. The less dense particles settle on top.

(d)Solid-solid mixtures/Alloys

Before solidifying, some heated molten/liquid metals dissolve in another metal to form a uniform mixture of the two. On solidifying, a uniform mixture of the metals is formed. A uniform mixture of two metals on solidifying is called **alloy**. In the alloy, one metallic particle occupies the spaces between the metallic particles of the other.

**c) Common alloys of metal.**

|  |  |  |
| --- | --- | --- |
| **Alloy name** | **Constituents of the alloy** | **Uses of the alloy** |
| Brass | Copper and Zinc | Making scews and bulb caps |
| Bronze | Copper and Tin | Making clock springs,electrical  contacts and copper coins |
| Soldier | Lead and Tin | Soldering, joining electrical contacts because of its low melting points and high thermal conductivity |
| Duralumin | Aluminium, Copper and Magnesium | Making aircraft, utensils, windows frames because of its light weight and corrosion resistant. |
| Steel | Iron, Carbon  ,Manganese and other metals | Railway lines, car bodies girders and utensils. |
| Nichrome | Nichrome and Chromium | Provide resistance in electric heaters and ovens |
| German silver | Copper, Zinc and Nickel | Making coins |

**METHODS OF SEPARATING MIXTURES**

Mixtures can be separated from applying the following methods:

**(a) Decantation**

Sediments can be separated from a liquid by pouring out the liquid. This process is called **decantation.**

Experiment

Put some sand in a beaker. Add about 200cm3 of water. Allow sand to settle. Pour off water carefully into another beaker.

Observation

Sand settles at the bottom as sediments.

Less clean water is poured out.

Explanation

Sand does not dissolve in water. Sand is denser than water and thus settles at the bottom as **sediment**. When poured out, the less dense water flows out.

**(b)Filtration**

Decantation leaves suspended particles in the liquid after separation. Filtration is thus improved decantation.

Filtration is the method of separating insoluble mixtures/particles/solids from a liquid.

Experiment **: To separate soil and water using filtration**

Fold a filter paper to fit well into a filter funnel. Place the funnel in an empty 250 cm3 beaker.

Put one spatula end full of soil into 50cm3 of water. Stir. Put the soil/water mixture into the filter funnel.

Observations

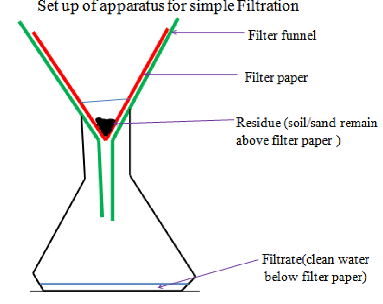
Clean water is collected **below** the filter funnel.

Soil remains **above** the filter paper.

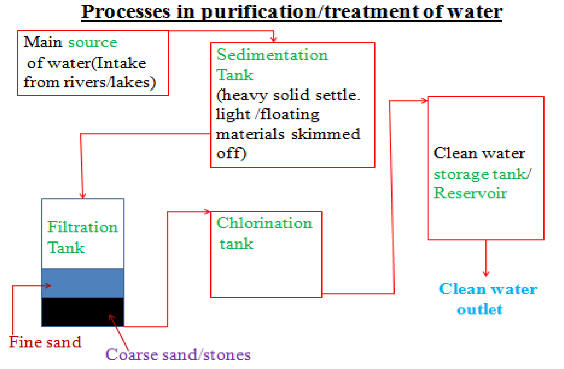
Explanation

A filter paper is **porous** which act like a fine sieve with very small **holes**. The holes allow smaller water particles to pass through but do not allow bigger soil particles. The liquid which passes through is called **filtrate**. The solid which do not pass through is called **residue**.

Set up of apparatus



In industries, filtration is used in engine filters to clean up air.



**(c)Evaporation**

Evaporation is a method of separating a solute/solid from its solution. This involves heating a solution (solvent and solute)to vapourize the solvent out of the solution mixture leaving pure solute/solid. If a mixture contain insoluble solid, they are filtered out.

Experiment: **: To separate a mixture of soil and salt(sodium chloride) .**

Procedure:

Put one spatula end full of soil on a filter paper.

Put one spatula full of common salt/sodium chloride into the same filter paper. Mix well using the spatula,.

Place about 200cm3 of water into a beaker.

Put the contents of the filter paper into the water. Stir thoroughly using a glass/stirring rod for about one minute.

Fold a filter paper into a filter funnel.

Pour half portion of the contents in the beaker into the filter funnel. Put the filtrate into an evaporating dish. Heat on a water bath.

Observation

(i)On mixing

Colourless crystals and brown soil particles appear on the filter paper.

(ii)On adding water

Common soil dissolves in water. Soil particles do not dissolve in water.

(iii)On filtration

Colourless liquid collected as filtrate below the filter funnel/paper.

Brown residue collected above the filter funnel/paper. (iv)On evaporation

Colourless crystals crystals collected after evaporation

Explanation

Solid mixture of sand and common salt take the colours of the two.

On adding water, common salt dissolve to form a solution .

Soil does not because it is insoluble in water and thus forms a suspension.

On filtration, a residue of insoluble soil does not pass through the filter paper.

It is collected as residue.

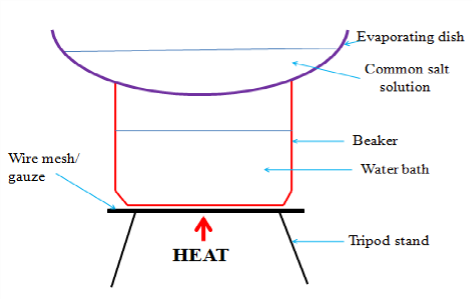
Common salt solution is collected as filtrate.

On heating the filtrate, the solvent/water evaporate/vapourize out of the evaporating dish leaving common salt crystals.

Vapourization/evaporation can take place even without heating.

This is the principle/process of drying wet clothes on the hanging line.

Set up of apparatus



**(d) Distillation**

Distillation is an improved evaporation where both the solute and the solvent in the solution are separated /collected. Distillation therefore is the process of separating a solution into constituent solid solute and the solvent. It involves heating the solution to evaporate/vapourize the solvent out. The solvent vapour is then condensed back to a liquid.

**Experiment: To obtain copper(II)sulphate (VI) crystals and water from copper (II) sulphate(VI) solution.**

Procedure:

Put one spatula end full of copper(II)sulphate (VI) crystals into a 250cm3 beaker.

Place about 200cm3 of water into the beaker.

Stir thoroughly using a glass/stirring rod for about one minute.

Pour half portion of the contents in the beaker into a round bottomed/flat/conical flask broken porcelain/sand/glass into the flask.

Put a few pieces of b Stopper the flask.

Connect the flask to a liebig condenser using delivery tube.

Place a 200cm3 clean empty beaker/conical flask as a receiver at the end of the liebig condenser.

Circulate water in the liebig condenser.

Heat the flask strongly on a tripod stand with wire mesh/gauze until there is no more visible **boiling bubbles** in the flask.

Observation

Copper (II)sulphate (VI) crystals dissolve in water to form a blue solution.

On heating, colourless liquid is collected in the receiver.

Blue crystals are left in the flask.

(if gently heated further, the blue crystals turn to white powder)

Explanation

On heating blue Copper (II)sulphate (VI) solution, the colourless liquid solvent evaporate/vapourize .

The liquid vapour/gas passes through the delivery tube to the liebig condenser. The liebig condenser has a cold water **inlet** near the receiver and cold water **out** let.

This ensures efficient cooling. If the cold water **outlet/inlet** is reversed, the water circulation would be less efficient.

The water in the receiver would be warm.In the liebig condenser, the cold water,condenses the liquid vapour into liquid.

The condensed liquid collects in the receiver as **distillate**.

The solute of blue Copper (II)sulphate (VI) crystals is left in the flask as **residue**. During simple distillation,therefore, the solution is heated to vapourize /evaporate the solvent/one component which is condensed at a different part of the apparatus.

The purpose of pieces of broken porcelain/porous pot/glass/sand/ is to:

(i)prevent bumping of the solution during boiling.

(ii)ensure smooth and even boiling.

Salty sea water can be made pure through simple distillation.

Any mixture with a large difference /40oC in boiling point can be separated using simple distillation.

Set up of apparatus



**(e)Fractional distillation**

Fractional distillation is an improved simple distillation used specifically to separate miscible mixtures with very **close /near** boiling points.

Fractional distillation involves:

(**i**)Heating the mixture in a conical/round bottomed /flat bottomed flask. The pure substance with a lower boiling point and thus more volatile evaporates/boils/vapourizes first. e.g.

Pure ethanol has a boiling point of 78oC.Pure water has a boiling point of 100 oC at sea level/one atmosphere pressure.

When a miscible mixture of ethanol and water is heated, ethanol vapourizes /boils/ evaporates first because it is more volatile.

(**ii**)The conical/round bottomed /flat bottomed flask is connected to a long glass tube called **fractionating column**.

The purpose of the fractionating column is to offer areas of condensation for the less volatile pure mixture.

The fractionating column is packed with glass beads/broken glass/ porcelain/ shelves to increase the surface area of condensation of the less volatile pure mixture.

(**iii**)When the vapours rise they condense on the glass beads/broken glass /porcelain / shelves which become hot.

When the temperature of the glass beads/broken glass/porcelain/shelves is beyond the boiling point of the less volatile pure substance, the pure substance rise and condensation take place on the glass beads/broken glass/porcelain/shelves at a higher level on the fractionating column.

The less volatile pure substance trickles/drips back down the fractionating column or back into the conical/round bottomed /flat bottomed flask to be heated again. e.g.

If the temperature on glass beads/broken glass/porcelain/shelves is beyond 78oC, the **more volatile** pure ethanol rise to condense on the glass beads/broken glass /porcelain/shelves **higher** in the fractionating column.

Water condenses and then drip/trickle to the glass beads/broken glass /porcelain /shelves **lower** in the fractionating column because it is **less volatile.**

(**iv**)The fractionating column is connected to a liebig condenser. The liebig condenser has a cold water inlet and outlet circulation.

The more volatile mixture that reach the top of the fractionating column is condenses by the liebig condenser into a receiver. It is collected as the first fraction.

(**v**)At the top of the fractionating column, a thermometer is placed to note/monitor the temperature of the boiling mixtures .

Pure substances have constant/fixed boiling point. When one mixture is completely separated, the thermometer reading rises.

e.g. The thermometer reading remains at78oC when ethanol is being separated. When no more ethanol is being separated, the mercury/alcohol level in the thermometer rises.

(**vi**)The second /subsequent fractions are collected in the receiver after noting a rise the mercury/alcohol level in the thermometer. e.g.

The thermometer reading rises to 100oC when water is being separated. It is passed through the liebig condenser with the cold water inlet and outlet circulation.

It is collected different receiver as the second/subsequent fraction.

(vii)Each fraction collected should be confirmed from known physical/chemical properties/characteristic. e.g.

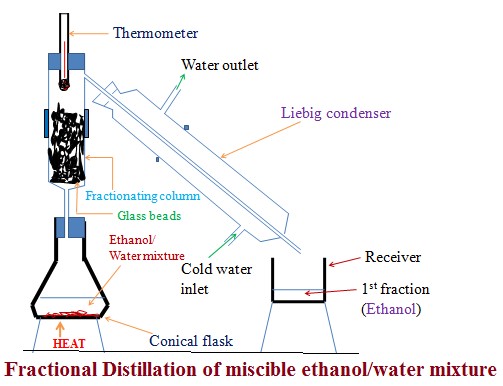
Ethanol

Ethanol is a colourless liquid that has a characteristic smell .When it is put in a watch glass then ignited, it catches fire and burn with a blue flame.

Water

Water is a colourless liquid that has no smell/odour .When it is put in a watch glass then ignited, it does not catch fire.

Set up of apparatus



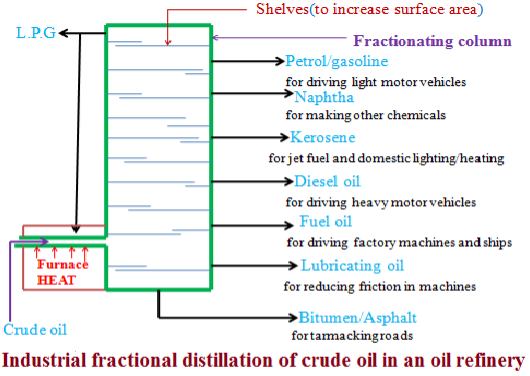
**Industrial application of Fractional distillation** On a large scale,fractional distillation is used:

**(i)**In fractional distillation of crude oil in an oil refinery.

Crude oil is a mixture of many fractions. When heated in a furnace, the different fractions separate out according to their boiling point. In Kenya,fractional distillation takes place at Changamwe in Mombasa.

**(ii)**In fractional distillation of air.

Air contain a mixture of three main useful gases which are condensed by coolin to very low temperature (-200oC) to form a liquid. The liquid is then heated. Nitrogen is the most volatile(-196 oC) and thus comes out as the first fraction. Argon (at -186 oC) is the second fraction. Oxygen ( at -183 oC) is the last fraction. The three gases are very useful industrial gases.



**(f)Separation of immiscibles (Using a separating funnel)**

Two or more liquids that form layers on mixing are immiscible. Immiscible mixture arrange themselves according to their densities

i.e The denser liquid sink to the bottom. The less dense liquid floats on the denser one. Immicible mixtures can be separated from each other by using a **separating funnel**.

**Experiment:** To separate an immiscible mixture of paraffin and water.

Procedure

Place about 100cm3 of water into a 250cm3 beaker. Add about 100cm3 of paraffin into the beaker. Stir.

Transfer the mixture into a separating funnel. Allow to settle for about one minute. Open the tap, run out the lower layer out slowly into a clean beaker. Close the tap when the upper layer is very close to the tap.

Run out the intermediate small amount of the mixture near the tap into a beaker.

Discard it.

Run out the remaining upper layer into a fresh beaker.

Place a portion of upper and lower layer into a watch glass separately after separating each. Ignite.

Observation

Water and paraffin are both colourless liquids.

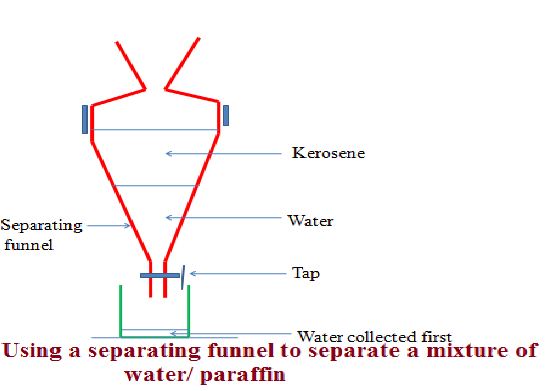
Two layers are formed on mixing.

Colourless odourless liquid collected first. It does not catch fire.

A colourless liquid with characteristic smell collected later/second. It catches fire and burn with a yellow smoky flame.

Explanation

Water and paraffin are immiscible. Water is denser than paraffin. When put in a separating funnel, paraffin float on water. On opening the tap, water runs out. A mixture of water and paraffin at the junction of the two is discarded. It is not pure. Set up of apparatus



**(g)Sublimation/deposition**

Some solids on heating do not melt to a liquid but change directly to a gas. The process by which a solid changes to a gas is called **sublimation**. The gas cools back and changes directly to a solid. The process by which a gas changes to a solid is called **deposition.** Sublimation and deposition therefore are the same but opposite processes.

Sublimation

Deposition

GAS

SOLID

Some common substances that undergo sublimation/ deposition include:

(i)Iodine (ii)Carbon(IV)oxide (iii)Camphor

(iv) ammonium chloride (v)Iron(III)chloride (vi)Aluminium(III)chloride (vii) benzoic acid

If a mixture has any of the above as a component, then on heating it will change to a gas and be deposited away from the source of heating.

Procedure

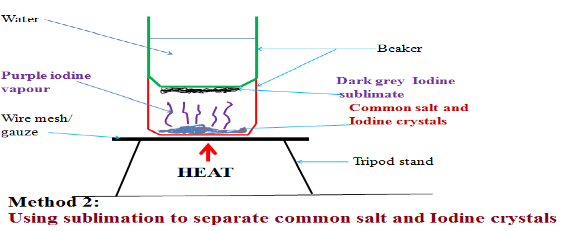
Place about one spatula full of ammonium chloride crystals into a clean dry 100cm3 beaker. Add equal amount of sodium chloride crystals into the beaker.

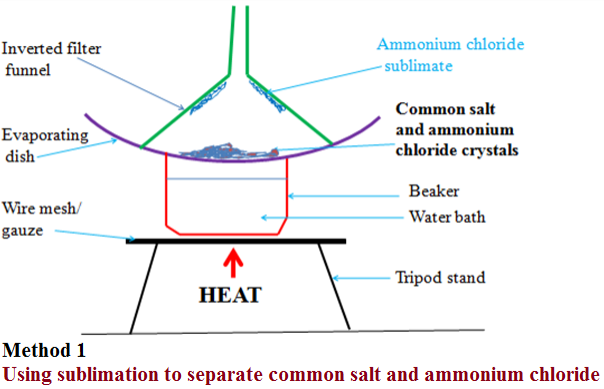
Swirl to mix.

Place the beaker on a tripod stand.

Put about 100cm3 of water into another beaker. Place carefully the beaker containing water on top of the beaker containing the solid mixture. Light/ignite a burner and heat the solid.

Set up of apparatus:





Observation

(i)With ammonium chloride/common salt mixture White fumes produced .

White sublimate deposited

Colourless residue left

(ii)With Iodine/common salt mixture Purple fumes produced .

Dark grey sublimate deposited

Colourless residue left

Explanation

(i)On heating a mixture of ammonium chloride and common salt, a white fumes of ammonium chloride is produced. The white fumes solidify as white sublimate on the cooler parts. Common salt remains as residue.

Chemical equation:

Ammonium chloride **solid** Ammonium chloride **gas**

NH4Cl(**s**) NH4Cl(**g**)

(ii)On heating a mixture of Iodine and common salt, a purple fumes of Iodine vapour is produced. The purple fumes solidify as dark grey sublimate on the cooler parts. Common salt remains as residue.

Chemical equation:

Iodine **solid** Iodine **gas**

I2(**s**) I2 (**g**)

**(h)Chromatography**

Chromatography is a method of separating components of a solution mixture by passing it through a medium where the different components move at different rates. The medium through which the solution mixture is passed is called **absorbent material**.

Paper chromatography is a method of separating coloured dyes by using paper as the absorbent material.

Since dyes are insoluble/do not dissolve in water, ethanol and propanone are used as suitable solvents for dissolving the dye.

Practically, a simple paper chromatography involve placing a dye/material on the absorbent material, adding slowly a suitable soluble solvent on the dye/material using a dropper, the solvent spread out on the absorbent material carrying the soluble dye away from the origin.

The spot on which the dye is initially/originally placed is called **baseline**. The farthest point the solvent spread is called **solvent front**. The farthest a dye can be spread by the solvent depend on:

1. density of the dye-the denser the dye, the less it spread from the basely ne by the solvent.
2. Stickiness of the dye-some dyes sticks on the absorbent material more than other thus do not spread far from baseline. **Experiment: To investigate the colours in ink**

Procedure

**Method 1**

Place a filter paper on a an empty beaker. Put a drop of black/blue ink in the centre of the filter paper. Wait for about one minute for the ink drop to spread. Using a clean teat pipette/dropper add one drop of ethanol/propanone. Wait for about one minute for the ink drop to spread further. Add about twenty other drops of ethanol waiting for about one minute before each addition. Allow the filter paper to dry.

**Experiment: To investigate the colours in ink**

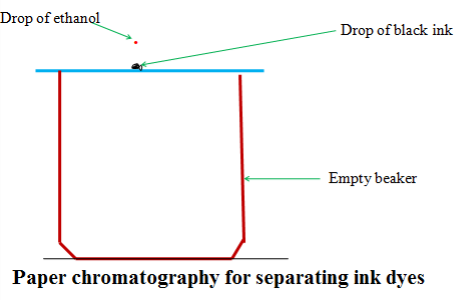
Procedure

**Method 2**

Cut an 8 centimeter thin strip of a filter paper. At about 3cm on the strip, place a drop of ink. Place the filter paper in a 10cm length boiling tube containing 5cm3 of ethanol. Ensure the cut strip of the filter paper just dips into the ethanol towards the ink mark. Cover the boiling tube. Wait for about twenty minutes. Remove the boiling tube and allow the filter paper to dry.

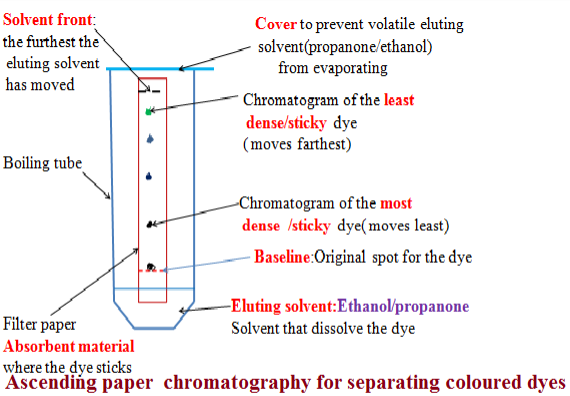
Set up of apparatus

Method 1



Set up of apparatus

Method 2



Explanation

When a drop of ink is placed on an absorbent material it sticks. On adding an eluting solvent, it dissolves the dye spread out with it. The denser and sticky pure dye move least. The least dense/sticky pure dye move farthest. A pure dye will produce the same chromatogram/spot if the same eluting solvent is used on the same absorbent material. Comparing the distance moved by a pure dye with a mixture ,the coloured dyes in a mixture can be deduced as below:

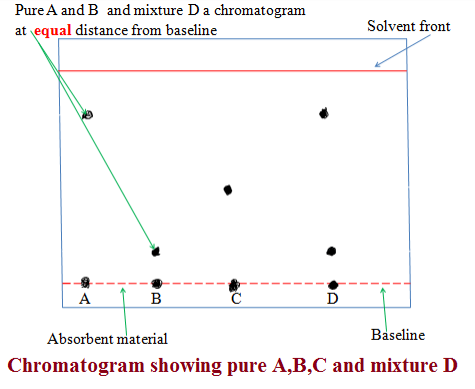
**Example 1**

The chromatogram of pure dyes A, B ,C and a dye mixture D is shown below Determine the pure dyes present in D. On the diagram show:

(i)the solvent front

(ii)baseline

(iii)the most soluble pure dye



**(i) Solvent extraction**

Solvent extraction is a method of separating oil from nuts/seeds. Most nuts contain oil. First the nuts are crushed to reduce their size and increase the surface area. A suitable volatile solvent is added. The mixture is filtered. The filtrate solvent is then allowed to crystallize leaving the oil/fat. If a filter paper is rubbed/smeared with the oil/fat, it becomes translucent. This is the test for the presence of oil/fat.

**Experiment: To extract oil from Macadamia nut seeds**

Procedure

Crush Macadamia nut seeds form the hard outer cover .Place the inner soft seed into a mortar. Crush(add a little sand to assist in crushing).

Add a little propanone and continue crushing. Continue crushing and adding a little propanone until there is more liquid mixture than the solid. Decant/filter. Put the filtrate into an evaporating dish. Vapourize the solvent using solar energy /sunlight. Smear/rub a portion of the residue left after evaporation on a clean dry filter paper.

Observation /Explanation

Propanone dissolve fat/oil in the macadamia nuts. Propanone is more volatile(lower boiling point)than oil/fat. In sunlight/solar energy, propanone evaporate/vapourize leaving oil/fat(has a higher boiling point).Any seed like corn, wheat , rice, soya bean may be used instead of macadamia seed. When oil/fat is rubbed/ smeared on an opaque paper, it becomes translucent.

**(j) Crystallization**

Crystallization is the process of using solubility of a solute/solid to obtain the solute/solid crystals from a saturated solution by cooling or heating the solution. A crystal is the smallest regular shaped particle of a solute. Every solute has unique shape of its crystals.

Some solutions form crystals when heated. This is because less solute dissolve at higher temperature. Some other solutions form crystals when cooled. This is because less solute dissolve at lower temperature.

**Experiment; To crystallize copper(II)sulphate(VI)solution**  Procedure:

Place about one spatula full of hydrated copper sulphate(VI) crystals into 200cm3 of distilled water in a beaker. Stir. Continue adding a little more of the hydrated copper sulphate (VI) crystals and stirring until no more dissolve. Decant/filter. Cover the filtrate with a filter paper. Pierce and make small holes on the filter paper cover. Preserve the experiment for about seven days.

Observation/Explanation

Large blue crystals formed

When hydrated copper(II)sulphate crystals are placed in water, they dissolve to form copper(II)sulphate solution. After some days water slowly evaporate leaving large crystals of copper(II)sulphate. If the mixture is heated to dryness, small crystals are formed.

**Physical/Temporary and Chemical changes**

A physical/temporary change is one which **no new** substance is formed and is **reversible** back to original.

A chemical/permanent change is one which **a new** substance is formed and is **irreversible** back to original.

The following experiments illustrates physical and chemical changes

**(a)Heating ice**

Place about 10g of pure ice in a beaker. Determine its temperature.Record it at time “0.0” in the table below. Heat the ice on a strong Bunsen flame and determine its temperature after every 60seconds/1minute to complete the table below:

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Time/minutes | 0 | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 |
| Temperature (oC) | -2 | 0 | 0 | 40 | 80 | 90 | 95 | 95 | 96 |

Plot a graph of time against Temperature(y-axes)

Explain the shape of your graph

**Melting/freezing/fusion/solidification** and **boiling /vaporization /evaporation** are the two physical processes.

Melting /freezing point of pure substances is fixed /constant.

The boiling point of pure substance depend on **external** atmospheric **pressure**.

Melting/fusion is the physical change of a **solid** to **liquid**.

Freezing is the physical change of a **liquid** to **solid**.

Melting/freezing/fusion/solidification are therefore two **opposite** but **same** reversible physical processes i.e

1. (**s**) A(**l**)

Boiling/vaporization/evaporation is the physical change of a **liquid** to **gas**.

Condensation/ liquidification is the physical change of **gas** to **liquid**.

Boiling/vaporization/evaporation and condensation/ liquidification are therefore two **opposite** but **same** reversible physical processes i.e

1. (**l**) B(**g**)

Practically

(i) Melting/liquidification/fusion involves **heating** a solid to **weaken** the strong bonds holding the solid particles together.

Solids are made up of very strong bonds holding the particles **very close** to each other (**Kinetic Theory of matter**).

On heating these particles gain energy/heat from the surrounding heat source to form a liquid with **weaker** bonds holding the particles close together but with some degree of **freedom**.

(ii)Freezing/fusion/solidification involves cooling a liquid to reform /rejoin the very strong bonds to hold the particles **very close** to each other as solid and thus lose their degree of **freedom** (**Kinetic Theory of matter**).

Freezing /fusion / solidification is an **exothermic** (**-**∆H)process that require particles holding the liquid together to lose energy to the surrounding. (iii)Boiling/vaporization/evaporation involves **heating** a liquid to completely **break/free** the bonds holding the liquid particles together.

Gaseous particles have high degree of **freedom** (**Kinetic Theory of matter**). Boiling /vaporization / evaporation is an **endothermic** (**+**∆H) process that require/absorb energy from the surrounding.

(iv)Condensation/liquidification is **reverse** process of boiling /vaporization / evaporation.

It involves gaseous particles losing energy to the surrounding to form a liquid.

**AIR AND COMBUSTION**

**A.THE ATMOSPHERE .**

1.The atmosphere is made up of air. Air is a mixture of colourless , odourless gases which is felt as wind(air in motion).All living things breath in air for respiration . Plants use air for respiration and photosynthesis.

2.The main gases present in the atmosphere/air:

|  |  |
| --- | --- |
| Gas | Approximate % composition by volume |
| Nitrogen | 78.0 |
| Oxygen | 21.0 |
| Carbon(IV)oxide | 0.03 |
| Noble gases | 1.0 |
| Water vapour | Vary from region |

3. The following experiments below shows the presence and composition of the gases in air/atmosphere

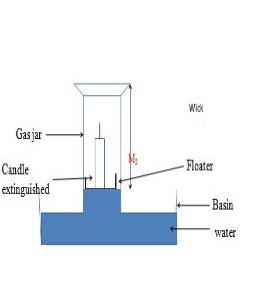
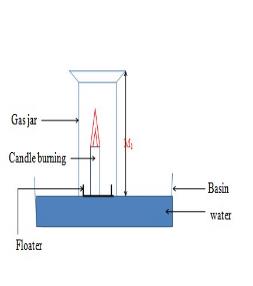
**(a)To find the composition of air supporting combustion using a candle stick**

Procedure

Measure the length of and empty gas jar M1. Place a candle stick on a petri dish. Float it on water in basin/trough. Cover it with the gas jar. Mark the level of the water in the gas jar M2. Remove the gas jar. Light the candle sick. Carefully cover it with the gas jar. Observe for two minutes. Mark the new level of the water M3.

Set up of apparatus

Sample observations



Candle continues to burn then extinguished/goes off

Level of water in the gas jar rises after igniting the candle

Length of empty gas jar = M1= 14cm

Length of gas jar **without** water before igniting candle = M2= 10 cm

Length of gas jar **with** water before igniting candle = M1 - M2= 14- 10 = 4 cm

Length of gas jar **with** water after igniting candle = M3 = 8 cm

Length of gas jar **without** water after igniting candle = M1 - M3 = 10 -8 = 2 cm

Explanation

Candle burns in air. In a closed system(vessel),the candle continues to burn using the part of air that support burning/combustion. This is called the **active part of air**.The candle goes off/extinguished when all theactive part of air is used up.The level of the water rises to occupy the space /volume occupied by the usedactive part of air.

The experiment is better when very dilute **sodium/potassium hydroxide** is used instead of water . Dilute Potassium/ sodium hydroxide absorb **Carbon(IV)oxide** gas that come out from burning/combustion of candle stick. From the experiment above the % composition of the:

(i)active part of air can be calculated:

M2 - M3 x 100% => 10- 8 x 100% = **20%**

M2 10cm

(ii)inactive part of air can be calculated:

100% -**20% = 80% //** M3 => 8 x 100% **= 80%**

M2 10cm

**(b)To find the composition of active part of air using heated copper turnings.**

Procedure

Clamp a completely packed/filled open ended glass tube with copper turnings. Seal the ends with glass/cotton wool.

Label two graduated syringes as “A” and “B” Push out air from syringe “A”. Pull in air into syringe “B”.

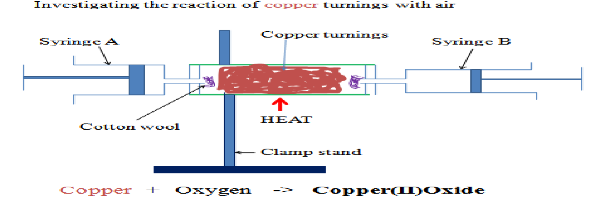
Attach both syringe “A” and “B” on opposite ends of the glass tube.

Determine and record the volume of air in syringe “B” V1. Heat the glass tube strongly for about three minutes.

Push all the air slowly from syringe “B” to syringe “A” as heating continues. Push all the air slowly from syringe “A” back to syringe “B” and repeatedly back and forth.

After about ten minutes, determine the new volume of air in syringe “B” V2

Set up of apparatus



**Sample observations**

Colour change from brown to black

Volume of air in syringe “B” before heating V1  = 158.0cm3

Volume of air in syringe “B” after heating V2 = 127.2cm3

Volume of air in syringe “B” used by copper V1 - V2 = 30.8cm3

**Sample questions**

**1.What is the purpose of**

1. **glass/cotton wool**

To prevent/stop copper turnings from being blown into the syringe/out of the glass tube

1. **passing air through the glass tube repeatedly** To ensure all the active part of air is used up

1. **passing air through the glass tube slowly**

To allow enough time of contact beteewn the active part of and the heated copper turnings.

1. **State and explain the observations made in the glass tube.**

Colour change from brown to black

Brown copper metal reacts with the active part of air/oxygen to form black copper(II)oxide. Chemical equation

Copper + Oxygen -> Copper(II)oxide

2Cu(s) + O2(g) -> 2CuO(s)

The reaction reduces the amount/volume of oxygen in syringe “B” leaving the inactive part of air. Copper only react with oxygen when heated.

1. **Calculate the % of**

**(i)active part of air**

% active part of air = V1 - V2 x 100% => 30.8cm3 x 100% = **19.493**%

V1 158.0cm3

**(ii) inactive part of air** Method 1

% inactive part of air = V2 x 100% =>127.2cm3 x 100% = **80.506**%

V1 158.0cm3

Method 2

% inactive part of air = 100% -% active part of air

=> 100 % - 19.493 % = **80.507**%

**4.The % of active part of air is theoretically higher than the above while % of inactive part of air is theoretically lower than the above. Explain.**

Not all the active part of air reacted with copper

**5.State the main gases that constitute:**

**(a)active part of air.**

Oxygen

(b) **inactive part of air**

Nitrogen, carbon(IV)oxide and noble gases

**6.If the copper turnings are replaced with magnesium shavings the % of active part of air obtained is extraordinary very high. Explain.**

Magnesium is more reactive than copper. The reaction is highly exothermic. It generates enough heat for magnesium to react with both oxygen and nitrogen in the air.

A white solid/ash mixture of Magnesium oxide and Magnesium nitride is formed. This considerably reduces the volume of air left after the experiment.

Chemical equation

Magnesium + Oxygen -> magnesium (II)oxide

2Mg(s) + O2(g) -> 2MgO(s)

Magnesium + Nitrogen -> magnesium (II)nitride

3Mg(s) + N2(g) -> Mg3N2 (s)

**(c)To find the composition of active part of air using alkaline pyrogallol.**

Procedure

Measure about 2cm3 of dilute sodium hydroxide into a graduated gas jar. Record the volume of the graduated cylinder V1.

Place about two spatula end full of pyrogallol/1,2,3-trihydroxobenzene into the gas jar. Immediately place a cover slip firmly on the mouth of the gas jar. Swirl thoroughly for about two minutes.

Invert the gas jar in a trough/basin containing water. Measure the volume of air in the gas jar V2

Sample observations

Colour of pyrogallol/1,2,3-trihydroxobenzene change to **brown**.

Level of water in gas jar rises when inverted in basin/trough. Volume of gas jar /air in gas jar V1= **800cm3**

Volume of gas jar /air in gas jar after shaking with alkaline pyrogallol/1,2,3trihydroxobenzene V2= **640 cm3**

Sample questions

1. **Which gas is absorbed by alkaline pyrogallol/1,2,3-trihydroxobenzene**

Oxygen

1. **Calculate the** 
   1. **% of active part of air**

V1-V2 x 100% => (800cm3 - 640 cm3) x 100% = **20%**

V1 800cm3

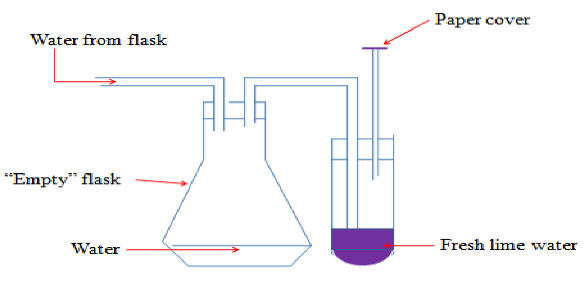
* 1. **% of inactive part of air**

V2 x 100% => 640 cm3 x 100% = **80%**

V1 800cm3

**(d)To establish the presence of carbon(IV)oxide in air using lime water**

Pass tap water slowly into an empty flask as in the set up below



Sample observation questions

**1.What is the purpose of paper cover?**

To ensure no air enters into the lime water.

1. **What happens when water enters the flask?**

It forces the air from the flask into the lime water.

1. **What is observed when the air is bubbled in the lime water**

A white precipitate is formed. The white precipitate dissolves on prolonged bubbling of air.

1. **(a) Identify the compound that form:**

**(i)lime water**

Calcium hydroxide / Ca(OH)2

**(ii)white precipitate**

Calcium carbonate/ CaCO3

**(iii)when the white precipitate dissolves**

Calcium hydrogen carbonate/ CaHCO3

**(b)Write the chemical equation for the reaction that tale place when:**

* 1. **white precipitate is formed**

Calcium hydroxide **+** carbon(IV)oxide **->** Calcium carbonate + water

Ca(OH)2(aq) + CO2 (g) -> CaCO3(s) + H2O(l)

* 1. **white precipitate dissolves**

Calcium carbonate + water**+** carbon(IV)oxide -> Calcium hydrogen carbonate CaCO3(s) + H2O(l) + CO2 (g) -> CaHCO3(aq)

1. **State the chemical test for the presence of carbon (IV)oxide gas based on 4(a) and (b)above:**

Carbon(IV)oxide forms a white precipitate with lime water that dissolves in excess of the gas.

1. **State the composition of carbon(IV)oxide gas by volume in the air.**

About 0.03% by volume

**B.OXYGEN.**

**a) Occurrence.**

1.Fifty 50% of the earths crust consist of Oxygen combined with other elements e.g.oxides of metals

2.About 70% of the earth is water made up of Hydrogen and Oxygen.

3.About 20% by volume of the atmospheric gases is Oxygen that form the active part of air.

**b)School laboratory preparation.**

Oxygen was first prepared in 1772 by Karl Scheele and later in 1774 by Joseph

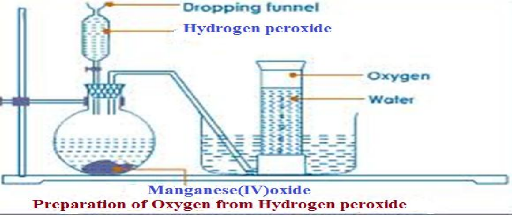
Priestly.It was Antony Lavoisier who gave it the name “Oxygen”

Procedure

Method 1: **Using Hydrogen peroxide**

Half fill a trough/basin with tap water. Place a bee hive shelf/stand into the water.

Completely fill the a gas jar with water and invert in onto the bee hive shelf/stand. Clamp a round bottomed flask and set up the apparatus as below.



Collect several gas jars of Oxygen covering each sample.

Sample observation questions

1.**What is observed when the hydrogen peroxide is added into the flask**

Rapid effervescence/bubbling/fizzing

**2.Describe the colour and smell of the gas**  Colourless and odourless.

**3.(a)Name the method of gas collection used.**

**-**Over water

-Upward delivery

-Down ward displacement of water

**(b)What property of Oxygen make it to be collected using the method above** -Slightly soluble in water

**4.What is the purpose of manganese(IV)oxide?**

Manganese(IV)oxide is **catalyst**.

A catalyst is a substance that speeds up the rate of a chemical reaction but remain chemically unchanged at the end of the reaction.

Hydrogen peroxide decomposes slowly to form water and Oxygen gas.

A little Manganese(IV)oxide speeds up the rate of decomposition by **reducing** the time taken for a given volume of Oxygen to be produced.

**5.Write the equation for the reaction.**

Hydrogen peroxide -> Water + Oxygen

2H2O2 (aq) -> 2H2O (l) + O2 (g)

**6. Lower a glowing splint slowly into a gas jar containing Oxygen gas. State what is observed.**

The glowing splint relights/rekindles

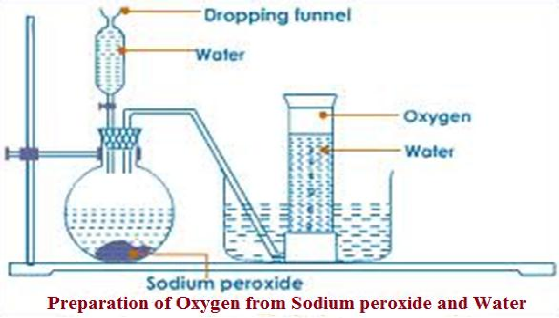
Oxygen relights/rekindles a glowing splint. This is the confirmatory test for the presence of Oxygen gas

Method 1: **Using Sodium peroxide**

Half fill a trough/basin with tap water. Add four drops of phenolphthalein indicator.

Place a bee hive shelf/stand into the water.

Completely fill a gas jar with water and invert in onto the bee hive shelf/stand. Clamp a round bottomed flask and set up the apparatus as below.



Collect several gas jars of Oxygen covering each sample.

Sample observation questions

1.**What is observed when water is added**

**(i)into the flask containing sodium peroxide**

Rapid effervescence/bubbling/fizzing

**(ii)phenolphththalein**

Remains colourless /Phenolphthalein indicator is colourless in neutral solution

**2. Describe the colour and smell of the gas**  Colourless and odourless.

**3.(a)Name the method of gas collection used.**

**-**Over water.Oxygen is slightly soluble in water.

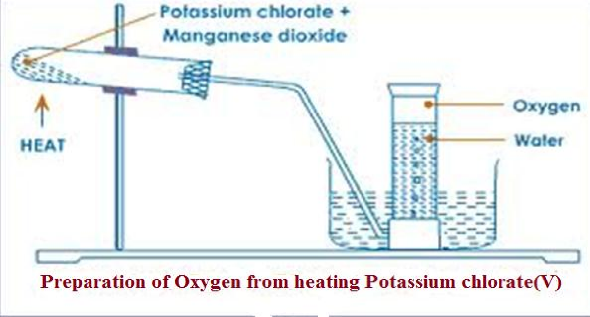
**4. Test the gas by lowering a glowing splint slowly into a gas jar containingthe prepared sample.**

The glowing splint relights/rekindles. This confirms the presence of Oxygen gas

**5.Write the equation for the reaction.**

Sodium peroxide + Water -> Sodium hydroxide + Oxygen

2Na2O2 (aq) + 2H2O (l) -> 4NaOH(aq) + O2 (g)



**1. Test the gas by lowering a glowing splint slowly into a gas jar containing the prepared sample.**

The glowing splint relights/rekindles.

This confirms the presence of Oxygen gas

**2.Write the equation for the reaction.**

Potassium Chlorate(V) -> Potassium Chloride + Oxygen

2KClO3 (aq) -> 2KCl(aq) + 3O2 (g)

**3.What is the purpose of manganese(IV)oxide?**

Manganese(IV)oxide is **catalyst**.

A catalyst is a substance that speeds up the rate of a chemical reaction but remain chemically unchanged at the end of the reaction.

Potassium Chlorate(V) decomposes slowly to form potassium chloride and Oxygen gas.

A little Manganese(IV)oxide speeds up the rate of decomposition by **reducing** the time taken for a given volume of Oxygen to be produced.

**(c)Uses of Oxygen**

1. Oxygen is put in cylinders for use where natural supply is not sufficiently enough. This is mainly in:

(i)Mountain climbing/Mountaineering-at high altitudes, the concentration of air/oxygen is low. Mountain climbers must therefore carry their own supply of oxygen for breathing.

* 1. Deep sea diving-Deep sea divers carry their own supply of Oxygen.
  2. Saving life in hospitals for patients with breathing problems and during anaethesia.

1. A mixture of oxygen and some other gases produces a flame that is very hot. (i) **Oxy-acetyline/ethyne** flame is produced when Ethyne/acetylene gas is burnt in pure oxygen. The flame has a temperature of about 3000oC.It is used for **welding /cutting** **metals**.

(ii)**Oxy-hydrogen** flame is produced when Hydrogen is burn in pure oxygen. The flame has a temperature of about 2000oC.It is used also for **welding /cutting** **metals**.

1. **Oxy-hydrogen** mixture is used as rocket fuel
2. A mixture of charcoal , petrol and liquid Oxygen is an explosive.

**(d) Chemical properties of Oxygen /combustion.**

Oxygen is a very reactive non metal. Many elements react with oxygen through burning to form a group of compounds called **Oxides**.

Burning/combustion is the reaction of Oxygen with an element/substances.

Reaction in which a substance is added oxygen is called **Oxidation reaction**.

Burning/combustion is an example of an oxidation reaction.

Most **non metals** burns in Oxygen/air to form an Oxide which in solution / dissolved in water is **acidic** in nature. They turn blue litmus red.e.g. Carbon(IV)oxide/CO2 , Nitrogen(IV)oxide/ NO2 , Sulphur(IV)oxide/ SO2

**Some non metals** burns in Oxygen/air to form an Oxide which in solution / dissolved in water is **neutral** in nature. They **don’t** turn blue or red litmus. e.g.

Carbon(II)oxide/CO, Water/ H2O.

All **metals** burns in Oxygen/air to form an Oxide which in solution/dissolved in water is **basic/alkaline** in nature. They turn red litmus blue.e.g. Magnesium oxide/MgO, Sodium Oxide/ Na2O ,Copper(II)oxide/CuO Elements/substances burn **faster** in pure Oxygen than in air.

Air contains the inactive part of air that **slows** the rate of burning of substances/elements.

**(i)Reaction of metals with Oxygen/air**

The following experiments show the reaction of metals with Oxygen and air.

**I. Burning Magnesium**

Procedure

(a)Cut a 2cm length piece of magnesium ribbon. Using a pair of tongs introduce it to a Bunsen flame. Remove it when it catches fire. Observe.

Place the products in a beaker containing about 5cm3 of water. Test the solution/mixture using litmus papers

(b)Cut another 2cm length piece of magnesium ribbon. Using a pair of tongs introduce it to a Bunsen flame. When it catches fire, lower it slowly into a gas jar containing Oxygen.

Place about 5cm3 of water into the gas jar. Test the solution/mixture using litmus papers. Test the solution/mixture using litmus papers

Observations

(a)In air

Magnesium burns with a bright blindening flame in air forming white solid/ash /powder. Effervescence/bubbles/ fizzing Pungent smell of urine. Blue litmus paper remains blue. Red litmus paper turns blue

(b) In pure Oxygen

Magnesium burns **faster** with a very bright blindening flame pure oxygen forming white solid/ash /powder. No effervescence/bubbles/ fizzing. No pungent smell of urine. Blue litmus paper remains blue. Red litmus paper turns blue

Explanation

Magnesium burns in air producing enough heat energy to react with both Oxygen and Nitrogen to form **Magnesium Oxide** and **Magnesium nitride.** Both Magnesium Oxide and Magnesium nitride are white solid/ash /powder.

Chemical equations

Magnesium + Oxygen -> Magnesium Oxide 2Mg(s) + O2(g) -> 2MgO(s)

Magnesium + Nitrogen -> Magnesium **Nitr**ide

3Mg(s) + N2(g) -> Mg3N2 (s)

Magnesium Oxide dissolves in water to form a basic/alkaline solution of

Magnesium hydroxide

Chemical equations

Magnesium Oxide + Water -> Magnesium hydroxide

2Mg(s) + O2(l) -> 2MgO(s)

Magnesium Nitride dissolves in water to form a basic/alkaline solution of Magnesium hydroxide and producing **Ammonia gas**. Ammonia is also an alkaline/basic gas that has a pungent smell of urine.

Chemical equations

Magnesium Nitride + Water -> Magnesium hydroxide + Ammonia gas

Mg3N2 (s) + 6H2O (l) -> 3Mg (OH)2 (aq) + 2NH3(g)

**II. Burning Sodium**

Procedure

(a)**Carefully** cut a very small piece of sodium . Using a deflagrating spoon introduce it to a Bunsen flame. Remove it when it catches fire. Observe. Place the products in a beaker containing about 20cm3 of water. Test the solution/mixture using litmus papers

(b) **Carefully** cut another very small piece of sodium. Using a deflagrating spoon introduce it to a Bunsen flame. When it catches fire, lower it slowly into a gas jar containing Oxygen.

Place about 20 cm3 of water into the gas jar. Test the solution/mixture using litmus papers. Test the solution/mixture using litmus papers

Observations

(a)In air

Sodium burns with a **yellow** flame in air forming a **black** solid. Blue litmus paper remains blue. Red litmus paper turns blue

(b) In pure Oxygen

Sodium burns **faster** with a golden **yellow** flame in pure oxygen forming a **yellow** solid. Effervescence/bubbles/ fizzing. Gas produced relights glowing splint.Blue litmus paper remains blue. Red litmus paper turns blue.

Explanation

(a)Sodium burns in air forming black **Sodium Oxide**

Chemical equations

Sodium + Oxygen/air -> Sodium Oxide

4Na(s) + O2(g) -> 2Na2O(s)

Sodium Oxide dissolves in water to form a basic/alkaline solution of Sodium hydroxide Chemical equations

Sodium Oxide + Water -> Sodium hydroxide

Na2O(s) + H2O (l) -> 2NaOH(aq)

(b)Sodium burns in pure oxygen forming yellow **Sodium peroxide**

Chemical equations

Sodium + Oxygen -> Sodium peroxide 2Na(s) + O2(g) -> Na2O2 (s)

Sodium peroxide dissolves in water to form a basic/alkaline solution of Sodium hydroxide. Oxygen is produced.

Chemical equations

Sodium Oxide + Water -> Sodium hydroxide + Oxygen

2Na2O2 (s) + 2H2O (l) -> 4NaOH(**aq**) + O2 (l)

**III. Burning Calcium**

Procedure

(a)Using a pair of tongs hold the piece of calcium on a Bunsen flame. Observe.

Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of calcium on a Bunsen flame.

Quickly lower it into a gas jar containing Oxygen gas .Observe.

Place about 2cm3 of water. Swirl.

Test the solution/mixture using litmus papers

Observations

(a)In air

Calcium burns with difficulty producing a faint **red** flame in air forming a **white** solid. Blue litmus paper remains blue. Red litmus paper turns blue (b) In pure Oxygen

Calcium burns with difficulty producing a less faint **red** flame Oxygen forming a **white** solid. Blue litmus paper remains blue. Red litmus paper turns blue

Explanation

(a)Calcium burns in air forming white **calcium Oxide.** Calcium Oxide coat/cover the calcium preventing further burning.

Chemical equations

Calcium + Oxygen/air -> calcium Oxide 2Ca(s) + O2(g) -> 2CaO(s)

Small amount of Calcium Oxide dissolves in water to form a basic/alkaline solution of Calcium hydroxide. The common name of Calcium hydroxide is **lime water.**

Chemical equations

Calcium Oxide + Water -> Calcium hydroxide CaO(s) + H2O (l) -> Ca(OH) 2 (aq)

**IV. Burning Iron**

Procedure

(a)Using a pair of tongs hold the piece of Iron wool/steel wire on a Bunsen flame.

Observe.

Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of Iron wool/steel wire on a Bunsen flame.

Quickly lower it into a gas jar containing Oxygen gas .Observe.

Place about 2cm3 of water. Swirl. Test the solution/mixture using litmus papers Observations

(a)In air

Iron wool/steel wire burns producing a **Orange** flame in air forming a **brown** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue

(b) In pure Oxygen

Iron wool/steel wire burns producing a golden **Orange** flame in Oxygen forming a **Brown** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue Explanation

(a)Iron burns in air forming brown **Iron(III) Oxide**

Chemical equations

Iron + Oxygen/air -> Iron(III) Oxide

4Fe(s) + 3O2(g) -> 2Fe2O3(s)

Very small amount of Iron(III)Oxide dissolves in water to form a weakly basic/alkaline **brown** solution of Iron(III) hydroxide.

Chemical equations

Calcium Oxide + Water -> Iron(III) hydroxide Fe2O3(s) + 3H2O (l) -> 2Fe(OH) 3 (**s**)

**V. Burning Copper**

Procedure

(a)Using a pair of tongs hold the piece of copper turnings/shavings on a Bunsen flame. Observe.

Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of Copper turnings/shavings on a Bunsen flame. Quickly lower it into a gas jar containing Oxygen gas .Observe.

Place about 2cm3 of water. Swirl. Test the solution/mixture using litmus papers Observations

(a)In air

Copper turnings/shavings burns with difficulty producing a **green** flame in air forming a **black** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue

(b) In pure Oxygen

Copper turnings/shavings burns less difficulty producing a **green** flame in Oxygen forming a **Brown** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue

Explanation

(a)Copper burns in air forming black **Copper(II) Oxide**

Chemical equations

Copper + Oxygen/air -> Copper(II) Oxide 2 Cu(s) + O2(g) -> 2CuO(s)

Very small amount of Copper(II)Oxide dissolves in water to form a weakly basic/alkaline **blue** solution of Copper(II) hydroxide.

Chemical equations

Copper(II) Oxide + Water -> Copper(II) hydroxide

CuO(s) + H2O (l) -> Cu(OH) 2 (**s**)

**(i)Reaction of non metals with Oxygen/air**

The following experiments show the reaction of non metals with Oxygen and air.

**I. Burning Carbon**

Procedure

(a)Using a pair of tongs hold a dry piece of charcoal on a Bunsen flame.

Observe.

Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of dry charcoal on a Bunsen flame. Quickly lower it into a gas jar containing Oxygen gas .Observe.

Place about 2cm3 of water. Swirl. Test the solution/mixture using litmus papers Observations

-Carbon **chars** then burns with a **blue** flame -Colourless and odourless gas produced

-Solution formed turn blue litmus paper faint red.

Red litmus paper remains red.

Explanation

Carbon burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas.

Carbon burns in limited supply of air with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas.

Carbon (IV) oxide gas dissolve in water to form weak acidic solution of Carbonic (IV)acid.

Chemical Equation

Carbon + Oxygen -> Carbon(IV)oxide

(excess air/oxygen)

C(s) + O2(g) -> CO2(g) (in excess air)

Carbon + Oxygen -> Carbon(II)oxide

(limited air/oxygen)

2C(s) + O2(g) -> 2CO(g) (in limited air)

Carbon(IV)oxide + Water -> Carbonic(IV)acid

CO2(g) + H2O (l) -> H2CO3 (aq) (very weak acid)

**II. Burning Sulphur**

Procedure

(a)Using a deflagrating spoon place sulphur powder on a Bunsen flame.

Observe.

Place the products in a beaker containing about 3cm3 of water. Test the solution/mixture using litmus papers

(b) Using a deflagrating spoon place sulphur powder on a Bunsen flame. Slowly lower it into a gas jar containing Oxygen gas. Observe.

Place about 5cm3 of water. Swirl. Test the solution/mixture using litmus papers. Observations

-Sulphur burns with a **blue** flame

-Gas produced that has pungent choking smell -Solution formed turn blue litmus paper faint red.

Red litmus paper remains red.

Explanation

Sulphur burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Sulphur (IV) oxide gas.

Sulphur (IV) oxide gas dissolve in water to form weak acidic solution of Sulphuric (IV)acid.

Chemical Equation

Sulphur + Oxygen -> Sulphur(IV)oxide

S(s) + O2(g) -> SO2(g) (in excess air)

Sulphur(IV)oxide + Water -> Sulphuric(IV)acid

SO2(g) + H2O (l) -> H2SO3 (aq) (very weak acid)

**III. Burning Phosphorus**

Procedure

(a)Remove a small piece of phosphorus from water and using a deflagrating spoon (with a lid cover)place it on a Bunsen flame.

Observe.

Carefully put the burning phosphorus to cover gas jar containing about 3cm3 of water. Test the solution/mixture using litmus papers

(b) Remove another small piece of phosphorus from water and using a deflagrating spoon (with a lid cover) place it on a Bunsen flame.

Slowly lower it into a gas jar containing Oxygen gas with about 5 cm3 of water.

Observe.

Swirl. Test the solution/mixture using litmus papers.

Observations

-Phosphorus catches fire before heating on Bunsen flame

-Dense white fumes of a gas produced that has pungent choking **poisonous** smell

-Solution formed turn blue litmus paper faint red.

Red litmus paper remains red.

Explanation

Phosphorus is stored in water.On exposure to air it instantaneously fumes then catch fire to burn in air and faster in Oxygen with a **yellow** flame producing dense white acidic fumes of Phosphorus(V) oxide gas.

Phosphoric(V) oxide gas dissolve in water to form weak acidic solution of Phosphoric (V)acid.

Chemical Equation

Phosphorus + Oxygen -> Phosphorous(V)oxide

4P(s) + 5O2(g) -> 2P2O5(s)

Phosphorous(V)oxide + Water -> Phosphoric(V)acid

P2O5(s) + 3H2O (l) -> 2H3PO4 (aq) (very weak acid)

**(e) Reactivity series/competition for combined Oxygen.**

The reactivity series is a list of elements/metals according to their affinity for oxygen.

Some metals have higher affinity for Oxygen than others.

A metal/element with higher affinity for oxygen is placed higher/on top of the one less affinity.

Most reactive

**The complete reactivity series of metals/elements**

|  |  |
| --- | --- |
| **Element/Metal** | **Symbol** |
| Potassium | K |
| Sodium | Na |

**Least reactive**

|  |  |
| --- | --- |
| Calcium | Ca |
| Magnesium | Mg |
| Aluminium | Al |
| **Carbon** | **C** |
| **Zinc** | **Zn** |
| **Iron** | **Fe** |
| **Tin** | **Sn** |
| **Lead** | **Pb** |
| **Hydrogen** | **H** |
| **Copper** | **Cu** |
| **Mercury** | **Hg** |
| **Silver** | **Ag** |
| **Gold** | **Au** |
| **Platinum** | **Pt** |

Metals compete for combined Oxygen. A metal/element with higher affinity for oxygen removes Oxygen from a metal lower in the reactivity series/less affinity for Oxygen.

When a metal/element gains/acquire Oxygen, the **process** is called **Oxidation**. When a metal/element donate/lose Oxygen, the **process** is called **Reduction**.

An element/metal/compound that undergo Oxidation is called **Reducing agent**.

An element/metal/compound that undergo Reduction is called **Oxidizing agent**.

A reaction in which **both** Oxidation and Reduction take place is called a **Redox** reaction.

Redox reaction between Magnesium and copper(II)Oxide

Procedure

Place about 2g of copper (II)oxide in a crucible with a lid. Place another 2g of Magnesium powder into the crucible. Mix thoroughly.

Cover the crucible with lid. Heat strongly for five minutes.

Allow the mixture to cool. Open the lid. Observe.

Observation

Colour change from black to brown. White solid power formed.

Explanation

Magnesium is higher in the reactivity series than Copper. It has therefore higher affinity for Oxygen than copper.

When a mixture of copper(II)oxide and Magnesium is heated, Magnesium reduces copper(II)oxide to brown copper metal and itself oxidized to Magnesium oxide. Magnesium is the reducing agent because it undergoes oxidation process. Copper(II)oxide is the oxidizing agent because it undergo **redox** reduction process.

The mixture should be cooled before opening the lid to prevent **hot** brown copper from being **reoxidized** back to black copper(II)oxide. The reaction of Magnesium and Copper(II)oxide is a reaction

Chemical equation

Red

uction process

Ox

idation process

1. Copper (II)oxide + Magnesium -> Magnesium oxide + Copper

(black) (white ash/solid) (brown)

**CuO**(s) + Mg(s) -> Mg**O**(s) + **Cu**(s) (Oxidizing Agent) (Reducing Agent)

1. Zinc (II)oxide + Magnesium -> Magnesium oxide + Zinc

(yellow when hot) (white ash/solid) (grey)

**ZnO**(s) + Mg(s) -> Mg**O**(s) + **Zn**(s) (Oxidizing agent) (Reducing agent)

1. Zinc (II)oxide + Carbon -> Carbon(IV) oxide gas + Zinc

(yellow when hot) (colourless gas) (grey)

**ZnO**(s) + C(s) -> C**O2**(g) + **Zn**(s) (Oxidizing agent) (Reducing agent)

The reactivity series is used during extraction of metals from their ore.An ore is a rock containing mineral element which can be extracted for commercial purposes. Most metallic ores occur naturally as:

(i)**ox**ides combined with Oxygen

(ii)**sulph**ides combined with Sulphur

(iii)**carbon**ates combined with carbon and Oxygen.

Metallic ores that naturally occur as metallic sulphides are first **roasted** in air to form the corresponding oxide. Sulphur(IV)oxide gas is produced. e.g.

Copper(I) sulphide + Oxygen -> Copper(I)Oxide + Sulphur(IV)oxide

Cu2S(s) + O2(g) -> 2Cu(s) + SO2(g)

Zinc(II) sulphide + Oxygen -> Zinc(II)Oxide + Sulphur(IV)oxide

ZnS(s) + O2(g) -> Zn(s) + SO2(g)

Lead(II) sulphide + Oxygen -> Lead(II)Oxide + Sulphur(IV)oxide

PbS(s) + O2(g) -> Pb(s) + SO2(g)

Iron(II) sulphide + Oxygen -> Iron(II)Oxide + Sulphur(IV)oxide FeS(s) + O2(g) -> Fe(s) + SO2(g)

Metallic ores that naturally occur as metallic carbonates are first **heated** in air. They **decompose**/split to form the corresponding oxide and produce Carbon (IV) oxide gas. e.g.

Copper (II)carbonate -> Copper(II)oxide + Carbon(IV)oxide

CuCO3(s) -> CuO(s) + CO2(g)

Zinc (II)carbonate -> Zinc(II)oxide + Carbon(IV)oxide ZnCO3(s) -> ZnO(s) + CO2(g)

Lead (II)carbonate -> Lead(II)oxide + Carbon(IV)oxide PbCO3(s) -> PbO(s) + CO2(g)

Iron(II)carbonate -> Iron(II)oxide + Carbon(IV)oxide

FeCO3(s) -> FeO(s) + CO2(g) Metallic ores

**WATER AND HYDROGEN**

A.WATER

Pure water is a **colourless**, **odourless**, **tasteless** ,**neutral** liquid. Pure water does not exist in nature but naturally in varying degree of purity. The main sources of water include rain, springs, borehole, lakes, seas and oceans: Water is generally **used** for the following purposes:

(i)drinking by animals and plants.

(ii)washing clothes.

(iii)bleaching and dyeing.

(iv) generating hydroelectric power.

(v)cooling industrial processes.

Water dissolves many substances/solutes.

It is therefore called **universal solvent**.

It contains about 35% **dissolved** Oxygen which support aquatic fauna and flora. Water naturally exist in three phases/states **solid** ice,**liquid** water and **gaseous** water vapour.

The three states of water are naturally **interconvertible**.

The natural interconvertion of the three phases/states of water forms the water cycle.

Precipitation

condensation

CLOUDS (Water in

**solid**

state)

RAIN

SPRING, RIVERS,WELLS.

OCEAN,LAKES,SEAS(water as

**liquid**

)

Evaporation

(

Water in

**gaseous**

state)

**Liquid** water in land, lakes , seas and oceans use the solar/sun **energy** to **evaporate/vapourize** to form water vapour/**gas**. Solar/sun energy is also used during transpiration by plants and respiration by animals.

During evaporation, the water vapour rises up the earths surface. Temperatures decrease with height above the earth surface increase. Water vapour therefore cools as it rises up. At a height where it is cold enough to below 373Kelvin/100oC Water vapour looses enough energy to form tiny droplets of liquid. The process by which a gas/water vapour changes to a liquid is called **condensation/liquidification**.

On further cooling, the liquid looses more energy to form **ice/solid.** The process by which a liquid/water changes to a ice/solid is called **freezing/solidification**. Minute/tiny ice/solid particles float in the atmosphere and coalesce/join together to form clouds. When the clouds become too heavy they fall to the earths surface as rain/snow as the temperature increase with the fall.

**Interconversion of the three phases/states water**

Liquidification/

/boiling/vapourization condensation

|  |  |
| --- | --- |
| Gas/water vapour | |
|  | Evaporation |
| Liquid/Water | |
|  | Melting liquidification |
| Solid/Ice | |

Freezing

Solidification

Pure water has :

1. fixed/constant/sharp freezing point/melting point of **273K/0oC**
2. fixed/constant/sharp boiling point of **373K/100oC** at **sea level**/1 atmosphere pressure
3. fixed density of **1gcm-3**

This is the **criteria** of identifying pure/purity of water.

Whether a substance is water can be determined by using the following methods:

**a)To test for presence of water using anhydrous copper(II)suphate(VI)**

**Procedure.**

Put about 2g of anhydrous copper(II)sulphate(VI)crystals into a clean test tube. Add three drops of tap water. Repeat the procedure using distilled water.

**Observation.**

Colour changes from white to blue **Explanation.**

Anhydrous copper(II)sulphate(VI)is white. On adding water ,**anhydrous** copper(II)sulphate(VI) gains/reacts with water to form **hydrated** copper(II) sulphate(VI).

Hydrated copper(II) sulphate(VI) is **blue**.Hydrated copper(II) sulphate(VI) contain water of crystallization.

The change of white **anhydrous** copper(II)sulphate(VI) to **blue** hydrated copper(II) sulphate(VI) is a confirmatory test for the **presence** of water **Chemical equation.**

Anhydrous Hydrated

copper(II)sulphate(VI) + Water -> copper (II)sulphate(VI)

(white) (**blue**)

CuSO4(s) + 5H2O(l) -> **CuSO4.5H2O(s)**

**b)To test for presence of water using anhydrous cobalt(II)chloride**

**Procedure.**

Put about 5cm3 of water into a clean test tube.

Dip a dry anhydrouscobalt(II)chloride **paper** into the test tube. Repeat the procedure using distilled water.

**Observation.**

Colour changes from blue to **pink** **Explanation.**

Anhydrous cobalt(II)chloride is **blue**. On adding water, **anhydrous** cobalt(II)chloride gains/reacts with water to form **hydrated** cobalt(II) chloride.

Hydrated cobalt(II)chloride is **pink**.

Hydrated cobalt (II)chloride contain water of crystallization.

The change of blue **anhydrous** cobalt(II)chloride to **pink** hydrated cobalt(II)chloride is a confirmatory test for the **presence** of water **Chemical equation.**

Anhydrous Hydrated

cobalt(II)chloride + Water -> cobalt (II)chloride

(**Blue**) (**pink**)

**CoCl2** (s) + 5H2O(l) -> **CoCl2.5H2O(s)**

**Burning a candle in air**

Most organic substances/fuels burn in air to produce water. Carbon(IV)oxide gas is also produced if the air is sufficient/excess.

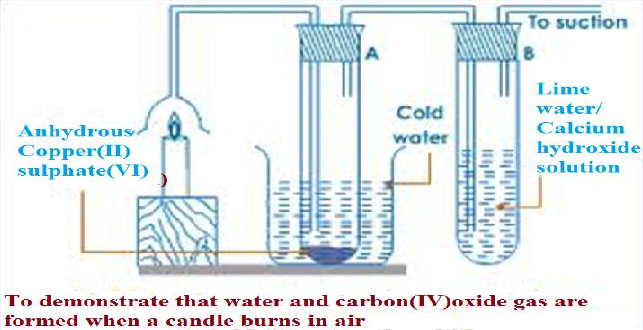
Procedure

Put about 2g of anhydrous copper(II)sulphate(VI)crystals in a boiling tube.

Put about 5cm3 of lime water in a boiling tube.

Light a small candle stick. Place it below an inverted thistle/filter funnel

Collect the products of the burning candle by setting the apparatus as below Set up of apparatus



Observation

The sunction pump pulls the products of burning into the inverted funnel. Colour of anhydrous copper(II) sulphate(VI)changes from white to blue. A **white precipitate** is formed in the lime water/calcium hydroxide.

Explanation

When a candle burn it forms a water and carbon(IV)oxide.

Water turns anhydrous copper(II) sulphate(VI)changes from white to blue . Carbon(IV)oxide gasforms **white precipitate** when bubbled in lime water/calcium hydroxide.

Since:

(i)hydrogen in the wax burn to form water

Hydrogen + Oxygen -> Water

(from candle) (from the air)

2H2(g) + O2(g) -> 2H2O (g/l)

(ii) carbon in the wax burn to form carbon(IV)oxide

Hydrogen + Oxygen -> Water

(from candle) (from the air)

C(s) + O2(g) -> CO2 (g)

The candle before burning therefore contained only **Carbon** and **Hydrogen** **only.**

A compound made up of **hydro**gen and carbon is called **Hydro**carbon.

A candle is a hydrocarbon.

Other hydrocarbons include: Petrol, diesel, Kerosene, and Laboratory gas. Hydrocarbons burn in air to form water and carbon(IV)oxide gas.

Hydrocarbons + Oxygen -> Water + Oxygen

**Water pollution**

Water pollution take place when undesirable substances are added into the water.

Sources of water pollution include:

(i)Industrial chemicals being disposed into water bodies like rivers, lakes and oceans.

(ii)Dicharging untreated /raw sewage into water bodies.

(iii)Leaching of insecticides/herbicides form agricultural activities into water bodies.

(iv)Discharging non-biodegradable detergents after domestic and industrial use into water bodies.

(v)Petroleum oil spilling by ships and oil refineries

(vi)Toxic/poisonous gases from industries dissolving in rain .

(vii) Acidic gases from industries dissolving in rain to form “acid rain” (viii)Discharging hot water into water bodies.This reduces the quantity of dissolved Oxygen in the water killing the aquatic fauna and flora.

Water pollution can be reduced by:

(i)reducing the use of agricultural fertilizers and chemicals in agricultural activities.

(ii)use of biological control method instead of insecticides and herbicides (iii)using biodegradable detergents

**Reaction of metals with water**

Some metals react with water while others do not. The reaction of metals with water depend on the reativity series. The higher the metal in the reactivity series the more reactive the metal with water .The following experiments shows the reaction of metals with cold water and water vapour/steam.

(a)Reaction of sodium/ potassium with cold water:

Procedure

Put about 500cm3 of water in a beaker. Add three drops of phenolphthalein indicator/litmus solution/universal indicator solution/methyl orange indicator into the water.

Cut a **very small** piece of sodium .Using a pair of forceps, put the metal into the water.

Observation

Sodium melts to a silvery ball that floats and darts on the surface decreasing in size.Effervescence/fizzing/ bubbles of colourless gas produced.

Colour of phenolphthalein turns **pink**

Colour of litmus solution turns **blue**

Colour of methy orange solution turns **Orange**

Colour of universal indicator solution turns **blue**

Explanation

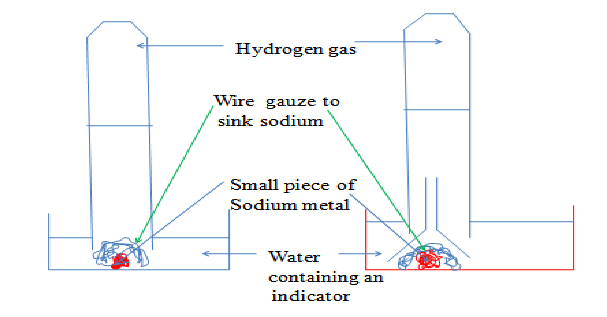
Sodium is less dense than water. Sodium floats on water and vigorously react to form an **alkaline** solution of sodium hydroxide and producing hydrogen gas. Sodium is thus stored in paraffin to prevent **contact** with water.

Chemical equation

Sodium + Water -> Sodium hydroxide + Hydrogen gas

2Na(s) + 2H2O(l) -> 2NaOH(aq) + H2(g)

To collect hydrogen gas , Sodium metal is forced to **sink** to the bottom of the trough/beaker by wrapping it in wire gauze/mesh.



Potassium is more reactive than Sodium. On contact with water it **explodes**/burst into flames. An alkaline solution of potassium hydroxide is formed and hydrogen gas

Chemical equation

Potassium + Water -> Potassium hydroxide + Hydrogen gas

2K(s) + 2H2O(l) -> 2KOH(aq) + H2(g)

**Caution: Reaction of Potassium with water is very risky to try in a school laboratory.**

(b)Reaction of Lithium/ Calcium with cold water:

Procedure

Put about 200cm3 of water in a beaker. Add three drops of phenolphthalein indicator/litmus solution/universal indicator solution/methyl orange indicator into the water.

Cut a small piece of Lithium .Using a pair of forceps, put the metal into the water. Repeat with a piece Calcium metal

Observation

Lithium sinksto the bottom of the water.Rapid effervescence/fizzing/ bubbles of colourless gas produced.

Colour of phenolphthalein turns **pink**

Colour of litmus solution turns **blue**

Colour of methy orange solution turns **Orange**

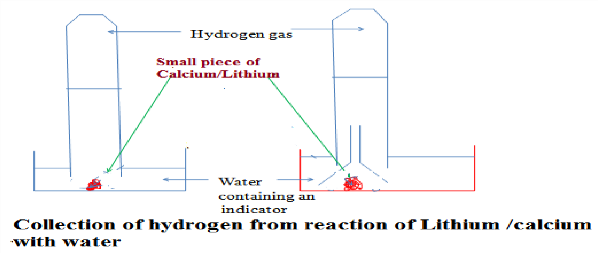
Colour of universal indicator solution turns **blue**

Explanation

Lithium and calcium are **denser** than water. Both sink in water and vigorously react to form an **alkaline** solution of Lithium hydroxide / calcium hydroxide and producing hydrogen gas. Lithium is more reactive than calcium. It is also stored in paraffin like Sodium to prevent **contact** with water.

Chemical equation

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Lithium | + | Water | -> | Lithium hydroxide + Hydrogen gas |
| 2Li(s) | + | 2H2O(l) | -> | 2LiOH(aq) + H2(g) |
| Calcium | + | Water | -> | Calcium hydroxide + Hydrogen gas |
| Ca(s) | + | 2H2O(l) | -> | Ca(OH)2(aq) + H2(g) |



(c) Reaction of Magnesium/Zinc/ Iron with Steam/water vapour:

Procedure method1

Place some wet sand or cotton/glass wool soaked in water at the bottom of an ignition/hard glass boiling tube.

Polish magnesium ribbon using sand paper.

Coil it at the centre of the ignition/hard glass boiling tube.

Set up the apparatus as below.

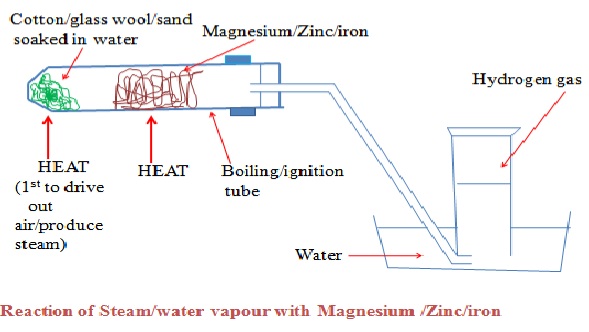
Heat the wet sand or cotton/glass wool soaked in water gently to:

(i)drive away air in the ignition/hard glass boiling tube.

(ii)generate steam

Heat the coiled ribbon strongly using another burner.Repeat the experiment using Zinc powder and fresh Iron filings.

Set up of apparatus



Observations

(i)With Magnesium ribbon:

The Magnesium glow with a bright flame (and continues to burn even if heating is stopped)

White solid /ash formed

White solid /ash formed dissolve in water to form a colourless solution

Colourless gas produced/collected that extinguish burning splint with “pop sound” (ii)With Zinc powder:

The Zinc powder turns red hot on strong heating

Yellow solid formed that turn white on cooling

White solid formed on cooling does not dissolve in water.

(iii)With Iron fillings:

The Iron fillings turns red hot on strong heating

Dark blue solid formed

Dark blue solid formed does not dissolve in water.

Procedure method 2

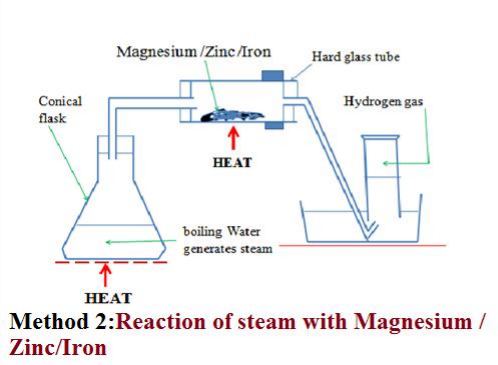
Put some water in a round bottomed flask Polish magnesium ribbon using sand paper. Coil it at the centre of a hard glass tube Set up the apparatus as below.

Heat water strongly to boil so as to:

(i)drive away air in the glass tube.

(ii)generate steam

Heat the coiled ribbon strongly using another burner. Repeat the experiment using Zinc powder and fresh Iron filings.



Observations

(i)With Magnesium ribbon:

The Magnesium glow with a bright flame (and continues to burn even if heating is stopped)

White solid /ash formed

White solid /ash formed dissolve in water to form a colourless solution

Colourless gas produced/collected that extinguish burning splint with “pop sound” (ii)With Zinc powder:

The Zinc powder turns red hot on strong heating

Yellow solid formed that turn white on cooling

White solid formed on cooling does not dissolve in water.

(iii)With Iron fillings:

The Iron fillings turns red hot on strong heating

Dark blue solid formed

Dark blue solid formed does not dissolve in water.

Explanations

(**a**)Hot magnesium burn vigorously in steam. The reaction is highly exothermic generating enough heat/energy to proceed without further heating. White Magnesium oxide solid/ash is left as residue.

Hydrogen gas is produced .It extinguishes a burning splint with a “pop sound”.

Chemical Equation

Magnesium + Steam -> Magnesium oxide + Hydrogen Mg(s) + H2O(**g**) -> MgO(s) + H2(g)

Magnesium oxide reacts /dissolves in water to form an alkaline solution of

Magnesium oxide

Chemical Equation

Magnesium oxide + Water -> Magnesium hydroxide MgO(s) + H2O(**l**) -> Mg(OH) 2 (**aq**)

(**b**)Hot Zinc react vigorously in steam forming yellow Zinc oxide solid/ash as residue which cools to white.

Hydrogen gas is produced .It extinguishes a burning splint with a “pop sound”.

Chemical Equation

Zinc + Steam -> Zinc oxide + Hydrogen Zn(s) + H2O(**g**) -> ZnO(s) + H2(g)

Zinc oxide does not dissolve in water.

(**c**)Hot Iron react with steam forming dark blue tri iron tetra oxide solid/ash as residue.

Hydrogen gas is produced .It extinguishes a burning splint with a “pop sound”.

Chemical Equation

Iron + Steam -> Tri iron tetra oxide + Hydrogen 2Fe(s) + 4H2O(**g**) -> Fe2O4(s) + 4H2(g)

Tri iron tetra oxide does not dissolve in water.

(**d**)Aluminium reacts with steam forming an **insoluble coat**/cover of **impervious** layer of aluminium oxide on the surface preventing further reaction.

(**e**) Lead, Copper, Mercury, Silver, Gold and Platinum **do not** react with either water or steam.

**HYDROGEN**

**Occurrence**

Hydrogen does not occur free in nature. It occurs as Water and in Petroleum.

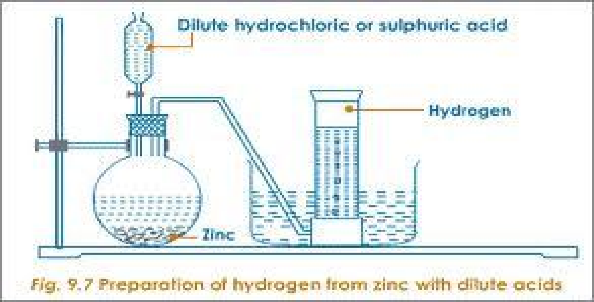
**School laboratory Preparation**

Procedure

Put Zinc granules in a round/flat/conical flask. Add dilute sulphuric(VI) /Hydrochloric acid.

Add about 3cm3 of copper(II)sulphate(VI) solution.

Collect the gas produced over water as in the set up below. Discard the first gas jar. Collect several gas jar.



Observation/Explanation

Zinc reacts with dilute sulphuric(VI)/hydrochloric acid to form a salt and produce hydrogen gas.

When the acid comes into contact with the metal, there is rapid effervescence/ bubbles /fizzing are produced and a colourless gas is produced that is collected:

(i) over water because it is insoluble in water

(ii)through downward displacement of air/upward delivery because it is less dense than air.

The first gas jar is impure. It contains air that was present in the apparatus.

Copper(II)sulphate(VI)solution act as catalyst.

Chemical equation

1. Zinc + Hydrochloric acid -> Zinc chloride + Hydrogen

Zn(s) + 2HCl(aq) -> ZnCl2(aq) + H2(g)

Ionic equation

Zn (s) + 2H+ (aq) -> Zn2+ (aq) + H2 (g)

Zinc + Sulphuric(VI)acid -> Zinc Sulphate(VI) + Hydrogen

|  |  |  |
| --- | --- | --- |
| Zn(s) + H2SO4(aq) Ionic equation | -> ZnSO4(aq) | + H2(g) |
| Zn (s) + 2H+ (aq) | -> Zn2+ (aq) + | H2 (g) |

1. Chemical equation

Magnesium + Hydrochloric acid -> Magnesium chloride + Hydrogen

Mg(s) + 2HCl(aq) -> MgCl2(aq) + H2(g)

Ionic equation

Mg (s) + 2H+ (aq) -> Mg2+ (aq) + H2 (g)

Magnesium + Sulphuric(VI)acid -> Magnesium Sulphate(VI) + Hydrogen

Mg(s) + H2SO4(aq) -> MgSO4(aq) + H2(g)

Ionic equation

Mg (s) + 2H+ (aq) -> Mg2+ (aq) + H2 (g)

1. Chemical equation

Iron + Hydrochloric acid -> Iron(II)chloride + Hydrogen

Fe(s) + 2HCl(aq) -> FeCl2(aq) + H2(g) Ionic equation

Fe (s) + 2H+ (aq) -> Fe2+ (aq) + H2 (g)

Iron + Sulphuric(VI)acid -> Iron(II) Sulphate(VI) + Hydrogen

|  |  |
| --- | --- |
| Fe(s) + H2SO4(aq) -> FeSO4(aq) Ionic equation | + H2(g) |
| Fe (s) + 2H+ (aq) -> Fe2+ (aq) | + H2 (g) |

**Note**

1.Hydrogen cannot be prepared from reaction of:

(i)Nitric(V)acid and a metal. Nitric(V)acid is a strong oxidizing agent. It **oxidizes** hydrogen gas to **water**.

(ii)dilute sulphuric(VI)acid with calcium/Barium/Lead because Calcium sulphate(VI),Barium sulphate(VI) and Lead(II)sulphate(VI) salts formed are insoluble. Once formed, they **cover/coat** the **unreacted** calcium/Barium/Lead **stopping** further reaction and producing very small amount/volume of hydrogen gas.

(iii)dilute acid with sodium/potassium. The reaction is **explosive.**

**Properties of Hydrogen gas**

**(a)Physical properties**

1. Hydrogen is a **neutral** ,**colourless** and **odourless** gas. When mixed with air it has a characteristic pungent choking smell
2. It is insoluble in water thus can be collected over water.
3. It is the lightest known gas. It can be transferred by inverting one gas jar over another.

**(b)Chemical properties.**

**(i)Burning**

1. Hydrogen does not support burning/combustion. When a burning splint is inserted into a gas jar containing Hydrogen, the flame is extinguished /put off.

1. Pure dry hydrogen burn with a blue quiet flame to form water. When a stream of pure dry hydrogen is ignited, it catches fire and continues to burn with a blue flame.

1. Impure (air mixed with) hydrogen burns with an explosion. Small amount/ volume of air **mixed** with hydrogen in a test tube produce a small explosion as a

“pop” sound. This is the confirmatory test for the presence of Hydrogen gas. A gas that burns with a “pop” sound is confirmed to be Hydrogen.

**(ii)Redox in terms of Hydrogen transfer**

Redox can also be defined in terms of Hydrogen transfer.

(i)Oxidation is removal of Hydrogen

(ii)Reduction is addition of Hydrogen

(iii)Redox is simultaneous addition and removal of Hydrogen

**Example**

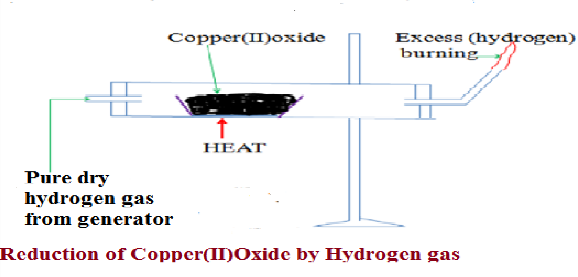
When a stream of dry hydrogen gas is passed through black copper (II) oxide, hydrogen gas gains the oxygen from copper(II)oxide.

Black copper (II) oxide is reduced to brown copper metal.

Black copper(II)oxide os thus the Oxidizing agent.

Hydrogen gas is oxidized to Water. Hydrogen is the Reducing agent.

Set up of apparatus



(**a**)Chemical equation

(i) In glass tube

Copper(II)Oxide + Hydrogen -> Copper + Hydrogen gas

(oxidizing agent) (reducing agent)

(**black**) (brown)

CuO (s) + H2(g) -> Cu(s) + H2O(l)

(ii)when excess Hydrogen is burning.

Oxygen + Hydrogen -> Water

O2(g) + 2H2(g) -> 2H2O(l)

(**b**)Chemical equation

(i) In glass tube

Lead(II)Oxide + Hydrogen -> Lead + Hydrogen gas

(oxidizing agent) (reducing agent)

(brown when hot/ (grey) yellow when cool)

PbO (s) + H2(g) -> Pb(s) + H2O(l)

(ii)when excess Hydrogen is burning.

Oxygen + Hydrogen -> Water

O2(g) + 2H2(g) -> 2H2O(l)

(**c**)Chemical equation

(i) In glass tube

Iron(III)Oxide + Hydrogen -> Iron + Hydrogen gas

(oxidizing agent) (reducing agent)

(Dark grey) (grey)

Fe2O3 (s) + 3H2(g) -> Fe(s) + 3H2O(l)

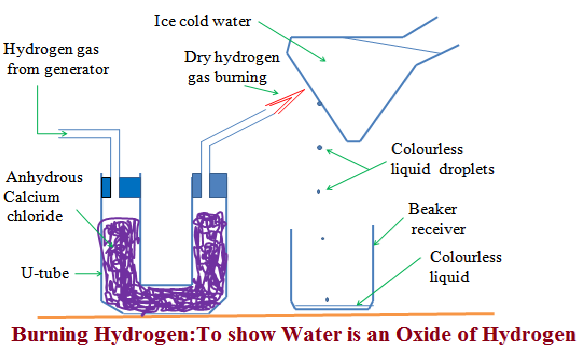
(ii)when excess Hydrogen is burning.

Oxygen + Hydrogen -> Water

O2(g) + 2H2(g) -> 2H2O(l)

**(iii) Water as an Oxide as Hydrogen**

Burning is a reaction of an element with Oxygen. The substance formed when an element burn in air is the oxide of the element. When hydrogen burns, it reacts/ combines with Oxygen to form the **oxide of Hydrogen**.Theoxide of Hydrogen is called water. Hydrogen is first dried because a mixture of Hydrogen and air explode. The gas is then ignited .The products condense on a cold surface/flask containing a freezing mixture. A freezing mixture is a mixture of water and ice.



The condensed products are collected in a receiver as a colourless liquid.

**Tests**

1. When about 1g of **white** anhydrous copper (II)sulphate(VI)is added to a sample of the liquid ,it turns to **blue**. This confirms the liquid formed is water.

1. When **blue** anhydrous cobalt (II)chloride paper is dipped in a sample of the liquid ,it turns to **pink**. This confirms the liquid formed is water.

(**c**)When the liquid is heated to boil, its **boiling point** is **100oC** at sea level/one atmosphere pressure. This confirms the liquid is **pure water.**

**Uses of Hydrogen gas**

1. **Hydrogenation/Hardening of unsaturated vegetable oils to saturated fats/margarine.**

When Hydrogen is passed through unsaturated compounds in presence of **Nickel** catalyst and about **150oC**, they become saturated. Most vegetable oil are unsaturated liquids at room temperature. They become saturated and hard through hydrogenation.

1. **In weather forecast balloons.**

Hydrogen is the lightest known gas**.** Meteorological data is collected for analysis by sending hydrogen filled weather balloons to the atmosphere. The data collected is then used to forecast weather conditions.

**3.In the Haber process for the manufacture of Ammonia**

Hydrogen is mixed with Nitrogen in presence of Iron catalyst to form Ammonia gas. Ammonia gas is a very important raw material for manufacture of agricultural fertilizers.

**4.In the manufacture of Hydrochloric acid.**

Limited volume/amount of Hydrogen is burnt in excess chlorine gas to form Hydrogen chloride gas. Hydrogen chloride gas is dissolved in water to form Hydrochloric acid. Hydrochloric acid is used in pickling/washing metal surfaces.

1. **As rocket fuel.**

Fixed proportions of Hydrogen and Oxygen when ignited explode violently producing a lot of energy/heat.This energy is used to power/propel a rocket to space.

1. **In oxy-hydrogen flame for welding.**

A cylinder containing Hydrogen when ignited in pure Oxygen from a second cylinder produces a flame that is very hot. It is used to cut metals and welding.

**Sample revision questions**

**1.A colourless liquid was added anhydrous copper(II)sulphate(VI) which turned blue.**

**(a)Why is it wrong to conclude the liquid was pure water?**

Anhydrous copper(II)sulphate(VI) test for presence of water. Purity of water is determined from freezing/melting/boiling point.

**(b)Write an equation for the reaction that take place with anhydrous copper(II)sulphate(VI)**

Anhydrous copper(II)sulphate(VI) + Water -> hydrated copper(II)sulphate(VI)

CuSO4(s) + 5H2O(l) -> CuSO4.5H2O(s)

**(c)(i)Which other compound would achieve the same results as** **anhydrous copper(II)sulphate(VI)**

Anhydrous cobalt (II)chloride/CoCl2.6H2O

**(ii)Write the equation for the reaction**

Anhydrous cobalt (II)chloride + Water -> hydrated cobalt (II)chloride CoCl2 (s) + 6H2O(l) -> CoCl2.6H2O (s)

|  |  |
| --- | --- |
| **(d)Complete the equation** |  |
| **(i) Sulphur(VI)oxide** | **+ Water -> Sulphuric(VI)acid** |
| **(ii) Sulphur(IV)oxide** | **+ Water -> Sulphuric(IV)acid** |
| **(iii) Carbon(IV)oxide** | **+ Water -> Carbonic(IV)acid** |
| **(iv) Nitrogen(IV)oxide** | **+ Water -> Nitric(V)acid** |
| **(v) Phosphorus(V)oxide** | **+ Water -> Phosphoric(V)acid** |
| **(vi) Sodium oxide** | **+ Water -> Sodium hydroxide** |
| **(vi) Sodium peroxide** | **+ Water -> Sodium hydroxide** |

**2. Metal B reacts with steam. Metal C reacts with cold water. Metal A does not react with water.**

**(a)Arrange the metals as they should appear in the reactivity series.**

B

C

A

1. **A product residue in D which was brown when hot but turned yellow on cooling during the reaction of metal B was formed. Gas E was also evolved.**

**Identify**

**(i)Metal B** Lead/Pb

**(ii)Residue D** Lead(II)oxide/PbO **(iii)Gas E** Hydrogen/H2

1. **A portion of product residue in D was added dilute nitric(V)acid. Another portion of product residue in D was added dilute sulphuric(VI)acid. State and explain the observations made.**

When added dilute nitric(V)acid, D dissolves to form a colourless solution.

Lead(II)Oxide + dilute nitric(V)acid -> Lead(II) nitrate(V) + Water

PbO (s) + 2HNO3(aq) -> Pb(NO3)2 (**aq**) + H2O(l)

When added dilute sulphuric(VI)acid, D does not dissolve. A white suspension/precipitate was formed. Lead(II)Oxide reacts with sulphuric(VI)acid to form insoluble Lead(II)sulphate(VI) that cover/coat unreacted Lead(II)Oxide, stopping further reaction.

Lead(II)Oxide + dilute sulphuric(VI)acid -> Lead(II) sulphate(VI) + Water

PbO (s) + H2SO4(aq) -> PbSO4 (**s**) + H2O(l)

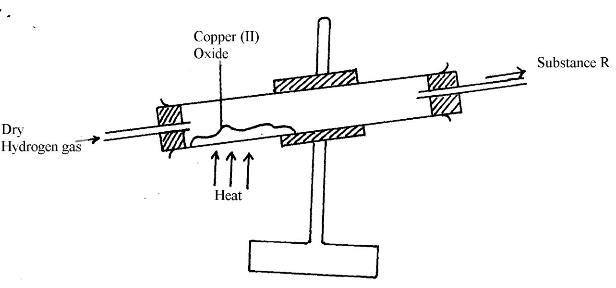
3. (a) Hydrogen can reduce copper(II)Oxide but not alluminium oxide. Explain

(b) When water reacts with potassium metal the hydrogen produced ignites explosively on the surface of water.

1. What causes this ignition?

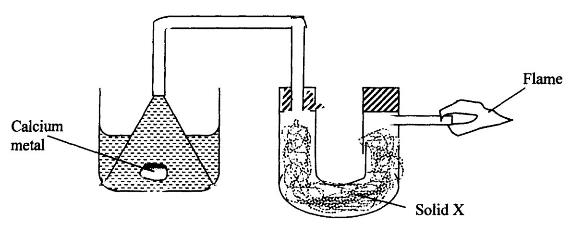
1. Write an equation to show how this ignition occurs

2. In an experiment, dry hydrogen gas was passed over hot copper (II) oxide in a combustion tube as shown in the diagram below:



* 1. Complete the diagram to show how the other product, substance **R** could be collected in the laboratory.

* 1. Describe how copper could be obtained from the mixture containing copper (II) oxide



3

.

The setup below was used to investigate the reaction between metals and

water.

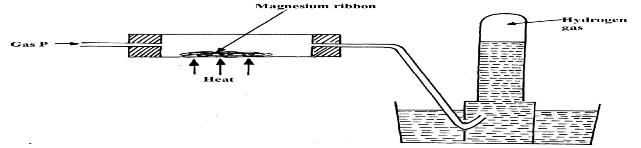
1. Identify solid **X** and state its purpose

Solid X .…………………………………………………………………..

Purpose …………………………………………………………………..

1. Write a chemical equation for the reaction that produces the flame.

1. Gas **P** was passed over heated magnesium ribbon and hydrogen gas was



collected as shown

in the diagram below:

(

i) Name gas

**P**

...............................................................................................................

* 1. Write an equation of the reaction that takes place in the combustion tube

* 1. State **one** precaution necessary at the end of this experiment

1. When hydrogen is burnt and the product cooled, the following results are obtained as shown in the diagram below:

Dry hydrogen

Liquid

**Y**



Burning

Ice cold water

Clamp

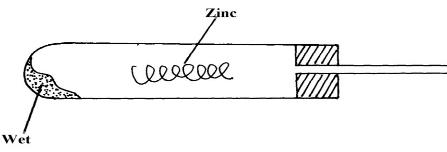


Clamp

* 1. Write the equation for the formation of liquid **Y**

* 1. Give a chemical test for liquid **Y**

1. Jane set-up the experiment as shown below to collect a gas. The wet

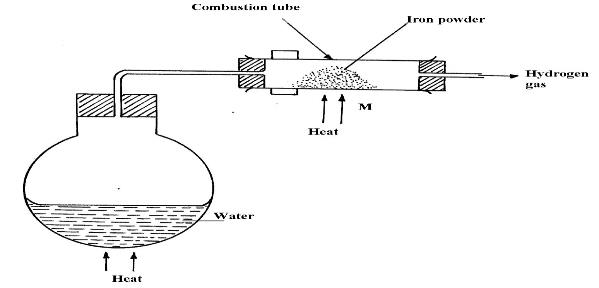


sand was heated before

heating Zinc granules

**Wet sand**

* 1. Complete the diagram for the laboratory preparation of the gas

* 1. Why was it necessary to heat wet sand before heating Zinc granules?

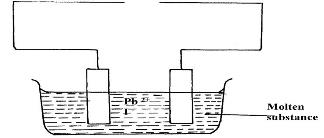
7.

**N**

* 1. Between **N** and **M** which part should be heated first? Explain

* 1. Write a chemical equation for the reaction occurring in the combustion tube.

1. The set-up below was used to investigate electrolysis of a certain molten compound;-



(a) Complete the circuit by drawing the cell in the gap left in the diagram (b) Write half-cell equation to show what happens at the cathode

(c) Using an arrow show the direction of electron flow in the diagram above

1. Hydrogen can be prepared by reacting zinc with dilute hydrochloric acid.
   * 1. Write an equation for the reaction.

* + 1. Name an appropriate drying agent for hydrogen gas.

* + 1. Explain why copper metal cannot be used to prepare hydrogen gas.

* + 1. Hydrogen burns in oxygen to form an oxide.
       1. Write an equation for the reaction.

* + - 1. State **two** precautions that must be taken before the combustion begins and at the end of the combustion.

* + 1. Give **two** uses of hydrogen gas.

* + 1. When zinc is heated to redness in a current of steam, hydrogen gas is

obtained. Write an equation for the reaction.

* + 1. Element **Q** reacts with dilute acids but not with cold water. Element **R**

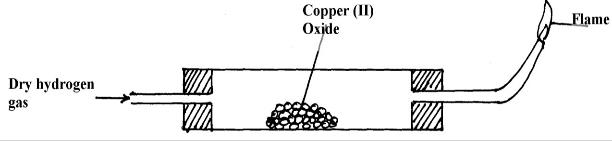
does not react with

dilute acids. Elements **S** displaces element **P** from its oxide. **P** reacts with

cold water. Arrange the four elements in order of their reactivity, starting with the most reactive.

* + 1. Explain how hydrogen is used in the manufacture of margarine.

1. a) The set-up below is used to investigate the properties of hydrogen.



* + 1. On the diagram, indicate what should be done for the reaction to occur

* + 1. Hydrogen gas is allowed to pass through the tube for some time before it is lit. Explain
    2. Write an equation for the reaction that occurs in the combustion tube

* + 1. When the reaction is complete, hydrogen gas is passed through the apparatus until they

cool down . Explain

* + 1. What property of hydrogen is being investigated?

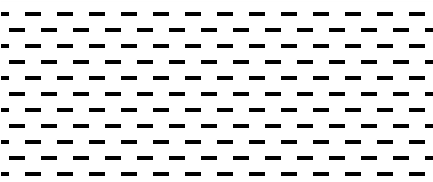
vi) What observation confirms the property stated in (**v)** above?

vii) Why is zinc oxide not used to investigate this property of hydrogen gas?

1. The set up below was used to collect gas **K,** produced by the reaction

between water and

calcium metal.



o

o

o

o

Water

Calcium metal

Gas

**K**

* + 1. Name gas **K**

……………………………………………………………..

* + 1. At the end of the experiment, the solution in the beaker was found to be a weak base. Explain why the solution is a weak base

**FORM TWO**

**A.ATOMIC STRUCTURE**

The atom is the smallest particle of an element that take part in a chemical reaction. The atom is made up of three subatomic particles:

(i)Protons

(ii)Electrons

(iii)Neutrons

(i)Protons

1.The proton is positively charged

2.Is found in the centre of an atom called nucleus

3.It has a relative mass 1

4.The number of protons in a atom of an element is its Atomic number

(ii)Electrons

1.The Electrons is negatively charged

2.Is found in fixed regions surrounding the centre of an atom called energy levels/orbitals.

3.It has a relative mass 1/1840

4.The number of protons and electrons in a atom of an element is always equal

(iii)Neutrons

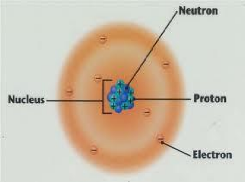
1.The Neutron is neither positively or negatively charged thus neutral.

2.Like protons it is found in the centre of an atom called nucleus

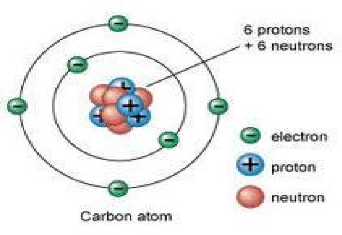
3.It has a relative mass 1

4.The number of protons and neutrons in a atom of an element is its Mass number

**Diagram showing the relative positions of protons ,electrons and neutrons in an atom of an element**



**Diagram showing the relative positions of protons, electrons and neutrons in an atom of Carbon**



The table below show atomic structure of the 1st twenty elements.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Symbol**  **Element** | **Protons** | **Electrons** | **Neutrons** | **Atomic number** | **Mass number** |
| Hydrogen **H** | **1** | **1** | **0** | **1** | **1** |
| Helium **He** | **2** | **2** | **2** | **2** | **4** |
| Lithium **Li** | **3** | **3** | **4** | **3** | **7** |
| Beryllium **Be** | **4** | **4** | **5** | **4** | **9** |
| Boron **B** | **5** | **5** | **6** | **5** | **11** |
| Carbon **C** | **6** | **6** | **6** | **6** | **12** |
| Nitrogen **N** | **7** | **7** | **7** | **7** | **14** |
| Oxygen **O** | **8** | **8** | **8** | **8** | **16** |
| Fluorine **F** | **9** | **9** | **10** | **9** | **19** |
| Neon **Ne** | **10** | **10** | **10** | **10** | **20** |
| Sodium **Na** | **11** | **11** | **12** | **11** | **23** |
| Magnesium **Mg** | **12** | **12** | **12** | **12** | **24** |
| Aluminium **Al** | **13** | **13** | **14** | **13** | **27** |
| Silicon **Si** | **14** | **14** | **14** | **14** | **28** |
| Phosphorus **P** | **15** | **15** | **16** | **15** | **31** |
| Sulphur **S** | **16** | **16** | **16** | **16** | **32** |
| Chlorine **Cl** | **17** | **17** | **18** | **17** | **35** |
| Argon **Ar** | **18** | **18** | **22** | **18** | **40** |
| Potassium **K** | **19** | **19** | **20** | **19** | **39** |
| Calcium **Ca** | **20** | **20** | **20** | **20** | **40** |

Most atoms of elements exist as **isotopes.**

Isotopes are atoms of the same **element**, having the same number of **protons**/atomic number but **different** number of **neutrons**/mass number. By convention, isotopes are written with the mass number as superscript and the atomic number as subscript to the left of the chemical symbol of the element. i.e.

mass number

atomic number m n X symbol of element

Below is the conventional method of writing the 1st twenty elements showing the mass numbers and atomic numbers;

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| 1 H 42He  1 | 7 Li  3 | 9 Be  4 | 11 B  5 | 126C |
| 147N 168O | 19 F  9 | 20 Ne  10 | 23 Na  11 | 2412Mg |
| 2713 Al 2814Si    39 K 4020C  19 | 31 P  15 | 32 S  16 | 35 Cl  17 | 40 Ar  18 |
|  |  |  |  |  |

The table below shows some common **natural** isotopes of some elements

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Element | Isotopes | Protons | Electrons | Neutrons | Atomic number | Mass number |
| Hydrogen | 1. H   1   1. H(deuterium)   1   1. H(Tritium)   1 | 1  1  1 | 1  1  1 | 0  2  3 | 1  1  1 | 1  2  3 |
| Chlorine | 35 Cl  17  37 Cl  17 | 17  17 | 17  17 | 18  20 | 17  17 | 35  37 |
| Potassium | 1. K   19   1. K   19   1. K   19 | 19  19  19 | 19  19  19 | 20  21  22 | 19  19  19 | 39  40  41 |
| Oxygen | 16 O  8  18 O  8 | 8  8 | 8  8 | 8  10 | 8  8 | 16  18 |
| Uranium | 235 U  92  238 U  92 | 92  92 | 92  92 | 143  146 | 92  92 | 235  238 |
| Neon | 22 Ne  10   1. Ne   10   1. Ne   10 | 10  10  10 | 10  10  10 | 12  10  11 | 10  10  10 | 22  20  21 |

The mass of an average atom is very small (10-22 g).Masses of atoms are therefore expressed in relation to a chosen element.

The atom recommended is 12C isotope whose mass is arbitrarily assigned as 12.000 atomic mass units(a.m.u) .

All other atoms are compared to the mass of 12C isotope to give the **relative at** The relative atomic mass(RAM) is therefore defined as “**the mass of average atom of an element compared to 1/12 an atom of 12C isotope whose mass is arbitrarily fixed as 12.000 atomic mass units(a.m.u)** ” i.e;

RAM = mass of atom of an element

1/12 of one atom of 12C isotope

Accurate relative atomic masses (RAM) are got from the **mass spectrometer.** Mass spectrometer determines the **isotopes** of the element and their relative **abundance/availability.**

Using the relative abundances/availability of the isotopes, the relative atomic mass (RAM) can be determined /calculated as in the below examples.

1. Chlorine occurs as 75% 3517Cl and 25% 3717Cl isotopes. Calculate the relative atomic mass of Chlorine.

Working

* 1. atoms of chlorine contains 75 atoms of 3517Cl isotopes

100 atoms of chlorine contains 75 atoms of 3717Cl isotopes

Therefore;

RAM of chlorine = ( 75/100 x 35) + 25/100 x 37 = **35.5**  Note that:

**Relative atomic mass has no units**

**More atoms of chlorine exist as 3517Cl(75%) than as 3717Cl(25%) therefore RAM is nearer to the more abundant isotope.**

1. Calculate the relative atomic mass of potassium given that it exist as;

93.1% 3919K , 0.01% 4019K , 6.89% 4119K ,

Working

* 1. atoms of potassium contains 93.1 atoms of 3919K isotopes

100 atoms of potassium contains 0.01 atoms of 4019K isotopes

100 atoms of potassium contains 6.89 atoms of 4119K isotopes

Therefore;

RAM of potassium = (93.1/100 x39) + (0.01/100 x 40) +(6.89 /100 x 39)

=

Note that:

**Relative atomic mass has no units**

**More atoms of potassium exist as 3919K (93.1%) therefore RAM is nearer to the more abundant 3919K isotope.**

1. Calculate the relative atomic mass of Neon given that it exist as;

90.92% 2010Ne , 0.26% 2110Ne , 8.82% 2210Ne,

Working

* 1. atoms of Neon contains 90.92 atoms of 2010Ne isotopes

100 atoms of Neon contains 0.26 atoms of 2110Ne isotopes

100 atoms of Neon contains 8.82 atoms of 2210 Ne isotopes

Therefore;

RAM of Neon = (90.92/100 x20) + (0.26/100 x 21) +(8.82 /100 x 22)

=

Note that:

**Relative atomic mass has no units**

**More atoms of Neon exist as 2010Ne (**90.92**%) therefore RAM is nearer to the more abundant 2010Ne isotope.**

1. Calculate the relative atomic mass of Argon given that it exist as;

90.92% 2010Ne , 0.26% 2110Ne , 8.82% 2210Ne,

NB

The relative atomic mass is a measure of the masses of atoms. The higher the relative atomic mass, the heavier the atom.

Electrons are found in **energy levels/orbital.**

An energy level is a **fixed region** around/surrounding the nucleus of an atom occupied by electrons of the **same (potential) energy.**

**By convention** energy levels are named 1,2,3… **outwards** from the region **nearest** to nucleus.

Each energy level is occupied by a fixed number of electrons:

The **1st** energy level is occupied by a maximum of **two** electrons

The **2nd** energy level is occupied by a maximum of **eight** electrons The  **3rd**  energy level is occupied by a maximum of **eight** electrons( or **eighteen** electrons if available)

The  **4th**  energy level is occupied by a maximum of **eight** electrons( or **eighteen or thirty two** electrons if available)

This arrangement of electrons in an atom is called **electron configuration** / **structure**.

**By convention** theelectron configuration / structure of an atom of an element can be shown in form of a diagram using either cross(**x**) or dot(**●**) to **Practice examples drawing electronic configurations**

a)11H has - in nucleus**1**proton and **0** neutrons -1 electron in the 1st energy levels thus:

Nucleus

Energy levels

Electrons(represented by cross(x)

Electronic structure of Hydrogen is thus: **1:**

1. 42He has - in nucleus **2** proton and **2** neutrons -2 electron in the 1st energy levels thus:

**Comment [s1]:**

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Helium is thus: **2:**

1. 73Li has - in nucleus **3** proton and **4** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Lithium is thus: **2:1**

1. 94Be has - in nucleus **4** proton and **5** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Beryllium is thus: **2:2**

1. 115B has - in nucleus **5** proton and **6** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Boron is thus: **2:3**

1. 126C has - in nucleus **6** proton and **6** neutrons - **2** electron in the 1st energy levels

-**4** electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Carbon is thus: **2:4**

1. 147N has - in nucleus **7** proton and **7** neutrons - **2** electron in the 1st energy levels

-**5** electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Nitrogen is thus: **2:5**

1. 168O has - in nucleus **8** proton and **8** neutrons - **2** electron in the 1st energy levels

-**6** electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Oxygen is thus: **2:6**

1. 199F has - in nucleus **9** proton and **10** neutrons - **2** electron in the 1st energy levels

-**7** electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Fluorine is thus: **2:7**

1. 2010Ne has - in nucleus **10** proton and **10** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Neon is thus: **2:8**

1. 2311Na has - in nucleus **11** proton and **12** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**1** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Sodium is thus: **2:8:1**

1. 2412Mg has - in nucleus **12** proton and **12** neutrons

- **2** electron in the 1st energy levels

* 1. electron in the 2nd energy levels

-**2** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Magnesium is thus: **2:8:2**

1. 2713Al has - in nucleus **13** proton and **14** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**3** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Aluminium is thus: **2:8:3**

1. 2814Si has - in nucleus **14** proton and **14** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**4** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Silicon is thus: **2:8:4**

1. 3115P has - in nucleus **14** proton and **15** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**5** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Phosphorus is thus: **2:8:5**

1. 3216S has - in nucleus **16** proton and **16** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**6** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Sulphur is thus: **2:8:6**

1. 3517Cl has - in nucleus **18** proton and **17** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**7** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Chlorine is thus: **2:8:7**

1. 4018Ar has - in nucleus **22** proton and **18** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**8** electron in the 3rd energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Argon is thus: **2:8:8**

1. 3919K has - in nucleus **20** proton and **19** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**8** electron in the 3rd energy levels

-**1** electron in the 4th energy levels thus Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Potassium is thus: **2:8:8:1**

1. 4020Ca has - in nucleus **20** proton and **20** neutrons - **2** electron in the 1st energy levels
   1. electron in the 2nd energy levels

-**8** electron in the 3rd energy levels

-**2** electron in the 4th energy levels thus

Nucleus

Energy levels

Electrons (represented by dot(.)

Electronic structure of Calcium is thus: **2:8:8:2**

**B.PERIODIC TABLE**

There are over 100 elements so far discovered. Scientists have tried to group them together in a periodic table.

A periodic table is a horizontal and vertical arrangement of elements according to their atomic numbers.

This table was successfully arranged in 1913 by the British scientist Henry Moseley from the previous work of the Russian Scientist Dmitri Mendeleev. The horizontal arrangement forms **period**. Atoms in the same period have the same the same number of energy levels in their electronic structure. i.e. The number of **energy levels** in the electronic configuration of an element determine the **period** to which the element is in the periodic table. e.g.

Which period of the periodic table are the following isotopes/elements/atoms? a) 126C

Electron structure 2:4 => 2 energy levels used thus **Period 2**

1. 2311Na

Electron structure 2:8:1 => 3 energy levels used thus **Period 3**

1. 3919K

Electron structure 2:8:8:1 => 4 energy levels used thus **Period 4**

1. 11H

Electron structure 1: => 1 energy level used thus **Period 1**

The vertical arrangement of elements forms a **group.** Atoms in the same have the same the same group have the same number of outer energy level electrons as per their electronic structure. i.e.

The number of electrons in the outer energy level an element determine the **group** to which the element is ,in the periodic table.

a) 126C

Electron structure 2:**4** => 4 electrons in outer energy level thus **Group IV** b) 2311C

Electron structure 2:8:**1** => 1 electron in outer energy level thus **Group I**

1. 3919K

Electron structure 2:8:8:**1**=>1 electron in outer energy level thus **Group I**

1. 11H

Electron structure **1**: => 1 electron in outer energy level thus **Group I**

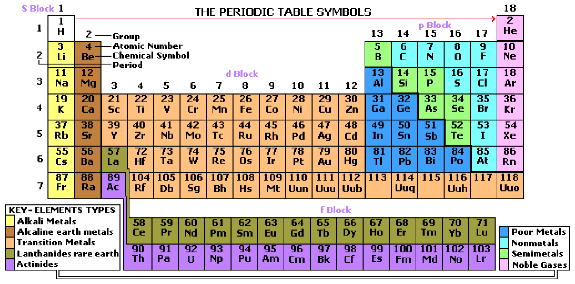
By convention;

(i)**Periods** are named using English numerals **1,2,3,4**,…

(ii)**Groups** are named using Roman numerals **I,II,III,IV**,…

There are eighteen groups in a standard periodic table.

There are seven periods in a standard periodic table.



**THE STANDARD PERIODIC TABLE OF ELEMENTS**

When an atom has maximum number of electrons in its outer energy level, it is said to be **stable.**

When an atom has no maximum number of electrons in its outer energy level, it is said to be **unstable.**

All stable atoms are in group **8/18** of the periodic table.All other elements are unstable.

All unstable atoms/isotopes try to be stable through chemical reactions. A chemical reaction involves gaining or losing outer electrons (electron transfer) .When electron transfer take place, an ion is formed.

An ion is formed when an unstable atom gain or donate electrons in its outer energy level inorder to be stable. Whether an atom gain or donate electrons depend on the relative energy required to donate or gain extra electrons i.e.

Examples

1. 199 F has electronic structure/configuration 2:7.

It can donate the seven outer electrons to have stable electronic structure/configuration 2:.

It can gain one extra electron to have stable electronic structure/configuration 2:8. Gaining requires less energy, and thus Fluorine reacts by gaining one extra electrons.

1. 2313 Al has electronic structure/configuration 2:8:3

It can donate the three outer electrons to have stable electronic structure/configuration 2:8.

It can gain five extra electrons to have stable electronic structure/configuration 2:8:8. Donating requires less energy, and thus Aluminium reacts by donating its three outer electrons.

Elements with **less** than four electrons in the outer energy level donates /lose the outer electrons to be stable and form a positively charged ion called **cation.** A cation therefore has more protons(positive charge) than electrons(negative charge)

Generally metals usually form cation

Elements with **more** than four electrons in the outer energy level gain /acquire extra electrons in the outer energy level to be stable and form a negatively charged ion called **anion.**

An anion therefore has less protons(positive charge) than electrons(negative charge)

Generally non metals usually form anion. Except Hydrogen

The charge carried by an ion is equal to the number of electrons gained/acquired or donated/lost.

Examples of ion formation

1.11H

|  |  |  |
| --- | --- | --- |
| H | | -> H+ + e |
| (atom) | | (monovalent cation) (electrons donated/lost) |
| Electronic configuration 1: | | (No electrons remains) |
| 2. 2713 Al |  |
| Al | -> | Al3+ + 3e |
| (atom) |  | (trivalent cation) (3 electrons donated/lost) |
| Electron 2:8:3 |  | 2:8 |
| structure (unstable)    3. 2311 Na |  | (stable) |
| Na | -> | Na+ + e |
| (atom) |  | (cation) ( 1 electrons donated/lost) |
| Electron 2:8:1 |  | 2:8 |
| structure (unstable) |  | (stable) |

4. 2412Mg

|  |  |  |  |
| --- | --- | --- | --- |
| Mg | -> Mg2+ + 2e | | |
| (atom) | (cation) ( 2 electrons donated/lost) | | |
| Electron 2:8:1 | 2:8 |  |  |
| structure (unstable)    5. 168O | (stable) |  |  |
| O + | 2e -> O2- |  |  |
| (atom) | ( 2 electrons gained/acquired) (anion) |  |  |
| Electron 2:6 | 2:8 |  |  |
| structure (unstable)    6. 147N | (stable) |  |  |
| N + | 3e -> N3- |  |  |
| (atom) | ( 3 electrons gained/acquired) (anion) |  |  |
| Electron 2:5 | 2:8 |  |  |
| structure (unstable)    7. 3115P | (stable) |  |  |
| P + | 3e -> P3- |  |  |
| (atom) | ( 3 electrons gained/acquired) (anion) |  |  |
| Electron 2:5 | 2:8 |  |  |
| structure (unstable)    8. 199F | (stable) |  |  |
| F + | e -> F- |  |  |
| (atom) | ( 1 electrons gained/acquired) (anion) |  |  |
| Electron 2:7 | 2:8 |  |  |
| structure (unstable)    9. 3517Cl | (stable) |  |  |
| Cl + | e -> Cl- |  |  |
| (atom) | ( 1 electrons gained/acquired) (anion) |  |  |
| Electron 2:8:7 | 2:8:8 |  |  |
| structure (unstable)    3. 3919 K | (stable) |  |  |
| K | -> K+ + e |  |  |
| (atom) | (cation) ( 1 electrons donated/lost) | | |
| Electron 2:8:8:1 | 2:8:8 | | |
| structure (unstable) | (stable) | | |

When an element donate/loses its outer electrons ,the process is called oxidation. When an element acquires/gains extra electrons in its outer energy level,the process is called reduction.The charge carried by an atom, cation or anion is its oxidation state.

Table showing the oxidation states of some isotopes

|  |  |  |  |
| --- | --- | --- | --- |
| Element | Symbol of element / isotopes | Charge of ion | Oxidation state |
| Hydrogen | 1. H   1   1. H(deuterium)   1   1. H(Tritium)   1 | H+  H+  H+ | +1  +1  +1 |
| Chlorine | 35 Cl  17  37 Cl  17 | Cl-  Cl- | -1  -1 |
| Potassium | 1. K   19   1. K   19   1. K   19 | K+  K+  K+ | +1  +1  +1 |
| Oxygen | 16 O  8  18 O  8 | O2-  O2- | -2  -2 |
| Magnesium | 24 Mg  12 | Mg2+ | +2 |
| sodium | 23 Na  11 | Na+ | +1 |
| Copper | Cu | Cu + Cu2+ | +1  +2 |
| Iron |  | Fe 2+ Fe3+ | +2  +3 |
| Lead |  | Pb 2+ Pb4+ | +2  +4 |
| Manganese |  | Mn 2+  Mn7+ | +2  +7 |
| Chromium |  | Cr 3+ Cr6+ | +3  +6 |
| Sulphur |  | S4+  S6+ | +4  +6 |
| Carbon |  | C 2+ C4+ | +2  +4 |

Note :

Some elements can exist in **more** than one oxidation state.They are said to have variable oxidation state.

Roman capital numeral is used to indicate the oxidation state of an element with a variable oxidation state in a compound.

Examples:

1. Copper (I) means Cu+ as in Copper(I)oxide
2. Copper (II) means Cu2+ as in Copper(II)oxide
3. Iron (II) means Fe2+ as in Iron(II)sulphide
4. Iron (III) means Fe3+ as in Iron(III)chloride
5. Sulphur(VI)mean S6+ as in Iron(III)sulphate(VI)
6. Sulphur(VI)mean S6+ as in sulphur(VI)oxide
7. Sulphur(IV)mean S4+ as in sulphur(IV)oxide
8. Sulphur(IV)mean S4+ as in sodium sulphate(IV)
9. Carbon(IV)mean C4+ as in carbon(IV)oxide
10. Carbon(IV)mean C4+ as in Lead(II)carbonate(IV)
11. Carbon(II)mean C2+ as in carbon(II)oxide
12. Manganese(IV)mean Mn4+ as in Manganese(IV)oxide

A compound is a combination of two or more elements in fixed proportions.

The ratio of the atoms making a compound is called the chemical formulae. Elements combine together to form a compound depending on their combining power.

The combining power of atoms in an element is called Valency.Valency of an element is equal to the **number** of:

(i)hydrogen atoms that an atom of element can combine with or displace. (ii)electrons gained /acquired in outer energy level by non metals to be stable/attain duplet/octet.

(iii)electrons donated/lost by outer energy level of metals to be stable/attain octet/duplet.

(iv)charges carried by ions/cations/ions

Group of atoms that react as a unit during chemical reactions are called **radicals**.Elements with variable oxidation state also have more than one valency.

**Table showing the valency of common radicals.**

|  |  |  |
| --- | --- | --- |
| **Radical name** | **Chemical formulae** | **Combining power / Valency** |
| Ammonium | NH4 + | 1 |
| Hydroxide | OH- | 1 |
| Nitrate(V) | NO3 - | 1 |
| Hydrogen carbonate | HCO3- | 1 |
| Hydrogen sulphate(VI) | HSO4- | 1 |
| Hydrogen sulphate(IV) | HSO3- | 1 |
| Manganate(VII) | MnO4- | 1 |
| Chromate(VI) | CrO42- | 2 |
| Dichromate(VI) | Cr2O72- | 2 |
| Sulphate(VI) | SO42- | 2 |
| Sulphate(IV) | SO32- | 2 |
| Carbonate(IV) | CO32- | 2 |
| Phosphate(V) | PO42- | 3 |

**Table showing the valency of some common metal and non metals**

|  |  |  |  |
| --- | --- | --- | --- |
| Element/metal | Valency | **Element/non metal** | **Valency** |
| Hydrogen | 1 | **Florine** | **1** |
| Lithium | 1 | **Chlorine** | **1** |
| Beryllium | 2 | **Bromine** | **1** |
| Boron | 3 | **Iodine** | **1** |
| Sodium | 1 | **Carbon** | **4** |
| Magnesium | 2 | **Nitrogen** | **3** |
| Aluminium | 3 | **Oxygen** | **2** |
| Potassium | 1 | **Phosphorus** | **3** |
| Calcium | 2 |  |  |
| Zinc | 2 |  |  |
| Barium | 2 |  |  |
| Mercury | 2 |  |  |
| Iron | 2 and 3 |  |  |
| Copper | 1 and 2 |  |  |
| Manganese | 2 and 4 |  |  |
| Lead | 2 and 4 |  |  |

From the valency of elements , the chemical formular of a compound can be derived using the following procedure:

(i)Identify the elements and radicals making the compound

(ii)Write the symbol/formular of the elements making the compound starting with the metallic element

(iii)Assign the valency of each element /radical as superscript.

(iv)Interchange/exchange the valencies of each element as subscript.

(v)Divide by the smallest/lowest valency to derive the smallest whole number ratios

Ignore a valency of 1.

This is the chemical formula.

**Practice examples**

**Write the chemical formula of**

**(a)Aluminium oxide**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Aluminium | Oxygen |
| Symbol of elements/radicals in compound | Al | O |
| Assign valencies as superscript | Al**3** | O**2** |
| Exchange/Interchange the valencies as subscript | Al**2** | O**3** |
| Divide by smallest valency to get whole number | - | - |

Chemical formula of Aluminium oxide is thus**:** **Al2 O3**

**This means:2atoms of Aluminium combine with 3 atoms of Oxygen**

**(b)Sodium oxide**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Sodium | Oxygen |
| Symbol of elements/radicals in compound | Na | O |
| Assign valencies as superscript | Na**1** | O**2** |
| Exchange/Interchange the valencies as subscript | Na**2** | O**1** |
| Divide by smallest valency to get whole number | - | - |

Chemical formula of Sodium oxide is thus**:** **Na2 O**

**This means:2atoms of Sodium combine with 1 atom of Oxygen**

**(c)Calcium oxide**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Calcium | Oxygen |
| Symbol of elements/radicals in compound | Ca | O |
| Assign valencies as superscript | Ca**2** | O**2** |
| Exchange/Interchange the valencies as subscript | Ca**2** | O**2** |
| Divide by two to get smallest whole number ratio | Ca**1** | O**1** |

Chemical formula of Calcium oxide is thus**:** **CaO**

**This means:1 atom of calcium combine with 1 atom of Oxygen.**

**(d)Lead(IV)oxide**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Lead | Oxygen |
| Symbol of elements/radicals in compound | Pb | O |
| Assign valencies as superscript | Pb**4** | O**2** |
| Exchange/Interchange the valencies as subscript | Pb**2** | O**4** |
| Divide by two to get smallest whole number ratio | Pb**1** | O**2** |

Chemical formula of Lead(IV) oxide is thus**:** **PbO2**

**This means:1 atom of lead combine with 2 atoms of Oxygen.**

**(e)Lead(II)oxide**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Lead | Oxygen |
| Symbol of elements/radicals in compound | Pb | O |
| Assign valencies as superscript | Pb**2** | O**2** |
| Exchange/Interchange the valencies as subscript | Pb**2** | O**2** |
| Divide by two to get smallest whole number ratio | Pb**1** | O**1** |

Chemical formula of Lead(II) oxide is thus**:** **PbO**

**This means:1 atom of lead combine with 1 atom of Oxygen.**

**(e)Iron(III)oxide**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Iron | Oxygen |
| Symbol of elements/radicals in compound | Fe | O |
| Assign valencies as superscript | Fe**3** | O**2** |
| Exchange/Interchange the valencies as subscript | Fe**2** | O**3** |
| Divide by two to get smallest whole number ratio | - | - |

Chemical formula of Iron(III) oxide is thus**:** **Fe2O3**

**This means:2 atom of lead combine with 3 atom of Oxygen.**

**(f)Iron(II)sulphate(VI)**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Iron | sulphate(VI) |
| Symbol of elements/radicals in compound | Fe | SO**4** |
| Assign valencies as superscript | Fe**2** | SO**4 2** |
| Exchange/Interchange the valencies as subscript | Fe**2** | SO**4 2** |
| Divide by two to get smallest whole number ratio | Fe1 | SO**4 1** |

Chemical formula of Iron(II) sulphate(VI) is thus**:** **FeSO4**

**This means:1 atom of Iron combine with 1 sulphate(VI) radical.**

**(g)Copper(II)sulphate(VI)**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Copper | sulphate(VI) |
| Symbol of elements/radicals in compound | Cu | SO**4** |
| Assign valencies as superscript | Cu**2** | SO**4 2** |
| Exchange/Interchange the valencies as subscript | Cu**2** | SO**4 2** |
| Divide by two to get smallest whole number ratio | Cu1 | SO**4 1** |

Chemical formula of Cu(II)sulphate(VI) is thus**:** **CuSO4**

**This means:1 atom of Copper combine with 1 sulphate(VI) radical.**

**(h)Aluminium sulphate(VI)**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Aluminium | sulphate(VI) |
| Symbol of elements/radicals in compound | Al | SO**4** |
| Assign valencies as superscript | Al**3** | SO**4 2** |
| Exchange/Interchange the valencies as subscript | Al**2** | SO**4 3** |
| Divide by two to get smallest whole number ratio | - | - |

Chemical formula of Aluminium sulphate(VI) is thus**:** **Al2(SO4)3**

**This means:2 atom of Aluminium combine with 3 sulphate(VI) radical.**

**(i)Aluminium nitrate(V)**

|  |  |  |
| --- | --- | --- |
| Elements making compound | Aluminium | nitrate(V) |
| Symbol of elements/radicals in compound | Al | NO**3** |
| Assign valencies as superscript | Al**3** | NO**3 1** |
| Exchange/Interchange the valencies as subscript | Al**1** | NO**3 3** |
| Divide by two to get smallest whole number ratio | - | - |

Chemical formula of Aluminium sulphate(VI) is thus**:** **Al (NO3)3** **This means:1 atom of Aluminium combine with 3 nitrate(V) radical.**

**(j)Potassium manganate(VII)**

|  |  |  |  |
| --- | --- | --- | --- |
| Elements making compound | Potassium | | manganate(VII) |
| Symbol of elements/radicals in compound | K | | MnO**4** |
| Assign valencies as superscript | K **1** | | MnO**4 1** |
| Exchange/Interchange the valencies as subscript | K**1** | | MnO**4 1** |
| Divide by two to get smallest whole number ratio | | - | - |

Chemical formula of Potassium manganate(VII) is thus**:** **KMnO4**

**This means:1 atom of Potassium combine with 4 manganate(VII) radical.**

**(k)Sodium dichromate(VI)**

|  |  |  |  |
| --- | --- | --- | --- |
| Elements making compound | Sodium | | dichromate(VI) |
| Symbol of elements/radicals in compound | Na | | Cr2O**7** |
| Assign valencies as superscript | Na **1** | | Cr2O**7 2** |
| Exchange/Interchange the valencies as subscript | Na**2** | | Cr2O**7 1** |
| Divide by two to get smallest whole number ratio | | - | - |

Chemical formula of Sodium dichromate(VI) is thus**:** **Na2 Cr2O7**

**This means:2 atom of Sodium combine with 1 dichromate(VI) radical.**

**(l)Calcium hydrogen carbonate**

|  |  |  |  |
| --- | --- | --- | --- |
| Elements making compound | Calcium | | Hydrogen carbonate |
| Symbol of elements/radicals in compound | Ca | | CO**3** |
| Assign valencies as superscript | Ca **2** | | HCO**3 1** |
| Exchange/Interchange the valencies as subscript | Ca**1** | | HCO**3 2** |
| Divide by two to get smallest whole number ratio | | - | - |

Chemical formula of Calcium hydrogen carbonate is thus**:** **Ca(HCO3)2** **This means:1 atom of Calcium combine with 2 hydrogen carbonate radical.**

**(l)Magnesium hydrogen sulphate(VI)**

|  |  |  |  |
| --- | --- | --- | --- |
| Elements making compound | Magnesium | | Hydrogen sulphate(VI) |
| Symbol of elements/radicals in compound | Mg | | HSO**4** |
| Assign valencies as superscript | Mg **2** | | HSO**4 1** |
| Exchange/Interchange the valencies as subscript | Mg**1** | | HSO**4 2** |
| Divide by two to get smallest whole number ratio | | - | - |

Chemical formula of Magnesium hydrogen sulphate(VI) is thus**:** **Mg(HSO4)2** **This means:1 atom of Magnesium combine with 2 hydrogen sulphate(VI) radical.**

Compounds are formed from chemical reactions. A chemical reaction is formed when atoms of the reactants break free to bond again and form products. A chemical reaction is a statement showing the movement of reactants to form products. The following procedure is used in writing a chemical equations:

1. Write the word equation
2. Write the correct chemical formula for each of the reactants and products 3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.
3. Multiply the chemical formula containing the unbalanced atoms with the lowest common multiple if the number of atoms on one side is not equal. This is called **balancing.**

**Do not change the chemical formula of the products/reactants.**

1. Assign in brackets, the physical state/state symbols of the reactants and products after each chemical formula as:
   * 1. **(s)** for solids
     2. **(l)** for liquids (iii) **(g)** for gas

(iv) **(aq)** for aqueous/dissolved in water to make a solution.

**Practice examples**

**Write a balanced chemical equation for the following**

* 1. Hydrogen gas is prepared from reacting Zinc granules with dilute hydrochloric acid.

Procedure

* 1. Write the word equation

**Zinc + Hydrochloric acid -> Zinc chloride + hydrogen gas**

* 1. Write the correct chemical formula for each of the reactants and products

**Zn + HCl -> ZnCl2 + H2**

* 1. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.  **Number of atoms of Zn on the reactant side is equal to product side One atom of H in HCl on the reactant side is not equal to two atoms in H2 on product side.**

**One atom of Cl in HCl on the reactant side is not equal to two atoms in ZnCl2 on product side.**

* 1. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

**Multiply HCl by “2” to get “2” Hydrogen and “2” Chlorine on product and reactant side.**

**Zn + 2 HCl -> ZnCl2 + H2**

* 1. Assign in brackets, the physical state/state symbols .

**Zn(s) + 2 HCl(aq) -> ZnCl2 (aq) + H2(g)**

* 1. Oxygen gas is prepared from decomposition of Hydrogen peroxide solution to water

Procedure

* 1. Write the word equation

**Hydrogen peroxide -> Water + oxygen gas**

* 1. Write the correct chemical formula for each of the reactants and products

**H2O2 -> H2O + O2**

* 1. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.  **Number of atoms of H on the reactant side is equal to product side Two atom of O in H2O2 on the reactant side is not equal to three atoms (one in H2O and two in O2) on product side.**

* 1. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

**Multiply H2O2 by “2” to get “4” Hydrogen and “4” Oxygen on reactants Multiply H2O by “2” to get “4” Hydrogen and “2” Oxygen on product side**

**When the “2” Oxygen in O2 and the“2” in H2O are added on product side they are equal to the“4” Oxygen on reactants side.**

**2H2O2 -> 2H2O + O2**

* 1. Assign in brackets, the physical state/state symbols .

**2H2O2(aq) -> 2H2O(l) + O2(g)**

* 1. Chlorine gas is prepared from Potassium manganate(VII) reacting with hydrochloric acid to form potassium chloride solution, manganese(II) chloride solution,water and chlorine gas.

Procedure

* 1. Write the word equation

**Potassium manganate(VII) + Hydrochloric acid ->** potassium chloride + manganese(II) chloride + chlorine **+**water

* 1. Write the correct chemical formula for each of the reactants and products **KMnO4 + HCl -> KCl + MnCl2 +H2O + Cl2**

* 1. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

**Number of atoms of K and Mn on the reactant side is equal to product side**

**Two atom of H in H2O on the product side is not equal to one atom on reactant side.**

**Four atom of O in KMnO4 is not equal to one in H2O**

**One atom of Cl in HCl on reactant side is not equal to three (one in H2O and two in Cl2)**

* 1. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.  **Multiply HCl by “16” to get “16” Hydrogen and “16” Chlorine on reactants**

**Multiply KMnO4  by “2” to get “2” Potassium and “2” manganese, “2 x4 =8” Oxygen on reactant side.**

**Balance the product side to get:**

**2 KMnO4 +16 HCl -> 2 KCl + 2 MnCl2 +8 H2O + 5 Cl2**

5. Assign in brackets, the physical state/state symbols .

**2KMnO4(s) +16 HCl(aq)-> 2 KCl (aq) + 2MnCl2(aq)+8 H2O(l)+5 Cl2(g)**

(d)Carbon(IV)oxide gas is prepared from Calcium carbonate reacting with hydrochloric acid to form calcium chloride solution, water and carbon(IV)oxide gas.

Procedure

* 1. Write the word equation

**Calcium carbonate + Hydrochloric acid ->**

**calcium chloride solution+ water +carbon(IV)oxide**

* 1. Write the correct chemical formula for each of the reactants and products

**CaCO3 + HCl -> CaCl2 +H2O + CO2**

* 1. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

* 1. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

* 1. Assign in brackets, the physical state/state symbols .

**CaCO3(s) + 2 HCl(aq) -> CaCl2(aq) + H2O(l) + CO2(g)**

(d)Sodium hydroxide solution neutralizes hydrochloric acid to form salt and water.

**NaOH(aq) + HCl(aq) -> NaCl (aq) + H2O(l)**

(e)Sodium reacts with water to form sodium hydroxide and hydrogen gas. **2Na(s) + 2H2O(l) -> 2NaOH(aq) + H2(g)**

(f)Calcium reacts withwater to form calcium hydroxide and hydrogen gas

**Ca(s) + 2H2O(l) -> Ca(OH)2(aq) + H2(g)**

(g)Copper(II)Oxide solid reacts with dilute hydrochloric acid to form copper(II)chloride and water.

**CuO(s) + 2HCl(aq) -> CuCl2(aq) + H2O(l)**

(h)Hydrogen sulphide reacts with Oxygen to form sulphur(IV)Oxide and water.

**2H2S(g) + 3O2(g) -> 2SO2(g) + 2H2O(l)**

(i)Magnesium reacts with steam to form Magnesium Oxide and Hydrogen gas.

**Mg(s) + 2H2O(g) -> MgO(s) + H2(g)**

(j)Ethane(C2H6) gas burns in air to form Carbon(IV)Oxide and water.

**2C2H6(g) + 7O2(g) -> 4CO2(g) + 6H2O(l)**

(k)Ethene(C2H4) gas burns in air to form Carbon(IV)Oxide and water.

**C2H4(g) + 3O2(g) -> 2CO2(g) + 2H2O(l)**

(l)Ethyne(C2H2) gas burns in air to form Carbon(IV)Oxide and water.

2**C2H2(g) + 5O2(g) -> 4CO2(g) + 2H2O(l)**

**C.PERIODICITY OF CHEMICAL FAMILES/DOWN THE GROUP.**

The number of valence electrons and the number of occupied energy levels in an atom of an element determine the position of an element in the periodic table.i.e

The number of occupied energy levels determine the Period and the valence electrons determine the Group.

Elements in the same group have similar physical and chemical properties. The trends in physical and chemical properties of elements in the same group vary down the group. Elements in the same group thus constitute a chemical family.

(a) Group I elements: Alkali metals

Group I elements are called **Alkali metals** except Hydrogen which is a non metal. The alkali metals include:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Electron structure | Oxidation state | Valency |
| Lithium | Li | 3 | 2:**1** | Li**+** | 1 |
| Sodium | Na | 11 | 2:8:**1** | Na**+** | 1 |
| Potassium | K | 19 | 2:8:8:**1** | K**+** | 1 |
| Rubidium | Rb | 37 | 2:8:18:8:**1** | Rb**+** | 1 |
| Caesium | Cs | 55 | 2:8:18:18:8:**1** | Cs**+** | 1 |
| Francium | Fr | 87 | 2:8:18:32:18:8:**1** | Fr**+** | 1 |

All alkali metals atom has one electron in the outer energy level. They therefore are **monovalent**. They donate /lose the outer electron to have oxidation state M+ The number of energy levels increases down the group from Lithium to Francium. The more the number of energy levels the bigger/larger the atomic size. e.g.

The atomic size of Potassium is bigger/larger than that of sodium because Potassium has more/4 energy levels than sodium (3 energy levels).

Atomic and ionic radius

The distance between the centre of the nucleus of an **atom** and the outermost energy level occupied by electron/s is called **atomic radius.** Atomic radius is measured in **nanometers**(n).The higher /bigger the atomic radius the bigger /larger the atomic size.

The distance between the centre of the nucleus of an **ion** and the outermost energy level occupied by electron/s is called **ionic radius.** Ionic radius is also measured in **nanometers**(n).The higher /bigger the ionic radius the bigger /larger the size of the ion.

Atomic radius and ionic radius depend on the number of energy levels occupied by electrons. The more the number of energy levels the bigger/larger the atomic /ionic radius. e.g.

The atomic radius of Francium is bigger/larger than that of sodium because Francium has more/7 energy levels than sodium (3 energy levels).

Atomic radius and ionic radius of alkali metals increase down the group as the number of energy levels increases.

The atomic radius of alkali metals is bigger than the ionic radius. This is because alkali metals react by losing/donating the outer electron and hence lose the outer energy level.

**Table showing the atomic and ionic radius of some alkali metals**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Atomic radius(nM) | Ionic radius(nM) |
| Lithium | Li | 3 | 0.133 | 0.060 |
| Sodium | Na | 11 | 0.157 | 0.095 |
| Potassium | K | 19 | 0.203 | 0.133 |

The atomic radius of sodium is 0.157nM .The ionic radius of Na+ is 0.095nM. This is because sodium reacts by donating/losing the outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the atomic radius.

Electropositivity

The ease of donating/losing electrons is called electropositivity. All alkali metals are electropositive. Electropositivity increase as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The outer electrons experience less nuclear attraction and can be lost/ donated easily/with ease. Francium is the most electropositive element in the periodic table because it has the highest/biggest atomic radius.

Ionization energy

The minimum amount of energy required to remove an electron from an atom of element in its gaseous state is called **1st ionization energy**. The SI unit of ionization energy is **kilojoules per mole/kJmole-1 .**Ionization energy depend on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer electrons/energy level and thus the lower the ionization energy. For alkali metals the 1st ionization energy decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease.

e.g. The 1st ionization energy of sodium is 496 kJmole-1while that of potassium is 419 kJmole-1 **.**This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from sodium to Potassium. It requires therefore less energy to donate/lose outer electrons in Potassium than in sodium.

Physical properties

**Soft/Easy to cut**: Alkali metals are soft and easy to cut with a knife. The softness and ease of cutting increase down the group from Lithium to Francium. This is because an increase in atomic radius, decreases the strength of metallic bond and the packing of the metallic structure

**Appearance:** Alkali metals have a shiny grey metallic luster when freshly cut. The surface rapidly/quickly tarnishes on exposure to air. This is because the metal surface rapidly/quickly reacts with elements of air/oxygen.

**Melting and boiling points:** Alkali metals have a relatively low melting/boiling point than common metals like Iron. This is because alkali metals use only one delocalized electron to form a weak metallic bond/structure.

**Electrical/thermal conductivity:** Alkali metals are good thermal and electrical conductors. Metals conduct using the outer mobile delocalized electrons. The delocalized electrons move randomly within the metallic structure.

**Summary of some physical properties of the 1st three alkali metals**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Alkali metal | Appearance | Ease of cutting | Melting point (oC) | Boiling point (oC) | Conductivity | 1st  ionization energy |
| Lithium | Silvery white | Not easy | 180 | 1330 | Good | 520 |
| Sodium | Shiny grey | Easy | 98 | 890 | Good | 496 |
| Potassium | Shiny grey | Very easy | 64 | 774 | Good | 419 |

Chemical properties

**(i)Reaction with air/oxygen**

On exposure to air, alkali metals reacts with the elements in the air. Example

On exposure to air, Sodium first reacts with Oxygen to form sodium oxide.

4Na(s) + O2(g) -> 2Na2O(s)

The sodium oxide formed further reacts with water/moisture in the air to form sodium hydroxide solution.

Na2O(s) + H2O(l) -> 2NaOH(aq)

Sodium hydroxide solution reacts with carbon(IV)oxide in the air to form sodium carbonate.

2NaOH(aq) + CO2(g) -> Na2CO3(g) + H2O(l)

**(ii)Burning in air/oxygen**

Lithium burns in air with a **crimson**/deep red flame to form Lithium oxide

4Li (s) + O2(g) -> 2Li2O(s)

Sodium burns in air with a **yellow** flame to form sodium oxide

4Na (s) + O2(g) -> 2Na2O(s)

Sodium burns in oxygen with a **yellow** flame to form sodium peroxide

2Na (s) + O2(g) -> Na2O2 (s)

Potassium burns in air with a **lilac/purple** flame to form potassium oxide

4K (s) + O2(g) -> 2K2O (s)

**(iii) Reaction with water:**

Experiment

Measure 500 cm3 of water into a beaker.

Put three drops of phenolphthalein indicator.

Put about 0.5g of Lithium metal into the beaker.

Determine the pH of final product

Repeat the experiment using about 0.1 g of Sodium and Potassium.

**Caution**: Keep a distance

Observations

|  |  |  |
| --- | --- | --- |
| Alkali metal | Observations | Comparative speed/rate of the reaction |
| Lithium | -Metal floats in water  -rapid effervescence/fizzing/bubbling -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn  phenolphthalein indicator pink  -pH of solution = 12/13/14 | Moderately vigorous |
| Sodium | -Metal floats in water  -very rapid effervescence /fizzing  /bubbling  -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn  phenolphthalein indicator pink  -pH of solution = 12/13/14 | Very vigorous |
| Potassium | -Metal floats in water  -explosive effervescence /fizzing  /bubbling  -colourless gas produced (that | Explosive/burst into flames |
|  | extinguishes burning splint with explosion /“pop” sound) -resulting solution turn  phenolphthalein indicator pink  -pH of solution = 12/13/14 |  |

Explanation

Alkali metals are less dense than water. They therefore float in water.They react with water to form a strongly alkaline solution of their hydroxides and producing hydrogen gas. The rate of this reaction increase down the group. i.e.

Potassium is more reactive than sodium .Sodium is more reactive than Lithium.

The reactivity increases as electropositivity increases of the alkali increases. This is because as the atomic radius increases , the ease of donating/losing outer electron increase during chemical reactions.

Chemical equations

2Li(s) + 2H2O(l) -> 2LiOH(aq) + H2(g)

2Na(s) + 2H2O(l) -> 2NaOH(aq) + H2(g)

2K(s) + 2H2O(l) -> 2KOH(aq) + H2(g)

2Rb(s) + 2H2O(l) -> 2RbOH(aq) + H2(g)

2Cs(s) + 2H2O(l) -> 2CsOH(aq) + H2(g)

2Fr(s) + 2H2O(l) -> 2FrOH(aq) + H2(g)

Reactivity increase down the group

**(iv) Reaction with chlorine:**

Experiment

Cut about 0.5g of sodium into a deflagrating spoon with a lid cover. Introduce it on a Bunsen flame until it catches fire. Quickly and carefully lower it into a gas jar containing dry chlorine to cover the gas jar.

Repeat with about 0.5g of Lithium.

**Caution:** This experiment should be done in fume chamber because chlorine is poisonous /toxic.

Observation

Sodium metal continues to burn with a yellow flame forming white solid/fumes. Lithium metal continues to burn with a crimson flame forming white solid / fumes.

Alkali metal react with chlorine gas to form the corresponding metal chlorides.

The reactivity increase as electropositivity increase down the group from Lithium to Francium.The ease of donating/losing the outer electrons increase as the atomic radius increase and the outer electron is less attracted to the nucleus.

Chemical equations

2Li(s) + Cl2(g) -> 2LiCl(s)

2Na(s) + Cl2(g) -> 2NaCl(s)

2K(s) + Cl2(g) -> 2KCl(s)

2Rb(s) + Cl2(g) -> 2RbCl(s)

2Cs(s) + Cl2(g) -> 2CsCl(s)

2Fr(s) + Cl2(g) -> 2FrCl(s)

Reactivity increase down the group

The table below shows some compounds of the 1st three alkali metals

|  |  |  |  |
| --- | --- | --- | --- |
|  | Lithium | sodium | Potassium |
| Hydroxide | LiOH | NaOH | KOH |
| Oxide | Li2O | Na2O | K2O |
| Sulphide | Li2S | Na2S | K2S |
| Chloride | LiCl | NaCl | KCl |
| Carbonate | Li2CO3 | Na2CO3 | K2CO3 |
| Nitrate(V) | LiNO3 | NaNO3 | KNO3 |
| Nitrate(III) | - | NaNO2 | KNO2 |
| Sulphate(VI) | Li2SO4 | Na2SO4 | K2SO4 |
| Sulphate(IV) | - | Na2SO3 | K2SO3 |
| Hydrogen carbonate | - | NaHCO3 | KHCO3 |
| Hydrogen sulphate(VI) | - | NaHSO4 | KHSO4 |
| Hydrogen sulphate(IV) | - | NaHSO3 | KHSO3 |
| Phosphate | - | Na3PO4 | K3PO4 |
| Manganate(VI) | - | NaMnO4 | KMnO4 |
| Dichromate(VI) | - | Na2Cr2O7 | K2Cr2O7 |
| Chromate(VI) | - | Na2CrO4 | K2CrO4 |

Some **uses** of alkali metals include:

(i)Sodium is used in making sodium cyanide for extracting gold from gold ore.

(ii)Sodium chloride is used in seasoning food.

(iii)Molten mixture of sodium and potassium is used as coolant in nuclear reactors.

(iv)Sodium is used in making sodium hydroxide used in making soapy and soapless detergents.

(v)Sodium is used as a reducing agent for the extraction of titanium from Titanium(IV)chloride.

(vi)Lithium is used in making special high strength glasses

(vii)Lithium compounds are used to make dry cells in mobile phones and computer laptops.

Group II elements: Alkaline earth metals

Group II elements are called **Alkaline earth metals** . The alkaline earth metals include:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Electron structure | Oxidation state | Valency |
| Beryllium | Be | 4 | 2:**2** | Be**2+** | 2 |
| Magnesium | Mg | 12 | 2:8:**2** | Mg**2+** | 2 |
| Calcium | Ca | 20 | 2:8:8:**2** | Ca**2+** | 2 |
| Strontium | Sr | 38 | 2:8:18:8:**2** | Sr**2+** | 2 |
| Barium | Ba | 56 | 2:8:18:18:8:**2** | Ba**2+** | 2 |
| Radium | Ra | 88 | 2:8:18:32:18:8:**2** | Ra**2+** | 2 |

All alkaline earth metal atoms have two electrons in the outer energy level. They therefore are **divalent**. They donate /lose the two outer electrons to have oxidation state M2+

The number of energy levels increases down the group from Beryllium to Radium. The more the number of energy levels the bigger/larger the atomic size. e.g.

The atomic size/radius of Calcium is bigger/larger than that of Magnesium because Calcium has more/4 energy levels than Magnesium (3 energy levels).

Atomic radius and ionic radius of alkaline earth metals increase down the group as the number of energy levels increases.

The atomic radius of alkaline earth metals is bigger than the ionic radius. This is because they react by losing/donating the two outer electrons and hence lose the outer energy level.

**Table showing the atomic and ionic radius of the 1st three alkaline earth metals**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Atomic radius(nM) | Ionic radius(nM) |
| Beryllium | Be | 4 | 0.089 | 0.031 |
| Magnesium | Mg | 12 | 0.136 | 0.065 |
| Calcium | Ca | 20 | 0.174 | 0.099 |

The atomic radius of Magnesium is 0.136nM .The ionic radius of Mg2+ is 0.065nM. This is because Magnesium reacts by donating/losing the two outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the atomic radius.

Electropositivity

All alkaline earth metals are also electropositive like alkali metals. The electropositivity increase with increase in atomic radius/size. Calcium is more electropositive than Magnesium. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The two outer electrons in calcium experience less nuclear attraction and can be lost/ donated easily/with ease because of the higher/bigger atomic radius.

Ionization energy

For alkaline earth metals the 1st ionization energy decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease.

e.g. The 1st ionization energy of Magnesium is 900 kJmole-1while that of Calcium is 590 kJmole-1 **.**This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from magnesium to calcium.

It requires therefore less energy to donate/lose outer electron in calcium than in magnesium.

The minimum amount of energy required to remove a second electron from an ion of an element in its gaseous state is called the **2nd ionization energy**. The 2nd ionization energy is always higher /bigger than the 1st ionization energy.

This because once an electron is donated /lost form an atom, the overall effective nuclear attraction on the remaining electrons/energy level increase. Removing a second electron from the ion require therefore more energy than the first electron.

The atomic radius of alkali metals is higher/bigger than that of alkaline earth metals.This is because across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level. Increase in nuclear charge increases the effective nuclear attraction on the outer energy level which pulls it closer to the nucleus. e.g.

Atomic radius of Sodium (0.157nM) is higher than that of Magnesium (0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus.

Physical properties

**Soft/Easy to cut**: Alkaline earth metals are **not** soft and easy to cut with a knife like alkali metals. This is because of the decrease in atomic radius of corresponding alkaline earth metal, increases the strength of metallic bond and the packing of the metallic structure. Alkaline earth metals are

(i)ductile(able to form wire/thin long rods)

(ii)malleable(able to be hammered into sheet/long thin plates)

(iii)have high tensile strength(able to be coiled without breaking/ not brittle/withstand stress)

**Appearance:** Alkali earth metals have a shiny grey metallic luster when their surface is freshly polished /scrubbed. The surface slowly tarnishes on exposure to air. This is because the metal surface slowly undergoes oxidation to form an oxide. This oxide layer should be removed before using the alkaline earth metals.

**Melting and boiling points:** Alkaline earth metals have a relatively high melting/ boiling point than alkali metals. This is because alkali metals use only one delocalized electron to form a weaker metallic bond/structure. Alkaline earth metals use two delocalized electrons to form a stronger metallic bond /structure.

Themelting and boiling points decrease down the group as the atomic radius/size increase reducing the strength of metallic bond and packing of the metallic structure. e.g.

Beryllium has a melting point of 1280oC. Magnesium has a melting point of 650oC.Beryllium has a smaller atomic radius/size than magnesium .The strength of metallic bond and packing of the metallic structure is thus stronger in beryllium.

**Electrical/thermal conductivity:** Alkaline earth metals are good thermal and electrical conductors. The two delocalized valence electrons move randomly within the metallic structure.

Electrical conductivity increase down the group as the atomic radius/size increase making the delocalized outer electrons less attracted to nucleus. Alkaline earth metals are better thermal and electrical conductors than alkali metals because they have more/two outer delocalized electrons.e.g.

Magnesium is a better conductor than sodium because it has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

Calcium is a better conductor than magnesium.

Calcium has bigger/larger atomic radius than magnesium because the delocalized electrons are less attracted to the nucleus of calcium and thus more free /mobile and thus better the electrical conductor

**Summary of some physical properties of the 1st three alkaline earth metals**

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Alkaline earth metal | Appearanc  e | Ease of cuttin g | Meltin g point (oC) | Boilin g point (oC) | Condu  ct- ivity | 1st  ionizatio n energy | 2nd ionizatio n energy |
| Beryllium | Shiny grey | Not easy | 1280 | 3450 | Good | 900 | 1800 |
| Magnesiu m | Shiny grey | Not Easy | 650 | 1110 | Good | 736 | 1450 |
| calcium | Shiny grey | Not easy | 850 | 1140 | Good | 590 | 970 |

Chemical properties

**(i)Reaction with air/oxygen**

On exposure to air, the surface of alkaline earth metals is slowly oxidized to its oxide on prolonged exposure to air.

Example

On exposure to air, the surface of magnesium ribbon is oxidized to form a thin film of Magnesium oxide

. 2Mg(s) + O2(g) -> 2MgO(s)

**(ii)Burning in air/oxygen**

Experiment

Hold a about 2cm length of Magnesium ribbon on a Bunsen flame. Stop heating when it catches fire/start burning. **Caution:** Do not look directly at the flame

Put the products of burning into 100cm3 beaker. Add about 5cm3 of distilled water. Swirl. Test the mixture using litmus papers.

Repeat with Calcium

Observations

-Magnesium burns with a bright blindening flame

-White solid /ash produced

-Solid dissolves in water to form a colourless solution

-Blue litmus paper remain blue

-Red litmus paper turns blue

-colourless gas with pungent smell of urine Explanation

Magnesium burns in air with a bright blindening flame to form a mixture of Magnesium oxide and Magnesium nitride.

2Mg (s) + O2(g) -> 2MgO(s)

3Mg (s) + N2 (g) -> Mg3N2 (s)

Magnesium oxide dissolves in water to form magnesium hydroxide.

MgO(s) + H2O (l) -> Mg(OH)2(aq)

Magnesium nitride dissolves in water to form magnesium hydroxide and produce ammonia gas.

Mg3N2 (s) + 6H2O(l) -> 3Mg(OH)2(aq) + 2NH3 (g)

Magnesium hydroxide and ammonia are weakly alkaline with pH 8/9/10/11 and turns red litmus paper blue.

Calcium burns in air with faint orange/red flame to form a mixture of both Calcium oxide and calcium nitride.

2Ca (s) + O2(g) -> 2CaO(s)

3Ca (s) + N2 (g) -> Ca3N2 (s)

Calcium oxide dissolves in water to form calcium hydroxide.

CaO(s) + H2O(l) -> Ca(OH)2(aq)

Calcium nitride dissolves in water to form calcium hydroxide and produce ammonia gas.

Ca3N2 (s) + 6H2O(l) -> 3Ca(OH)2(aq) + 2NH3 (g)

Calcium hydroxide is also weakly alkaline solution with pH 8/9/10/11 and turns red litmus paper blue.

**(iii)Reaction with water**

Experiment

Measure 50 cm3 of distilled water into a beaker.

Scrub/polish with sand paper 1cm length of Magnesium ribbon

Place it in the water. Test the product-mixture with blue and red litmus papers. Repeat with Calcium metal.

Observations

-Surface of magnesium covered by bubbles of colourless gas.

-Colourless solution formed.

-Effervescence/bubbles/fizzing takes place in Calcium.

-Red litmus paper turns blue.

-Blue litmus paper remains blue.

Explanations

Magnesium slowly reacts with cold water to form Magnesium hydroxide and bubbles of Hydrogen gas that stick on the surface of the ribbon.

Mg(s) + 2H2O (l) -> Mg(OH)2(aq) + H2 (g)

Calcium moderately reacts with cold water to form Calcium hydroxide and produce a steady stream of Hydrogen gas.

Ca(s) + 2H2O (l) -> Ca(OH)2(aq) + H2 (g)

**(iv)Reaction with water vapour/steam**

Experiment

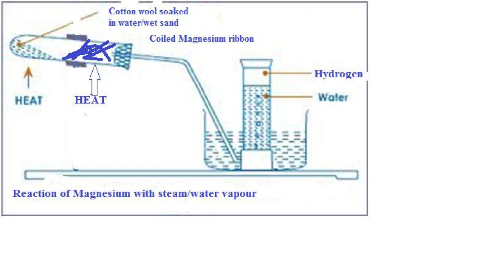
Put some cotton wool soaked in water/wet sand in a long boiling tube.

Coil a well polished magnesium ribbon into the boiling tube.

Ensure the coil touches the side of the boiling tube. Heat the cotton wool/sand slightly then strongly heat the Magnesium ribbon .

Set up of apparatus

Observations



-Magnesium glows red hot then burns with a blindening flame.

-Magnesium continues to glow/burning even without more heating.

-White solid/residue.

-colourless gas collected over water.

Explanation

On heating wet sand, steam is generated which drives out the air that would otherwise react with /oxidize the ribbon.

Magnesium burns in steam/water vapour generating enough heat that ensures the reaction goes to completion even without further heating. White Magnesium oxide is formed and hydrogen gas is evolved.

To prevent suck back, the delivery tube should be removed from the water before heating is stopped at the end of the experiment. Mg(s) + H2O (l) -> MgO(s) + H2 (g)

**(v)Reaction with chlorine gas.**

Experiment

Lower slowly a burning magnesium ribbon/shavings into a gas jar containing Chlorine gas. Repeat with a hot piece of calcium metal.

Observation

-Magnesium continues to burn in chlorine with a bright blindening flame.

-Calcium continues to burn for a short time.

-White solid formed .

-Pale green colour of chlorine fades.

Explanation

Magnesium continues to burn in chlorine gas forming white magnesium oxide solid.

Mg(s) + Cl2 (g) -> MgCl2 (s)

Calcium burns slightly in chlorine gas to form white calcium oxide solid.

Calcium oxide formed coat unreacted Calcium stopping further reaction

Ca(s) + Cl2 (g) -> CaCl2 (s)

**(v)Reaction with dilute acids.**

Experiment

Place about 4.0cm3 of 0.1M dilute sulphuric(VI)acid into a test tube. Add about 1.0cm length of magnesium ribbon into the test tube. Cover the mouth of the test tube using a thumb. Release the gas and test the gas using a burning splint.

Repeat with about 4.0cm3 of 0.1M dilute hydrochloric/nitric(V) acid.

Repeat with 0.1g of Calcium in a beaker with all the above acid

**Caution:** Keep distance when using calcium

Observation

-Effervescence/fizzing/bubbles with dilute sulphuric(VI) and nitric(V) acids -Little Effervescence/fizzing/bubbles with calcium and dilute sulphuric(VI) acid.

-Colourless gas produced that extinguishes a burning splint with an explosion/

“pop” sound.

-No gas is produced with Nitric(V)acid.

-Colourless solution is formed.

Explanation

Dilute acids react with alkaline earth metals to form a salt and produce hydrogen gas.

Nitric(V)acid is a strong oxidizing agent. It quickly oxidizes the hydrogen produced to water.

Calcium is very reactive with dilute acids and thus a very **small** piece of very **dilute** acid should be used.

Chemical equations

|  |  |
| --- | --- |
| Mg(s) + H2SO4 (aq) | -> MgSO4(aq) + H2 (g) |
| Mg(s) + 2HNO3 (aq) | -> Mg(NO3)2(aq) + H2 (g) |
| Mg(s) + 2HCl (aq) | -> MgCl2(aq) + H2 (g) |
| Ca(s) + H2SO4 (aq) | -> CaSO4(**s**) + H2 (g) |
|  | (insoluble CaSO4(**s**) coat/cover Ca(s)) |
| Ca(s) + 2HNO3 (aq) | -> Ca(NO3)2(aq) + H2 (g) |
| Ca(s) + 2HCl (aq) | -> CaCl2(aq) + H2 (g) |
| Ba(s) + H2SO4 (aq) | -> BaSO4(**s**) + H2 (g) |
|  | (insoluble BaSO4(**s**) coat/cover Ba(s)) |
| Ba(s) + 2HNO3 (aq) | -> Ba(NO3)2(aq) + H2 (g) |
| Ba(s) + 2HCl (aq) | -> BaCl2(aq) + H2 (g) |

The table below shows some compounds of some alkaline earth metals

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Beryllium | Magnesium | Calcium | Barium |
| Hydroxide | Be(OH)2 | Mg(OH)2 | Ca(OH)2 | Ba(OH)2 |
| Oxide | BeO | MgO | CaO | BaO |
| Sulphide | - | MgS | CaS | BaS |
| Chloride | BeCl2 | MgCl2 | CaCl2 | BaCl2 |
| Carbonate | BeCO3 | MgCO3 | CaCO3 | BaCO3 |
| Nitrate(V) | Be(NO3)2 | Mg(NO3)2 | Ca(NO3)2 | Ba(NO3)2 |
| Sulphate(VI) | BeSO4 | MgSO4 | CaSO4 | BaSO4 |
| Sulphate(IV) | - | - | CaSO3 | BaSO3 |
| Hydrogen carbonate | - | Mg(HCO3)2 | Ca(HCO3)2 | - |
| Hydrogen sulphate(VI) | - | Mg(HSO4)2 | Ca(HSO4)2 | - |

Some uses of alkaline earth metals include:

(i)Magnesium hydroxide is a non-toxic/poisonous mild base used as an anti acid medicine to relieve stomach acidity.

(ii)Making **duralumin.** Duralumin is an alloy of Magnesium and aluminium used for making aeroplane bodies because it is light.

(iii) Making plaster of Paris-Calcium sulphate(VI) is used in hospitals to set a fractures bone.

(iii)Making cement-Calcium carbonate is mixed with clay and sand then heated to form cement for construction/building.

(iv)Raise soil pH-Quicklime/calcium oxide is added to acidic soils to neutralize and raise the soil pH in agricultural farms.

(v)As nitrogenous fertilizer-Calcium nitrate(V) is used as an agricultural fertilizer because plants require calcium for proper growth.

(vi)In the blast furnace-Limestone is added to the blast furnace to produce more reducing agent and remove slag in the blast furnace for extraction of Iron.

(c)Group VII elements: Halogens

Group VII elements are called **Halogens.** They are all non metals. They include:

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Element** | **Symbol** | **Atomic number** | **Electronicc configuration** | **Charge**  **of ion** | **Valency** | **State at Room Temperature** |
| **Fluorine**  **Chlorine**  **Bromine**  **Iodine**  **Astatine** | **F**  **Cl**  **Br**  **I**  **At** | **9**  **17**  **35**  **53**  **85** | **2:7**  **2:8:7**  **2:8:18:7**  **2:8:18:18:7**  **2:8:18:32:18:7** | **F - Cl- Br-**  **I-**  **At-** | **1**  **1**  **1**  **1**  **1** | **Pale yellow gas**  **Pale green gas**  **Red liquid**  **Grey Solid**  **Radioactive** |

All halogen atoms have seven electrons in the outer energy level. They **acquire/gain one** electron in the outer energy level to be stable. They therefore are therefore **monovalent** .They exist in oxidation state X**-**

The number of energy levels increases down the group from Fluorine to Astatine. The more the number of energy levels the bigger/larger the atomic size. e.g.

The atomic size/radius of Chlorine is bigger/larger than that of Fluorine because Chlorine has more/3 energy levels than Fluorine (2 energy levels).

Atomic radius and ionic radius of Halogens increase down the group as the number of energy levels increases.

The atomic radius of Halogens is smaller than the ionic radius. This is because they react by gaining/acquiring extra one electron in the outer energy level. The effective nuclear attraction on the more/extra electrons decreases. The incoming extra electron is also repelled causing the outer energy level to expand to reduce the repulsion and accommodate more electrons.

**Table showing the atomic and ionic radius of four Halogens**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Atomic radius(nM) | Ionic radius(nM) |
| Fluorine | F | 9 | 0.064 | 0.136 |
| Chlorine | Cl | 17 | 0.099 | 0.181 |
| Bromine | Br | 35 | 0.114 | 0.195 |
| Iodine | I | 53 | 0.133 | 0.216 |

The atomic radius of Chlorine is 0.099nM .The ionic radius of Cl- is 0.181nM. This is because Chlorine atom/molecule reacts by gaining/acquiring extra one electrons. The more/extra electrons/energy level experience less effective nuclear attraction /pull towards the nucleus .The outer enegy level expand/increase to reduce the repulsion of the existing and incoming gained /acquired electrons.

Electronegativity

The ease of gaining/acquiring extra electrons is called electronegativity. All halogens are electronegative. Electronegativity decreases as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius.

The outer electrons experience less nuclear attraction and thus ease of gaining/acquiring extra electrons decrease.

It is measured using Pauling’s scale.

Where Fluorine with Pauling scale 4.0 is the most electronegative element and thus the highest tendency to acquire/gain extra electron.

Table showing the electronegativity of the halogens.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Halogen | F | Cl | Br | I | At |
| Electronegativity (Pauling scale) | 4.0 | 3.0 | 2.8 | 2.5 | 2.2 |

The electronegativity of the halogens decrease down the group from fluorine to Astatine. This is because atomic radius increases down the group and thus decrease electron – attracting power down the group from fluorine to astatine.

Fluorine is the most electronegative element in the periodic table because it has the small atomic radius.

Electron affinity

The minimum amount of energy required to gain/acquire an extra electron by an atom of element in its gaseous state is called **1st electron affinity**. The SI unit of electron affinity is **kilojoules per mole/kJmole-1 .** Electron affinity depend on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer energy level electrons and thus the lower the electron affinity. For halogens the 1st electron affinity decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease. Due to its small size/atomic radius Fluorine shows exceptionally low electron affinity. This is because a lot of energy is required to overcome the high repulsion of the existing and incoming electrons.

Table showing the election affinity of halogens for the process X + e -> X**-**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Halogen | F | Cl | Br | I |
| Electron affinity kJmole -1 | -333 | -364 | -342 | -295 |

The higher the electron affinity the more stable theion.i.e

Cl- is a more stable ion than Br- because it has a more negative / exothermic electron affinity than Br- Electron affinity is different from:

1. Ionization energy.

Ionization energy is the energy required to lose/donate an electron in an atom of an element in its gaseous state while electron affinity is the energy required to gain/acquire extra electron by an atom of an element in its gaseous state.

1. Electronegativity.

-Electron affinity is the energy required to gain an electron in an atom of an element in gaseous state. It involves the process:

X(g) + e -> X-(g)

Electronegativity is the ease/tendency of gaining/ acquiring electrons by an element during chemical reactions.

It does not involve use of energy but theoretical arbitrary Pauling’ scale of measurements.

Physical properties

State at room temperature

Fluorine and Chlorine are gases, Bromine is a liquid and Iodine is a solid.

Astatine is radioactive .

All halogens exist as **diatomic** molecules bonded by strong covalent bond. Each molecule is joined to the other by weak intermolecular forces/ Van-der-waals forces.

Melting/Boiling point

The strength of intermolecular/Van-der-waals forces of attraction increase with increase in molecular size/atomic radius.

Iodine has therefore the largest atomic radius and thus strongest intermolecular forces to make it a solid.

Iodine sublimes when heated to form (**caution:** highly toxic/poisonous) purple vapour.

This is because Iodine molecules are held together by weak van-derwaals/intermolecular forces which require little heat energy to break.

Electrical conductivity

All Halogens are poor conductors of electricity because they have no free delocalized electrons.

Solubility in polar and non-polar solvents All halogens are soluble in water(polar solvent).

When a boiling tube containing either chlorine gas or bromine vapour is separately inverted in a beaker containing distilled water and tetrachloromethane (non-polar solvent), the level of solution in boiling tube rises in both water and tetrachloromethane.

This is because halogen are soluble in both polar and non-polar solvents.

Solubility of halogens in water/polar solvents decrease down the group.

Solubility of halogens in non-polar solvent increase down the group. The level of water in chlorine is higher than in bromine and the level of tetrachloromethane in chlorine is lower than in bromine.

**Caution:** Tetrachloromethane , Bromine vapour and Chlorine gas are all **highly** toxic/poisonous.

**Table showing the physical properties of Halogens**

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Haloge n** | **Formul**  **a of**  **molecul**  **e** | **Electrical**  **conductivit**  **y** | **Solubility in water** | **Melting point(oC**  **)** | **Boiling**  **point(oC**  **)** |
| Fluorin e | F2 | Poor | Insoluble/soluble in  tetrachlorometha  ne | -238 | -188 |
| Chlorin e | Cl2 | Poor | Insoluble/soluble in  tetrachlorometha  ne | -101 | -35 |
| Bromin  e | Br2 | Poor | Insoluble/soluble in  tetrachlorometha  ne | 7 | 59 |
| Iodine | I2 | Poor | Insoluble/soluble in  tetrachlorometha  ne | 114 | sublimes |

Chemical properties

**(i)Displacement**

Experiment

Place separately in test tubes about 5cm3 of sodium chloride, Sodium bromide and Sodium iodide solutions.

Add 5 drops of chlorine water to each test tube:

Repeat with 5 drops of bromine water instead of chlorine water

Observation

**Using Chlorine water**

-Yellow colour of chlorine water fades in all test tubes except with sodium chloride.

-Coloured Solution formed.

**Using Bromine water**

Yellow colour of bromine water fades in test tubes containing sodium iodide.

-Coloured Solution formed.

Explanation

The halogens displace each other from their solution. The more electronegative displace the less electronegative from their solution.

Chlorine is more electronegative than bromine and iodine.

On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.

Bromine is more electronegative than iodide but less 6than chlorine.

On adding Bromine water, iodine is displaced from its solution but not chlorine.

**Table showing the displacement of the halogens**

(V) means there is displacement (x ) means there is no displacement

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Halogen ion in    solution  Halogen | F - | Cl - | Br - | I - |
| F2 | X |  |  |  |
| Cl2 | X | X |  |  |
| Br2 | X | X | X |  |
| I2 | X | X | X | X |

|  |  |  |  |
| --- | --- | --- | --- |
| Chemical /ionic equations  With Fluorine |  |  |  |
| F2(g) + 2NaCl-(aq) -> | 2NaF(aq) | + | Cl2(aq) |
| F2(g) + 2Cl-(aq) -> | 2F-(aq) | + | Cl2(aq) |
| F2(g) + 2NaBr-(aq) -> | 2NaF(aq) | + | Br2(aq) |
| F2(g) + 2Br-(aq) -> | 2F-(aq) | + | Br2(aq) |
| F2(g) + 2NaI-(aq) -> | 2NaF(aq) | + | I2(aq) |
| F2(g) + 2I-(aq) ->    With chlorine | 2F-(aq) | + | I2(aq) |
| Cl2(g) + 2NaCl-(aq) -> 2NaCl(aq) + Br2(aq)  Cl2(g) + 2Br-(aq) -> 2Cl-(aq) + Br2(aq)    Cl2(g) + 2NaI-(aq) -> 2NaCl(aq) + I2(aq)  Cl2(g) + 2I-(aq) -> 2Cl-(aq) + I2(aq)    With Bromine  Br2(g) + 2NaI-(aq) -> 2NaBr(aq) + I2(aq)  Br2(g) + 2I-(aq) -> 2Br-(aq) + I2(aq) | | | |

Uses of halogens

1. Florine – manufacture of P.T.F.E (Poly tetra fluoroethene) synthetic fiber.
   * Reduce tooth decay when added in small amounts/quantities in tooth paste.

NB –large small quantities of fluorine /fluoride ions in water cause browning of teeth/flourosis.

* + Hydrogen fluoride is used to engrave words /pictures in glass.

1. Bromine - Silver bromide is used to make light sensitive photographic paper/films.

1. Iodide – Iodine dissolved in alcohol is used as medicine to kill bacteria in skin cuts. It is called tincture of iodine.

The table below to show some compounds of halogens.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Element  Halogen | H | Na | Mg | Al | Si | C | P |
| F | **HF** | **NaF** | **MgH2** | **AlF3** | **SiF4** | **CF4** | **PF3** |
| Cl | **HCl** | **NaCl** | **MgCl** | **AlCl3** | **SiCl3** | **CCl4** | **PCl3** |
| Br | **HBr** | **NaBr** | **MgBr2** | **AlBr3** | **SiBr4** | **CBr4** | **PBr3** |
| I | **Hl** | **Nal** | **Mgl2** | **All3** | **SiI4** | **Cl2** | **PBr3** |

(i) Below is the table showing the bond energy of four halogens.

Bond Bond energy k J mole-1

Cl-Cl 242

Br-Br 193

I-I 151

1. What do you understand by the term “bond energy”

**Bond energy is the energy required to break/ form one mole of chemical bond**

1. Explain the trend in bond Energy of the halogens above:

-**Decrease down the group from chlorine to Iodine**

**-Atomic radius increase down the group decreasing the energy required to break the covalent bonds between the larger atom with reduced effective nuclear @ charge an outer energy level that take part in bonding.**

(c)Group VIII elements: Noble gases

Group VIII elements are called **Noble gases.** They are all non metals. Noble gases occupy about 1.0% of the atmosphere as colourless gaseous mixture.

Argon is the most abundant with 0.9%.

They exists as **monatomic** molecules with very weak van-der-waals /intermolecular forces holding the molecules.

They include:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Electron structure | State at room temperature |
| Helium | He | 2 | 2: | Colourless gas |
| Neon | Ne | 10 | 2:8 | Colourless gas |
| Argon | Ar | 18 | 2:8:8 | Colourless gas |
| Krypton | Kr | 36 | 2:8:18:8 | Colourless gas |
| Xenon | Xe | 54 | 2:8:18:18:8 | Colourless gas |
| Radon | Rn | 86 | 2:8:18:32:18:8 | Radioctive |

All noble gas atoms have a stable duplet(two electrons in the 1st energy level) or octet(eight electrons in other outer energy level)in the outer energy level. They therefore **do not** acquire/gain extra electron in the outer energy level or donate/lose. They therefore are therefore **zerovalent** .

The number of energy levels increases down the group from Helium to Randon. The more the number of energy levels the bigger/larger the atomic size/radius. e.g.

The atomic size/radius of Argon is bigger/larger than that of Neon because Argon has more/3 energy levels than Neon (2 energy levels).

Atomic radius noble gases increase down the group as the number of energy levels increases.

The effective nuclear attraction on the outer electrons thus decrease down the group.

The noble gases are generally unreactive because the outer energy level has the stable octet/duplet. The stable octet/duplet in noble gas atoms lead to a comparatively very high 1st ionization energy. This is because losing /donating an electron from the stable atom require a lot of energy to lose/donate and make it unstable.

As atomic radius increase down the group and the 1st ionization energy decrease, very electronegative elements like Oxygen and Fluorine are able to react and bond with lower members of the noble gases.e.g

Xenon reacts with Fluorine to form a covalent compound XeF6.This is because the outer electrons/energy level if Xenon is far from the nucleus and thus experience less effective nuclear attraction.

Noble gases have low melting and boiling points. This is because they exist as monatomic molecules joined by very weak intermolecular/van-der-waals forces that require very little energy to weaken and form liquid and break to form a gas.

The intermolecular/van-der-waals forces increase down the group as the atomic radius/size increase from Helium to Radon. The melting and boiling points thus increase also down the group.

Noble gases are insoluble in water and are poor conductors of electricity.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Element | Formula of molecule | Electrical conductivity | Solubility in water | Atomic radius(nM) | 1st  ionization energy | Melting point(0C) | Boiling point(0C) |
| Helium | He | Poor | Insoluble | 0.128 | 2372 | -270 | -269 |
| Neon | Ne | Poor | Insoluble | 0.160 | 2080 | -249 | -246 |
| Argon | Ar | Poor | Insoluble | 0.192 | 1520 | -189 | -186 |
| Krypton | Kr | Poor | Insoluble | 0.197 | 1350 | -157 | -152 |
| Xenon | Xe | Poor | Insoluble | 0.217 | 1170 | -112 | -108 |
| Radon | Rn | Poor | Insoluble | 0.221 | 1134 | -104 | -93 |

Uses of noble gases

Argon is used in light bulbs to provide an inert environment to prevent oxidation of the bulb filament

Argon is used in arch welding as an insulator.

Neon is used in street and advertisement light

Helium is mixed with Oxygen during deep sea diving and mountaineering.

Helium is used in weather balloon for meteorological research instead of Hydrogen because it is unreactive/inert.Hydrogen when impure can ignite with an explosion.

Helium is used in making thermometers for measuring very low temperatures.

**C. PERIODICITY OF ACROSS THE PERIOD.**

**(See Chemical bonding and Structure)**

**CHEMICAL BONDING AND STRUCTURE**

**A. CHEMICAL BONDING**

A chemical bond is formed when atoms of the same or different elements **share**, **gain**, **donate or delocalize** their **outer** energy level **electrons** to combine during **chemical reactions** inorder to be **stable**.

Atoms have equal number of negatively charged electrons in the energy levels and positively charged protons in the nucleus.

Atoms are chemically stable if they have filled outer energy level. An energy level is full if it has duplet (2) or octet (8) state in outer energy level.

Noble gases have duplet /octet. All other atoms try to be like noble gases through chemical reactions and forming molecules.

Only electrons in the outer energy level take part in formation of a chemical bond. There are three main types of chemical bonds formed by atoms:

1. covalent bond
2. ionic/electrovalent bond
3. metallic bond

(i)**COVALENT BOND**

A covalent bond is formed when atoms of the same or different element share some or all the outer energy level electrons to combine during chemical reactions inorder to attain duplet or octet.

A shared pair of electrons is attracted by the nucleus (protons) of the two atoms sharing.

Covalent bonds are mainly formed by non-metals to form molecules. A molecule is a group of atoms of the same or different elements held together by a covalent bond. The number of atoms making a molecule is called **atomicity**. Noble gases are **monatomic** because they are stable and thus do not bond with each other or other atoms. Most other gases are **diatomic**

The more the number of electrons shared, the stronger the covalent bond.

A pair of electrons that do not take part in the formation of a covalent bond is called a **lone pair of electrons.**

Mathematically, the number of electrons to be shared by an atom is equal to the number of electrons remaining for the atom to be stable/attain duplet/octet /have maximum electrons in outer energy level.

**The following diagrams illustrate the formation of covalent bonds:**

a)hydrogen molecule is made up of two hydrogen atoms in the outer energy level each requiring one electron to have a stable duplet.

To show the formation of covalent bonding in the molecule then the following data/information is required;

Symbol of atom/element taking part in bonding H H

Number of protons/electrons 1 1

Electron configuration/structure 1: 1:

Number of electron in outer energy level 1 1

Number of electrons remaining to be stable/shared 1 1

Number of electrons not shared(lone pairs) 0 0

**Diagram method 1**

**Diagram method 2**

**H H**

**●x**

**Note:**

After bonding the following **intramolecular** forces exist:

(i)the attraction of the shared electrons by both nucleus /protons of the atoms (ii) the repulsion of the nucleus of one atom on the other.

(iii)balance of the attraction and repulsion is maintained inside/intramolecular/within the molecule as follows;

E1

P1 P1

E1

(iv)Protons(P1) from nucleus of atom 1 **repel** protons (P2) from nucleus of atom 2.

(v)Electron (E1) in the energy levels of atom 1 **repel** electron (E2) in the energy levels of atom 2.

1. Protons(P1) from nucleus of atom 1 **attract** electron (E2) in the energy levels of atom 2.
2. protons (P2) from nucleus of atom 2 **attract** electron (E2) in the energy levels of atom 2.

1. Fluorine, chlorine, bromine and iodine molecules are made up also of two atoms sharing the outer energy level electrons to have a stable octet.

To show the formation of covalent bonding in the molecule then the following data/information is required;

* 1. fluorine

Symbol of atom/element taking part in bonding F F

Number of protons/electrons 9 9

Electron configuration/structure 2:7 2:7

Number of electron in outer energy level 7 7

Number of electrons remaining to be stable/shared 1 1

Number of outer electrons not shared( 3-lone pairs) 6 6

**Diagram method 1**

**Diagram method 2**

* 1. chlorine

Symbol of atom/element taking part in bonding Cl Cl

Number of protons/electrons 17 17

Electron configuration/structure 2:8:7 2:8:7

Number of electron in outer energy level 7 7

Number of electrons remaining to be stable/shared 1 1

Number of outer electrons not shared( 3-lone pairs) 6 6

**Diagram method 1**

**Diagram method 2**

* 1. Bromine

Symbol of atom/element taking part in bonding Br Br

Number of protons/electrons 35 35

Electron configuration/structure 2:8:18:7 2:8:18:7

Number of electron in outer energy level 7 7

Number of electrons remaining to be stable/shared 1 1

Number of outer electrons not shared( 3-lone pairs) 6 6

**Diagram method 1**

**Diagram method 2**

* 1. Iodine

Symbol of atom/element taking part in bonding I I

Number of protons/electrons 53 53

Electron configuration/structure 2:8:18:18:7 2:8:18:18:7

Number of electron in outer energy level 7 7

Number of electrons remaining to be stable/shared 1 1

Number of outer electrons not shared( 3-lone pairs) 6 6

**Diagram method 1**

**Diagram method 2**

1. Oxygen molecule is made up of two atoms sharing each two outer energy level electrons to have a stable octet as shown below;

Symbol of atom/element taking part in bonding O O

Number of protons/electrons 8 8

Electron configuration/structure 2:6 2:6

Number of electron in outer energy level 6 6

Number of electrons remaining to be stable/shared 2 2

Number of outer electrons not shared( 2-lone pairs) 4 4

**Diagram method 1**

**Diagram method 2**

1. Nitrogen and phosphorus molecule is made up of two atoms sharing each three outer energy level electrons to have a stable octet as shown below;
   1. Nitrogen

Symbol of atom/element taking part in bonding N N

Number of protons/electrons 7 7

Electron configuration/structure 2:5 2:5

Number of electron in outer energy level 5 5

Number of electrons remaining to be stable/shared 3 3

Number of outer electrons not shared ( 3-lone pairs) 2 2

**Diagram method 1**

**Diagram method 2**

* 1. Phosphorus

Symbol of atom/element taking part in bonding P P

Number of protons/electrons 15 15

Electron configuration/structure 2:8:5 2:8:5

Number of electron in outer energy level 5 5

Number of electrons remaining to be stable/shared 3 3

Number of outer electrons not shared ( 3-lone pairs) 2 2

**Diagram method 1**

**Diagram method 2**

1. Water molecule is made up of hydrogen and oxygen. Hydrogen requires to share one electron with oxygen to be stable/attain duplet. Oxygen requires to share two electrons to be stable/attain octet. Two hydrogen atoms share with one oxygen atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding O H

Number of protons/electrons 8 1

Electron configuration/structure 2:6 1

Number of electron in outer energy level 6 1

Number of electrons remaining to be stable/shared 2 1

Number of electrons not shared( 2-Oxygen lone pairs) 4 0

**Diagram method 1**

**Diagram method 2**

1. Ammonia molecule is made up of Hydrogen and Nitrogen. Hydrogen requires to share one electron with Nitrogen to be stable/attain duplet. Nitrogen requires to share three electrons to be stable/attain octet. Three hydrogen atoms share with one nitrogen atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding N H

Number of protons/electrons 7 1

Electron configuration/structure 2:5 1:

Number of electron in outer energy level 5 1

Number of electrons remaining to be stable/shared 3 1

Number of electrons not shared( 1-Nitrogen lone pairs) 2 0

**Diagram method 1**

**Diagram method 2**

g)Carbon(IV) oxide molecule is made up of carbon and oxygen. Carbon requires to share four electrons with oxygen to be stable/attain octet. Oxygen requires to share two electrons to be stable/attain octet. Two oxygen atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding O C

Number of protons/electrons 8 6

Electron configuration/structure 2:6 2:4

Number of electron in outer energy level 6 4

Number of electrons remaining to be stable/shared 2 4

2-lone pairs from each Oxygen atom) 2 0

**Diagram method 1**

**Diagram method 2**

1. Methane molecule is made up of hydrogen and carbon. Hydrogen requires sharing one electron with carbon to be stable/attain duplet. Carbon requires sharing four electrons to be stable/attain octet. Four hydrogen atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding C H

Number of protons/electrons 6 1

Electron configuration/structure 2:4 1

Number of electron in outer energy level 4 1

Number of electrons remaining to be stable/shared 4 1

Number of electrons not shared ( No lone pairs) 0 0

**Diagram method 1**

**Diagram method 2**

1. Tetrachloromethane molecule is made up of chlorine and carbon. Chlorine requires sharing one electron with carbon to be stable/attain octet. Carbon requires sharing four electrons to be stable/attain octet. Four chlorine atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding C Cl

Number of protons/electrons 6 17

Electron configuration/structure 2:4 2:8:7

Number of electron in outer energy level 4 7

Number of electrons remaining to be stable/shared 4 1

3-lone pairs from each Chlorine atom(24 electrons) 0 6

**Diagram method 1**

**Diagram method 2**

1. Ethane molecule is made up of six hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Three hydrogen atoms share with one carbon atom while another three hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing a pair of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding C H

Number of protons/electrons 6 1

Electron configuration/structure 2:4 1

Number of electron in outer energy level 4 1

Number of electrons remaining to be stable/shared 4 1

Number of electrons not shared( No lone pairs) 0 0

**Diagram method 1**

**Diagram method 2**

1. Ethene molecule is made up of four hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Two hydrogen atoms share with one carbon atom while another two hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing two pairs of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding C H

Number of protons/electrons 6 1

Electron configuration/structure 2:4 1

Number of electron in outer energy level 4 1

Number of electrons remaining to be stable/shared 4 1

Number of electrons not shared( No lone pairs) 0 0

**Diagram method 1**

**Diagram method 2**

1. Ethyne molecule is made up of two hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. One hydrogen atoms share with one carbon atom while another hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing three pairs of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding C H

Number of protons/electrons 6 1

Electron configuration/structure 2:4 1

Number of electron in outer energy level 4 1

Number of electrons remaining to be stable/shared 4 1

Number of electrons not shared( No lone pairs) 0 0

**Diagram method 1**

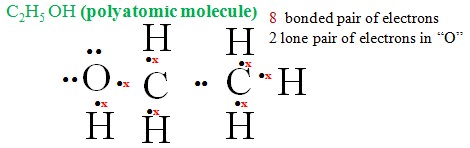
**Diagram method 2**

j) Ethanol molecule is made up of six hydrogen one Oxygen atom two carbon atoms.

Five Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet

Each Carbon uses four electrons to share with ―O‖and ―H‖attain octet/duplet.

NB: Oxygen has two lone pairs



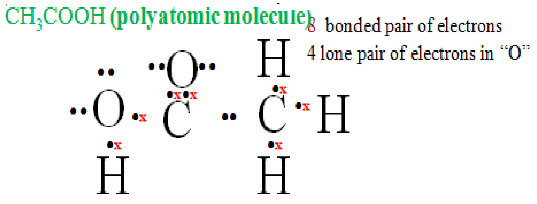
**j)Ethanoic molecule is made up of four hydrogen two Oxygen atom two carbon atoms.**

**Three Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet**

**Each Carbon uses four electrons to share with “O”and “H”attain octet/duplet.**

**NB: Each Oxygen atom has two lone pa**

**irs**



By convention (as a rule), a

1. **single** covalent bond made up of two shared( **a pair**) electrons is represented by a dash(---)
2. **double** covalent bond made up of four shared( **two pairs)** electrons is represented by a double dash(==)
3. **triple** covalent bond made up of six shared( **three pairs)** electrons is represented by a triple dash(==)

The representation below show the molecules covered in (a) to (k) above:

1. Hydrogen molecule(H2) H--H

1. Fluorine molecule(F2) F--F

1. Chlorine molecule(Cl2) Cl--Cl

1. Bromine molecule(Br2) Br--Br

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| e) Iodine molecule(I2) |  | I--I |  |  |
| f) Oxygen molecule(O2) |  | O=O |  |  |
| g) Nitrogen molecule(N2) |  | N=N |  |  |
| h) Phosphorus molecule(P2) |  | P=P |  |  |
| i) Water molecule (H2O) |  | H--O--H |  |  |
| j Ammonia molecule(NH3) |  | H--N--H |  |  |
|  |  |  |  |  |
|  |  | H |  |  |
| k)Carbon(IV) oxide molecule(CO2) |  | O==C==O |  |  |
|  |  | H |  |  |
| l)Methane molecule(CH4) |  | H--C--H |  |  |
|  |  | H |  |  |
|  |  |  |  |  |
|  |  | Cl |  |  |
| m)Tetrachloromethane molecule(CCl4 |  | Cl--C--Cl |  |  |
|  |  | Cl |  |  |
|  |  | H H |  |  |
| n)Ethane molecule(C2H6) |  | H--C—C--H |  |  |

H H

p)Ethene molecule(C2H4) H-C==C-H

H H

q)Ethyne molecule(C2H6) H-C—C-H

Dative /coordinate bond

A dative/coordinate bond is a covalent bond formed when a lone pair of electrons is donated then shared to an electron-deficient species/ion/atom.

During dative/coordinate bonding, all the shared pair of electrons are donated by one of the combining/bonding species/ ion/atom.

Like covalent bonding, coordinate /dative bond is mainly formed by non-metals.

**Illustration of coordinate /dative bond**

a)Ammonium ion(NH4+)

The ammonium ion is made up of ammonia (NH3) molecule and hydrogen (H+) ion. (H+) ion has no electrons. NH3 is made up of covalent bonding from Nitrogen and Hydrogen. One lone pair of electrons is present in Nitrogen atom after the bonding. This lone pair is donated and shared with an electron-deficient H+ ion Diagram method 1

Diagram method 2

b)Phosphine ion (PH4+)

The Phosphine ion is made up of phosphine(NH3) molecule and hydrogen (H+) ion. (H+) ion has no electrons. PH3 is made up of covalent bonding from Phosphorus and Hydrogen. One lone pair of electrons is present in Phosphorus atom. After the bonding this lone pair is donated and shared with the electron-deficient H+ ion Diagram method 1

Diagram method 2

1. Hydroxonium (H3O+) ion

The hydroxonium ion is made up of water (H2O) molecule and hydrogen (H+) ion. (H+) ion has no electrons. The H2O molecule is made up of covalent bonding from Oxygen and Hydrogen. One lone pair of electrons out of the two present in Oxygen atom after the bonding is donated and shared with the electron-deficient H+ ion Diagram method 1

Diagram method 2

1. Carbon (II) oxide (CO)

Carbon (II) oxide is made up of carbon and Oxygen atoms sharing each two outer electron and not sharing each two electrons. Oxygen with an extra lone pair of electrons donates and share with the carbon atom for both to be stable.

Diagram method 1

Diagram method 2

1. Aluminium (III) chloride (AlCl3/Al2Cl6)

Aluminium (III) chloride is made up of aluminium and chlorine. One aluminium atom shares its outer electrons with three separate chlorine atoms. All chlorine atoms attain stable octet but aluminium does not. Another molecule of aluminium chloride shares its chlorine lone pair of electrons with the aluminium atom for both to be stable. This type of bond exists only in vapour phase after aluminium chloride sublimes.

Diagram method 1

Diagram method 2

A dative/coordinate bond is by convention represented by an arrow (→) heading from the donor of the shared pair of electrons.

Below is the representation of molecules in the above examples;

a)Ammonium ion.

|  |  |  |
| --- | --- | --- |
|  | H |  |
|  | H− N→H |  |
|  | H |  |
| b)Phosphine ion | H |  |
|  | H− P→H |  |
|  | H |  |
| c)Hydroxonium ion |  |  |
|  | H− O→H |  |
|  | H |  |
| d)Carbon(II) oxide | O→C |  |
| d) Aluminium(III)chloride Cl Cl | | Cl |

Al Al

Cl Cl Cl

(ii)**IONIC/ELECTROVALENT BOND**

An ionic/electrovalent bond is **extreme** of a covalent bond.

During ionic/electrovalent bonding there is complete **transfer** of valence electrons to one electronegative atom from an electropositive atom.

All metals are electropositive and easily/readily donate/lose their valence electrons.

All non-metals are electronegative and easily/readily gain/acquire extra electrons.

Ionic/electrovalent bonding therefore mainly involves transfer of electrons from metal/metallic radical to non-metallic radical.

When an electropositive atom donates /loses the valence electrons, it forms a positively charged **cation** to attain stable octet/duplet.

When an electronegative atom gains /acquires extra valence electrons, it forms a negatively charged **anion** to attain stable octet/duplet**.**

The **electrostatic attraction force** between the **stable** positively charged cation and the stable negatively charged anion with **opposite** charges constitute the ionic bond.

Like in covalent/dative/coordinate bonding, only the outer energy level electrons take part in the formation of ionic/electrovalent bond

Like in covalent/dative/coordinate bonding, the **more** electrons taking part / involved in the formation of ionic/electrovalent bond, the **stronger** the ionic /electrovalent bond.

**Illustration of ionic /electrovalent bond**

a)Sodium chloride(NaCl)

Sodium chloride(NaCl) is formed when a sodium atom donate its outer valence electrons to chlorine atom for both to attain stable octet:

|  |  |
| --- | --- |
| Symbol of atoms/elements taking part in bonding | Na Cl |
| Number of protons/electrons | 11 17 |
| Electron configuration/structure | 2:8:1 2:8:7 |
| Number of electron in outer energy level | 11 7 |
| Number of electrons donated and gained to be stable | 1 1 |
| New electron configuration/structure | 2:8: 2:8: |
| Symbol of cation/anion after bonding | Na+ Cl- |

**Diagram**

b)Magnesium chloride(MgCl2)

Magnesium chloride (MgCl2) is formed when a magnesium atom donate its two outer valence electrons to chlorine atoms. Two chlorine atoms are required to gain each one electron. All the ions (cations and anions) attain stable octet:

|  |  |
| --- | --- |
| Symbol of atoms/elements taking part in bonding | Mg Cl |
| Number of protons/electrons | 11 17 |
| Electron configuration/structure | 2:8:2 2:8:7 |
| Number of electron in outer energy level | 2 7 |
| Number of electrons donated and gained to be stable | 2 1 |
| New electron configuration/structure | 2:8: 2:8: |
| Symbol of cation/anion after bonding | Mg2+ Cl- |

**Diagram**

c)Lithium oxide(Li2O)

Lithium oxide(Li2O)is formed when a Lithium atom donate its outer valence electrons to Oxygen atom. Two Lithium atoms are required to donate/lose each one electron and attain stable duplet. Oxygen atom acquires the two electrons and attain stable octet:

Symbol of atoms/elements taking part in bonding Li O

Number of protons/electrons 3 8

Electron configuration/structure 2:1 2:6

Number of electron in outer energy level 1 6

Number of electrons donated and gained to be stable 1 2

New electron configuration/structure 2: 2:8:

Symbol of cation/anion after bonding Li+ O2-

**Diagram**

d)Aluminium(III) oxide(Al2O3)

Aluminium(III) oxide(Al2O3)is formed when a Aluminium atom donate its three outer valence electrons to Oxygen atom. Two Aluminium atoms are required to donate/lose each three electron and attain stable octet. Three Oxygen atoms gain/ acquire the six electrons and attain stable octet:

|  |  |
| --- | --- |
| Symbol of atoms/elements taking part in bonding | Al O |
| Number of protons/electrons | 13 8 |
| Electron configuration/structure | 2:8:3 2:6 |
| Number of electron in outer energy level | 3 6 |
| Number of electrons donated and gained to be stable | 3 2 |
| New electron configuration/structure | 2:8: 2:8: |
| Symbol of cation/anion after bonding | Al3+ O2- |

**Diagram**

e)Calcium oxide(CaO)

Calcium oxide(CaO)is formed when a Calcium atom donate its two outer valence electrons to Oxygen atom. Both attain stable octet:

Symbol of atoms/elements taking part in bonding Ca O Number of protons/electrons 20 8

Electron configuration/structure 2:8:8:2 2:6

Number of electron in outer energy level 2 6

Number of electrons donated and gained to be stable 2 2

New electron configuration/structure 2:8:8: 2:8:

Symbol of cation/anion after bonding Ca2+ O2-

**Diagram**

Some compounds can be formed from ionic/electrovalent, covalent and dative/coordinate bonding within their atoms/molecules:

a)Formation of ammonium chloride:

Ammonium chloride is formed from the reaction of ammonia gas and hydrogen chloride gas. Both ammonia and hydrogen chloride gas are formed from covalent bonding. During the reaction of ammonia and hydrogen chloride gas to form Ammonium chloride;

-ammonia forms a dative/coordinate bond with electron deficient H+ ion from Hydrogen chloride to form ammonium ion(NH4+)ion.

-the chloride ion Cl- and ammonium ion(NH4+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

b) Dissolution/dissolving of hydrogen chloride:

Hydrogen chloride is formed when hydrogen and chlorine atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When hydrogen chloride gas is dissolved in water;

-water molecules forms a dative/coordinate bond with electron deficient H+ ion from Hydrogen chloride to form hydroxonium ion(H3O+)ion.

-the chloride ion Cl- and hydroxonium ion(H3O+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

.

c)Dissolution/dissolving of ammonia gas:

Ammmonia gas is formed when hydrogen and Nitrogen atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When Ammonia gas is dissolved in water;

-ammonia forms a dative/coordinate bond with electron deficient H+ ion from a water molecule to form ammonium ion(NH4+)ion.

-the hydroxide ion OH- and ammonium ion(NH4+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

(iii)**METALLIC BOND**

A metallic bond is formed when metallic atoms **delocalize** their outer electrons inorder to be stable.

Metals delocalize their outer electrons to form positively charged cation .

The electrostatic attraction force between the metallic cation and the negatively charged electrons constitute the metallic bond.

The more delocalized electrons the stonger the metallic bond.

**Illustration of ionic /electrovalent bond**

1. Sodium (Na) is made of one valence electron. The electron is donated to form Na+ ion. The electron is delocalized /free within many sodium ions.

Symbol of atoms/elements taking part in bonding Na Na Na

|  |  |
| --- | --- |
| Number of protons/electrons | 11 11 11 |
| Electron configuration/structure | 2:8:1 2:8:1 2:8:1 |
| Number of electron in outer energy level | 1 1 1 |
| Number of electrons delocalized/free within | 1 1 1 |
| New electron configuration/structure | 2:8: 2:8: 2:8: |
| Symbol of cation after metallic bonding | Na+ Na+ Na+ |

**Diagram**

(three)Metallic cations **attract**

(three) free/delocalized electrons

1. Aluminium (Al) is made of three valence electron. The three electrons are donated to form Al3+ ion. The electrons are delocalized /free within many aluminium ions.

|  |  |
| --- | --- |
| Symbol of atoms/elements taking part in bonding Al Al Al | |
| Number of protons/electrons | 13 13 13 |
| Electron configuration/structure | 2:8:3 2:8:3 2:8:3 |
| Number of electron in outer energy level | 3 3 3 |
| Number of electrons delocalized/free within | 3 3 3 |
| New electron configuration/structure | 2:8: 2:8: 2:8: |
| Symbol of cation after metallic bonding | Al3+ Al3+ Al3+ |

**Diagram**

(three)Metallic cations **attract**

(nine) free/delocalized electrons

c)Calcium (Ca) is made of two valence electron.The two electrons are donated to form Ca2+ ion.The electrons are delocalized /free within many Calcium ions.

Symbol of atoms/elements taking part in bonding Ca Ca Ca

Number of protons/electrons 20 20 20

Electron configuration/structure 2:8:8:2 2:8:8:2 2:8:8:2

|  |  |
| --- | --- |
| Number of electron in outer energy level | 2 2 2 |
| Number of electrons delocalized/free within | 2 2 2 |
| New electron configuration/structure | 2:8:8: 2:8:8: 2:8:8: |
| Symbol of cation after metallic bonding | Ca2+ Ca2+ Ca2+ |

**Diagram**

(three)Metallic cations **attract**

(six) free/delocalized electrons

d) Magnesium (Mg) is made of two valence electron. The two electrons are donated to form Mg2+ion.The electrons are delocalized /free within many Magnesium ions.

|  |  |  |  |
| --- | --- | --- | --- |
| Symbol of atoms/elements taking part in bonding | | Mg | Mg |
| Number of protons/electrons |  | 12 | 12 |
| Electron configuration/structure |  | 2:8:2 | 2:8:2 |
| Number of electron in outer energy level |  | 2 | 2 |
| Number of electrons delocalized/free within |  | 2 | 2 |
| New electron configuration/structure |  | 2:8: | 2:8: |
| Symbol of cation after metallic bonding |  | Mg2+ Mg2+ | |

**Diagram**

(two)Metallic cations **attract**

(four) free/delocalized electrons

e)Lithium (Li) is made of one valence electron.The electron is donated to form Li+ ion.The electron is delocalized /free within many Lithium ions.ie;

|  |  |  |
| --- | --- | --- |
| Symbol of atoms/elements taking part in bonding Li | | Li Li Li |
| Number of protons/electrons | 3 | 3 3 3 |
| Electron configuration/structure | 2:1 | 2:1 2:1 2:1 |
| Number of electron in outer energy level | 1 | 1 1 1 |
| Number of electrons delocalized/free within 1 | 1 | 1 1 |
| New electron configuration/structure | 2:1: | 2:1: 2:1: 2:1: |
| Symbol of cation after metallic bonding | Li+ | Li+ Li+  Li+ |

**Diagram**

(four)Metallic cations **attract**

(four) free/delocalized electrons

**CHEMICAL FAMILES**

**(Patterns down the group)**

The number of valence electrons and the number of occupied energy levels in an atom of an element determine the position of an element in the periodic table. i.e The number of occupied energy levels determine the Period and the valence electrons determine the Group.

Elements in the same group have similar physical and chemical properties. The trends in physical and chemical properties of elements in the same group vary down the group. Elements in the same group thus constitute a chemical family.

(a) Group I elements: Alkali metals

Group I elements are called **Alkali metals** except Hydrogen which is a non metal. The alkali metals include:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Electron structure | Oxidation state | Valency |
| Lithium | Li | 3 | 2:**1** | Li**+** | 1 |
| Sodium | Na | 11 | 2:8:**1** | Na**+** | 1 |
| Potassium | K | 19 | 2:8:8:**1** | K**+** | 1 |
| Rubidium | Rb | 37 | 2:8:18:8:**1** | Rb**+** | 1 |
| Caesium | Cs | 55 | 2:8:18:18:8:**1** | Cs**+** | 1 |
| Francium | Fr | 87 | 2:8:18:32:18:8:**1** | Fr**+** | 1 |

All alkali metals atom has one electron in the outer energy level. They therefore are **monovalent**. They donate /lose the outer electron to have oxidation state M+ The number of energy levels increases down the group from Lithium to Francium. The more the number of energy levels the bigger/larger the atomic size. e.g. The atomic size of Potassium is bigger/larger than that of sodium because Potassium has more/4 energy levels than sodium (3 energy levels).

Atomic and ionic radius

The distance between the centre of the nucleus of an **atom** and the outermost energy level occupied by electron/s is called **atomic radius.** Atomic radius is measured in **nanometers**(n).The higher /bigger the atomic radius the bigger /larger the atomic size.

The distance between the centre of the nucleus of an **ion** and the outermost energy level occupied by electron/s is called **ionic radius.** Ionic radius is also measured in **nanometers**(n).The higher /bigger the ionic radius the bigger /larger the size of the ion.

Atomic radius and ionic radius depend on the number of energy levels occupied by electrons. The more the number of energy levels the bigger/larger the atomic /ionic radius. e.g.

The atomic radius of Francium is bigger/larger than that of sodium because Francium has more/7 energy levels than sodium (3 energy levels).

Atomic radius and ionic radius of alkali metals increase down the group as the number of energy levels increases.

The atomic radius of alkali metals is bigger than the ionic radius. This is because alkali metals react by losing/donating the outer electron and hence lose the outer energy level.

**Table showing the atomic and ionic radius of some alkali metals**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Atomic radius(nM) | Ionic radius(nM) |
| Lithium | Li | 3 | 0.133 | 0.060 |
| Sodium | Na | 11 | 0.157 | 0.095 |
| Potassium | K | 19 | 0.203 | 0.133 |

The atomic radius of sodium is 0.157nM .The ionic radius of Na+ is 0.095nM. This is because sodium reacts by donating/losing the outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the atomic radius.

Electropositivity

The ease of donating/losing electrons is called electropositivity. All alkali metals are electropositive. Electropositivity increase as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The outer electrons experience less nuclear attraction and can be lost/ donated easily/with ease. Francium is the most electropositive element in the periodic table because it has the highest/biggest atomic radius.

Ionization energy

The minimum amount of energy required to remove an electron from an atom of element in its gaseous state is called **1st ionization energy**. The SI unit of ionization energy is **kilojoules per mole/kJmole-1 .**Ionization energy depend on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer electrons/energy level and thus the lower the ionization energy. For alkali metals the 1st ionization energy decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease.

e.g. The 1st ionization energy of sodium is 496 kJmole-1while that of potassium is 419 kJmole-1 **.**This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from sodium to Potassium. It requires therefore less energy to donate/lose outer electrons in Potassium than in sodium.

Physical properties

**Soft/Easy to cut**: Alkali metals are soft and easy to cut with a knife. The softness and ease of cutting increase down the group from Lithium to Francium. This is because an increase in atomic radius, decreases the strength of metallic bond and the packing of the metallic structure

**Appearance:** Alkali metals have a shiny grey metallic luster when freshly cut. The surface rapidly/quickly tarnishes on exposure to air. This is because the metal surface rapidly/quickly reacts with elements of air/oxygen.

**Melting and boiling points:** Alkali metals have a relatively low melting/boiling point than common metals like Iron. This is because alkali metals use only one delocalized electron to form a weak metallic bond/structure.

**Electrical/thermal conductivity:** Alkali metals are good thermal and electrical conductors. Metals conduct using the outer mobile delocalized electrons. The delocalized electrons move randomly within the metallic structure.

**Summary of some physical properties of the 1st three alkali metals**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Alkali metal | Appearance | Ease of cutting | Melting point (oC) | Boiling point (oC) | Conductivity | 1st  ionization energy |
| Lithium | Silvery white | Not easy | 180 | 1330 | Good | 520 |
| Sodium | Shiny grey | Easy | 98 | 890 | Good | 496 |
| Potassium | Shiny grey | Very easy | 64 | 774 | Good | 419 |

Chemical properties

**(i)Reaction with air/oxygen**

On exposure to air, alkali metals reacts with the elements in the air. Example

On exposure to air, Sodium first reacts with Oxygen to form sodium oxide.

4Na(s) + O2(g) -> 2Na2O(s)

The sodium oxide formed further reacts with water/moisture in the air to form sodium hydroxide solution.

Na2O(s) + H2O(l) -> 2NaOH(aq)

Sodium hydroxide solution reacts with carbon(IV)oxide in the air to form sodium carbonate.

2NaOH(aq) + CO2(g) -> Na2CO3(g) + H2O(l)

**(ii)Burning in air/oxygen**

Lithium burns in air with a **crimson**/deep red flame to form Lithium oxide

4Li (s) + O2(g) -> 2Li2O(s)

Sodium burns in air with a **yellow** flame to form sodium oxide

4Na (s) + O2(g) -> 2Na2O(s)

Sodium burns in oxygen with a **yellow** flame to form sodium peroxide

2Na (s) + O2(g) -> Na2O2 (s)

Potassium burns in air with a **lilac/purple** flame to form potassium oxide

4K (s) + O2(g) -> 2K2O (s)

**(iii) Reaction with water:**

Experiment

Measure 500 cm3 of water into a beaker.

Put three drops of phenolphthalein indicator.

Put about 0.5g of Lithium metal into the beaker.

Determine the pH of final product

Repeat the experiment using about 0.1 g of Sodium and Potassium. **Caution**: Keep a distance

Observations

|  |  |  |
| --- | --- | --- |
| Alkali metal | Observations | Comparative speed/rate of the reaction |
| Lithium | -Metal floats in water  -rapid effervescence/fizzing/bubbling -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn  phenolphthalein indicator pink  -pH of solution = 12/13/14 | Moderately vigorous |
| Sodium | -Metal floats in water  -very rapid effervescence /fizzing  /bubbling  -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn  phenolphthalein indicator pink  -pH of solution = 12/13/14 | Very vigorous |
| Potassium | -Metal floats in water  -explosive effervescence /fizzing  /bubbling  -colourless gas produced (that extinguishes burning splint with explosion /“pop” sound) -resulting solution turn  phenolphthalein indicator pink  -pH of solution = 12/13/14 | Explosive/burst into flames |

Explanation

Alkali metals are less dense than water. They therefore float in water.They react with water to form a strongly alkaline solution of their hydroxides and producing hydrogen gas. The rate of this reaction increase down the group. i.e. Potassium is more reactive than sodium .Sodium is more reactive than Lithium.

The reactivity increases as electropositivity increases of the alkali increases. This is because as the atomic radius increases , the ease of donating/losing outer electron increase during chemical reactions.

Chemical equations

2Li(s) + 2H2O(l) -> 2LiOH(aq) + H2(g)

2Na(s) + 2H2O(l) -> 2NaOH(aq) + H2(g)

2K(s) + 2H2O(l) -> 2KOH(aq) + H2(g)

2Rb(s) + 2H2O(l) -> 2RbOH(aq) + H2(g)

2Cs(s) + 2H2O(l) -> 2CsOH(aq) + H2(g)

2Fr(s) + 2H2O(l) -> 2FrOH(aq) + H2(g)

Reactivity increase down the group

**(iv) Reaction with chlorine:**

Experiment

Cut about 0.5g of sodium into a deflagrating spoon with a lid cover. Introduce it on a Bunsen flame until it catches fire. Quickly and carefully lower it into a gas jar containing dry chlorine to cover the gas jar.

Repeat with about 0.5g of Lithium.

**Caution:** This experiment should be done in fume chamber because chlorine is poisonous /toxic.

Observation

Sodium metal continues to burn with a yellow flame forming white solid/fumes.

Lithium metal continues to burn with a crimson flame forming white solid / fumes. Alkali metal react with chlorine gas to form the corresponding metal chlorides. The reactivity increase as electropositivity increase down the group from Lithium to Francium.The ease of donating/losing the outer electrons increase as the atomic radius increase and the outer electron is less attracted to the nucleus.

Chemical equations

2Li(s) + Cl2(g) -> 2LiCl(s)

2Na(s) + Cl2(g) -> 2NaCl(s)

2K(s) + Cl2(g) -> 2KCl(s)

2Rb(s) + Cl2(g) -> 2RbCl(s)

2Cs(s) + Cl2(g) -> 2CsCl(s)

2Fr(s) + Cl2(g) -> 2FrCl(s)

Reactivity increase down the group

The table below shows some compounds of the 1st three alkali metals

|  |  |  |  |
| --- | --- | --- | --- |
|  | Lithium | sodium | Potassium |
| Hydroxide | LiOH | NaOH | KOH |
| Oxide | Li2O | Na2O | K2O |
| Sulphide | Li2S | Na2S | K2S |
| Chloride | LiCl | NaCl | KCl |
| Carbonate | Li2CO3 | Na2CO3 | K2CO3 |
| Nitrate(V) | LiNO3 | NaNO3 | KNO3 |
| Nitrate(III) | - | NaNO2 | KNO2 |
| Sulphate(VI) | Li2SO4 | Na2SO4 | K2SO4 |
| Sulphate(IV) | - | Na2SO3 | K2SO3 |
| Hydrogen carbonate | - | NaHCO3 | KHCO3 |
| Hydrogen sulphate(VI) | - | NaHSO4 | KHSO4 |
| Hydrogen sulphate(IV) | - | NaHSO3 | KHSO3 |
| Phosphate | - | Na3PO4 | K3PO4 |
| Manganate(VI) | - | NaMnO4 | KMnO4 |
| Dichromate(VI) | - | Na2Cr2O7 | K2Cr2O7 |
| Chromate(VI) | - | Na2CrO4 | K2CrO4 |

Some **uses** of alkali metals include:

(i)Sodium is used in making sodium cyanide for extracting gold from gold ore.

(ii)Sodium chloride is used in seasoning food.

(iii)Molten mixture of sodium and potassium is used as coolant in nuclear reactors. (iv)Sodium is used in making sodium hydroxide used in making soapy and soapless detergents.

(v)Sodium is used as a reducing agent for the extraction of titanium from Titanium(IV)chloride.

(vi)Lithium is used in making special high strength glasses

(vii)Lithium compounds are used to make dry cells in mobile phones and computer laptops.

Group II elements: Alkaline earth metals

Group II elements are called **Alkaline earth metals** . The alkaline earth metals include:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Electron structure | Oxidation state | Valency |
| Beryllium | Be | 4 | 2:**2** | Be**2+** | 2 |
| Magnesium | Mg | 12 | 2:8:**2** | Mg**2+** | 2 |
| Calcium | Ca | 20 | 2:8:8:**2** | Ca**2+** | 2 |
| Strontium | Sr | 38 | 2:8:18:8:**2** | Sr**2+** | 2 |
| Barium | Ba | 56 | 2:8:18:18:8:**2** | Ba**2+** | 2 |
| Radium | Ra | 88 | 2:8:18:32:18:8:**2** | Ra**2+** | 2 |

All alkaline earth metal atoms have two electrons in the outer energy level. They therefore are **divalent**. They donate /lose the two outer electrons to have oxidation state M2+

The number of energy levels increases down the group from Beryllium to Radium. The more the number of energy levels the bigger/larger the atomic size. e.g.

The atomic size/radius of Calcium is bigger/larger than that of Magnesium because Calcium has more/4 energy levels than Magnesium (3 energy levels).

Atomic radius and ionic radius of alkaline earth metals increase down the group as the number of energy levels increases.

The atomic radius of alkaline earth metals is bigger than the ionic radius. This is because they react by losing/donating the two outer electrons and hence lose the outer energy level.

**Table showing the atomic and ionic radius of the 1st three alkaline earth metals**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic | Atomic radius(nM) | Ionic radius(nM) |
|  |  | number |  |  |
| Beryllium | Be | 4 | 0.089 | 0.031 |
| Magnesium | Mg | 12 | 0.136 | 0.065 |
| Calcium | Ca | 20 | 0.174 | 0.099 |

The atomic radius of Magnesium is 0.136nM .The ionic radius of Mg2+ is 0.065nM. This is because Magnesium reacts by donating/losing the two outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the atomic radius.

Electropositivity

All alkaline earth metals are also electropositive like alkali metals. The electropositivity increase with increase in atomic radius/size. Calcium is more electropositive than Magnesium. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The two outer electrons in calcium experience less nuclear attraction and can be lost/ donated easily/with ease because of the higher/bigger atomic radius.

Ionization energy

For alkaline earth metals the 1st ionization energy decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease.

e.g. The 1st ionization energy of Magnesium is 900 kJmole-1while that of Calcium is 590 kJmole-1 **.**This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from magnesium to calcium.

It requires therefore less energy to donate/lose outer electron in calcium than in magnesium.

The minimum amount of energy required to remove a second electron from an ion of an element in its gaseous state is called the **2nd ionization energy**.

The 2nd ionization energy is always higher /bigger than the 1st ionization energy. This because once an electron is donated /lost form an atom, the overall effective nuclear attraction on the remaining electrons/energy level increase. Removing a second electron from the ion require therefore more energy than the first electron. The atomic radius of alkali metals is higher/bigger than that of alkaline earth metals.This is because across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level. Increase in nuclear charge increases the effective nuclear attraction on the outer energy level which pulls it closer to the nucleus. e.g.

Atomic radius of Sodium (0.157nM) is higher than that of Magnesium (0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus.

Physical properties

**Soft/Easy to cut**: Alkaline earth metals are **not** soft and easy to cut with a knife like alkali metals. This is because of the decrease in atomic radius of corresponding alkaline earth metal, increases the strength of metallic bond and the packing of the metallic structure. Alkaline earth metals are

(i)ductile(able to form wire/thin long rods)

(ii)malleable(able to be hammered into sheet/long thin plates)

(iii)have high tensile strength(able to be coiled without breaking/ not brittle/withstand stress)

**Appearance:** Alkali earth metals have a shiny grey metallic luster when their surface is freshly polished /scrubbed. The surface slowly tarnishes on exposure to air. This is because the metal surface slowly undergoes oxidation to form an oxide. This oxide layer should be removed before using the alkaline earth metals.

**Melting and boiling points:** Alkaline earth metals have a relatively high melting/ boiling point than alkali metals. This is because alkali metals use only one delocalized electron to form a weaker metallic bond/structure. Alkaline earth metals use two delocalized electrons to form a stronger metallic bond /structure. Themelting and boiling points decrease down the group as the atomic radius/size increase reducing the strength of metallic bond and packing of the metallic structure. e.g.

Beryllium has a melting point of 1280oC. Magnesium has a melting point of 650oC.Beryllium has a smaller atomic radius/size than magnesium .The strength of metallic bond and packing of the metallic structure is thus stronger in beryllium.

**Electrical/thermal conductivity:** Alkaline earth metals are good thermal and electrical conductors. The two delocalized valence electrons move randomly within the metallic structure.

Electrical conductivity increase down the group as the atomic radius/size increase making the delocalized outer electrons less attracted to nucleus. Alkaline earth metals are better thermal and electrical conductors than alkali metals because they have more/two outer delocalized electrons.e.g.

Magnesium is a better conductor than sodium because it has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

Calcium is a better conductor than magnesium.

Calcium has bigger/larger atomic radius than magnesium because the delocalized electrons are less attracted to the nucleus of calcium and thus more free /mobile and thus better the electrical conductor

**Summary of some physical properties of the 1st three alkaline earth metals**

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Alkaline earth metal | Appearance | Ease of cutting | Melting point (oC) | Boiling point (oC) | Conduct  - ivity | 1st  ionization energy | 2nd ionization energy |
| Beryllium | Shiny grey | Not easy | 1280 | 3450 | Good | 900 | 1800 |
| Magnesium | Shiny grey | Not Easy | 650 | 1110 | Good | 736 | 1450 |
| calcium | Shiny grey | Not easy | 850 | 1140 | Good | 590 | 970 |

Chemical properties

**(i)Reaction with air/oxygen**

On exposure to air, the surface of alkaline earth metals is slowly oxidized to its oxide on prolonged exposure to air.

Example

On exposure to air, the surface of magnesium ribbon is oxidized to form a thin film of Magnesium oxide

. 2Mg(s) + O2(g) -> 2MgO(s)

**(ii)Burning in air/oxygen**

Experiment

Hold a about 2cm length of Magnesium ribbon on a Bunsen flame. Stop heating when it catches fire/start burning. **Caution:** Do not look directly at the flame

Put the products of burning into 100cm3 beaker. Add about 5cm3 of distilled water. Swirl. Test the mixture using litmus papers.

Repeat with Calcium

Observations

-Magnesium burns with a bright blindening flame

-White solid /ash produced

-Solid dissolves in water to form a colourless solution

-Blue litmus paper remain blue

-Red litmus paper turns blue

-colourless gas with pungent smell of urine

Explanation

Magnesium burns in air with a bright blindening flame to form a mixture of Magnesium oxide and Magnesium nitride.

2Mg (s) + O2(g) -> 2MgO(s)

3Mg (s) + N2 (g) -> Mg3N2 (s)

Magnesium oxide dissolves in water to form magnesium hydroxide.

MgO(s) + H2O (l) -> Mg(OH)2(aq)

Magnesium nitride dissolves in water to form magnesium hydroxide and produce ammonia gas.

Mg3N2 (s) + 6H2O(l) -> 3Mg(OH)2(aq) + 2NH3 (g)

Magnesium hydroxide and ammonia are weakly alkaline with pH 8/9/10/11 and turns red litmus paper blue.

Calcium burns in air with faint orange/red flame to form a mixture of both Calcium oxide and calcium nitride.

2Ca (s) + O2(g) -> 2CaO(s)

3Ca (s) + N2 (g) -> Ca3N2 (s)

Calcium oxide dissolves in water to form calcium hydroxide.

CaO(s) + H2O(l) -> Ca(OH)2(aq)

Calcium nitride dissolves in water to form calcium hydroxide and produce ammonia gas.

Ca3N2 (s) + 6H2O(l) -> 3Ca(OH)2(aq) + 2NH3 (g)

Calcium hydroxide is also weakly alkaline solution with pH 8/9/10/11 and turns red litmus paper blue.

**(iii)Reaction with water**

Experiment

Measure 50 cm3 of distilled water into a beaker.

Scrub/polish with sand paper 1cm length of Magnesium ribbon

Place it in the water. Test the product-mixture with blue and red litmus papers. Repeat with Calcium metal.

Observations

-Surface of magnesium covered by bubbles of colourless gas.

-Colourless solution formed.

-Effervescence/bubbles/fizzing takes place in Calcium.

-Red litmus paper turns blue.

-Blue litmus paper remains blue.

Explanations

Magnesium slowly reacts with cold water to form Magnesium hydroxide and bubbles of Hydrogen gas that stick on the surface of the ribbon.

Mg(s) + 2H2O (l) -> Mg(OH)2(aq) + H2 (g)

Calcium moderately reacts with cold water to form Calcium hydroxide and produce a steady stream of Hydrogen gas.

Ca(s) + 2H2O (l) -> Ca(OH)2(aq) + H2 (g)

**(iv)Reaction with water vapour/steam**

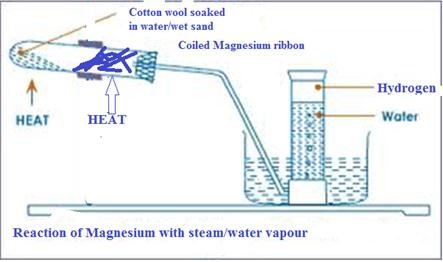
Experiment

Put some cotton wool soaked in water/wet sand in a long boiling tube.

Coil a well polished magnesium ribbon into the boiling tube.

Ensure the coil touches the side of the boiling tube. Heat the cotton wool/sand slightly then strongly heat the Magnesium ribbon .

Set up of apparatus



Observations

-Magnesium glows red hot then burns with a blindening flame.

-Magnesium continues to glow/burning even without more heating.

-White solid/residue.

-colourless gas collected over water.

Explanation

On heating wet sand, steam is generated which drives out the air that would otherwise react with /oxidize the ribbon.

Magnesium burns in steam/water vapour generating enough heat that ensures the reaction goes to completion even without further heating. White Magnesium oxide is formed and hydrogen gas is evolved.

To prevent suck back, the delivery tube should be removed from the water before heating is stopped at the end of the experiment.

Mg(s) + H2O (l) -> MgO(s) + H2 (g)

**(v)Reaction with chlorine gas.**

Experiment

Lower slowly a burning magnesium ribbon/shavings into a gas jar containing Chlorine gas. Repeat with a hot piece of calcium metal.

Observation

-Magnesium continues to burn in chlorine with a bright blindening flame.

-Calcium continues to burn for a short time.

-White solid formed .

-Pale green colour of chlorine fades.

Explanation

Magnesium continues to burn in chlorine gas forming white magnesium oxide solid.

Mg(s) + Cl2 (g) -> MgCl2 (s)

Calcium burns slightly in chlorine gas to form white calcium oxide solid. Calcium oxide formed coat unreacted Calcium stopping further reaction Ca(s) + Cl2 (g) -> CaCl2 (s)

**(v)Reaction with dilute acids.**

Experiment

Place about 4.0cm3 of 0.1M dilute sulphuric(VI)acid into a test tube. Add about 1.0cm length of magnesium ribbon into the test tube. Cover the mouth of the test tube using a thumb. Release the gas and test the gas using a burning splint.

Repeat with about 4.0cm3 of 0.1M dilute hydrochloric/nitric(V) acid.

Repeat with 0.1g of Calcium in a beaker with all the above acid **Caution:** Keep distance when using calcium

Observation

-Effervescence/fizzing/bubbles with dilute sulphuric(VI) and nitric(V) acids -Little Effervescence/fizzing/bubbles with calcium and dilute sulphuric(VI) acid. -Colourless gas produced that extinguishes a burning splint with an explosion/

“pop” sound.

-No gas is produced with Nitric(V)acid.

-Colourless solution is formed.

Explanation

Dilute acids react with alkaline earth metals to form a salt and produce hydrogen gas.

Nitric(V)acid is a strong oxidizing agent. It quickly oxidizes the hydrogen produced to water.

Calcium is very reactive with dilute acids and thus a very **small** piece of very **dilute** acid should be used. Chemical equations

|  |  |
| --- | --- |
| Mg(s) + H2SO4 (aq) | -> MgSO4(aq) + H2 (g) |
| Mg(s) + 2HNO3 (aq) | -> Mg(NO3)2(aq) + H2 (g) |
| Mg(s) + 2HCl (aq) | -> MgCl2(aq) + H2 (g) |
| Ca(s) + H2SO4 (aq) | -> CaSO4(**s**) + H2 (g) |
|  | (insoluble CaSO4(**s**) coat/cover Ca(s)) | |
| Ca(s) + 2HNO3 (aq) | -> Ca(NO3)2(aq) + H2 (g) | |
| Ca(s) + 2HCl (aq) | -> CaCl2(aq) + H2 (g) | |
| Ba(s) + H2SO4 (aq) | -> BaSO4(**s**) + H2 (g) | |
|  | (insoluble BaSO4(**s**) coat/cover Ba(s)) | |
| Ba(s) + 2HNO3 (aq) | -> Ba(NO3)2(aq) + H2 (g) | |
| Ba(s) + 2HCl (aq) | -> BaCl2(aq) + H2 (g) | |

The table below shows some compounds of some alkaline earth metals

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Beryllium | Magnesium | Calcium | Barium |
| Hydroxide | Be(OH)2 | Mg(OH)2 | Ca(OH)2 | Ba(OH)2 |
| Oxide | BeO | MgO | CaO | BaO |
| Sulphide | - | MgS | CaS | BaS |
| Chloride | BeCl2 | MgCl2 | CaCl2 | BaCl2 |
| Carbonate | BeCO3 | MgCO3 | CaCO3 | BaCO3 |
| Nitrate(V) | Be(NO3)2 | Mg(NO3)2 | Ca(NO3)2 | Ba(NO3)2 |
| Sulphate(VI) | BeSO4 | MgSO4 | CaSO4 | BaSO4 |
| Sulphate(IV) | - | - | CaSO3 | BaSO3 |
| Hydrogen carbonate | - | Mg(HCO3)2 | Ca(HCO3)2 | - |
| Hydrogen sulphate(VI) | - | Mg(HSO4)2 | Ca(HSO4)2 | - |

Some uses of alkaline earth metals include:

(i)Magnesium hydroxide is a non-toxic/poisonous mild base used as an anti acid medicine to relieve stomach acidity.

(ii)Making **duralumin.** Duralumin is an alloy of Magnesium and aluminium used for making aeroplane bodies because it is light.

(iii) Making plaster of Paris-Calcium sulphate(VI) is used in hospitals to set a fractures bone.

(iii)Making cement-Calcium carbonate is mixed with clay and sand then heated to form cement for construction/building.

(iv)Raise soil pH-Quicklime/calcium oxide is added to acidic soils to neutralize and raise the soil pH in agricultural farms.

(v)As nitrogenous fertilizer-Calcium nitrate(V) is used as an agricultural fertilizer because plants require calcium for proper growth.

(vi)In the blast furnace-Limestone is added to the blast furnace to produce more reducing agent and remove slag in the blast furnace for extraction of Iron.

(c)Group VII elements: Halogens

Group VII elements are called **Halogens.** They are all non metals. They include:

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Element** | **Symbol** | **Atomic number** | **Electronicc configuration** | **Charge**  **of ion** | **Valency** | **State at Room Temperature** |
| **Fluorine**  **Chlorine**  **Bromine**  **Iodine**  **Astatine** | **F**  **Cl**  **Br**  **I**  **At** | **9**  **17**  **35**  **53**  **85** | **2:7**  **2:8:7**  **2:8:18:7**  **2:8:18:18:7**  **2:8:18:32:18:7** | **F - Cl- Br-**  **I-**  **At-** | **1**  **1**  **1**  **1**  **1** | **Pale yellow gas**  **Pale green gas**  **Red liquid**  **Grey Solid**  **Radioactive** |

All halogen atoms have seven electrons in the outer energy level. They **acquire/gain one** electron in the outer energy level to be stable. They therefore are therefore **monovalent** .They exist in oxidation state X**-**

The number of energy levels increases down the group from Fluorine to Astatine.

The more the number of energy levels the bigger/larger the atomic size. e.g. The atomic size/radius of Chlorine is bigger/larger than that of Fluorine because Chlorine has more/3 energy levels than Fluorine (2 energy levels).

Atomic radius and ionic radius of Halogens increase down the group as the number of energy levels increases.

The atomic radius of Halogens is smaller than the ionic radius. This is because they react by gaining/acquiring extra one electron in the outer energy level. The effective nuclear attraction on the more/extra electrons decreases. The incoming extra electron is also repelled causing the outer energy level to expand to reduce the repulsion and accommodate more electrons.

**Table showing the atomic and ionic radius of four Halogens**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Atomic radius(nM) | Ionic radius(nM) |
| Fluorine | F | 9 | 0.064 | 0.136 |
| Chlorine | Cl | 17 | 0.099 | 0.181 |
| Bromine | Br | 35 | 0.114 | 0.195 |
| Iodine | I | 53 | 0.133 | 0.216 |

The atomic radius of Chlorine is 0.099nM .The ionic radius of Cl- is 0.181nM. This is because Chlorine atom/molecule reacts by gaining/acquiring extra one electrons. The more/extra electrons/energy level experience less effective nuclear attraction /pull towards the nucleus .The outer enegy level expand/increase to reduce the repulsion of the existing and incoming gained /acquired electrons.

Electronegativity

The ease of gaining/acquiring extra electrons is called electronegativity. All halogens are electronegative. Electronegativity decreases as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius.

The outer electrons experience less nuclear attraction and thus ease of gaining/acquiring extra electrons decrease.

It is measured using Pauling’s scale.

Where Fluorine with Pauling scale 4.0 is the most electronegative element and thus the highest tendency to acquire/gain extra electron.

Table showing the electronegativity of the halogens.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Halogen | F | Cl | Br | I | At |
| Electronegativity (Pauling scale) | 4.0 | 3.0 | 2.8 | 2.5 | 2.2 |

The electronegativity of the halogens decrease down the group from fluorine to Astatine. This is because atomic radius increases down the group and thus decrease electron – attracting power down the group from fluorine to astatine.

Fluorine is the most electronegative element in the periodic table because it has the small atomic radius.

Electron affinity

The minimum amount of energy required to gain/acquire an extra electron by an atom of element in its gaseous state is called **1st electron affinity**. The SI unit of electron affinity is **kilojoules per mole/kJmole-1 .** Electron affinity depend on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer energy level electrons and thus the lower the electron affinity. For halogens the 1st electron affinity decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease. Due to its small size/atomic radius Fluorine shows exceptionally low electron affinity. This is because a lot of energy is required to overcome the high repulsion of the existing and incoming electrons.

Table showing the election affinity of halogens for the process X + e -> X**-**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Halogen | F | Cl | Br | I |
| Electron affinity kJmole -1 | -333 | -364 | -342 | -295 |

The higher the electron affinity the more stable theion.i.e

Cl- is a more stable ion than Br- because it has a more negative / exothermic electron affinity than Br- Electron affinity is different from:

1. Ionization energy.

Ionization energy is the energy required to lose/donate an electron in an atom of an element in its gaseous state while electron affinity is the energy required to gain/acquire extra electron by an atom of an element in its gaseous state.

1. Electronegativity.

-Electron affinity is the energy required to gain an electron in an atom of an element in gaseous state. It involves the process:

X(g) + e -> X-(g)

Electronegativity is the ease/tendency of gaining/ acquiring electrons by an element during chemical reactions.

It does not involve use of energy but theoretical arbitrary Pauling’ scale of measurements.

Physical properties

State at room temperature

Fluorine and Chlorine are gases, Bromine is a liquid and Iodine is a solid.

Astatine is radioactive .

All halogens exist as **diatomic** molecules bonded by strong covalent bond. Each molecule is joined to the other by weak intermolecular forces/ Van-der-waals forces.

Melting/Boiling point

The strength of intermolecular/Van-der-waals forces of attraction increase with increase in molecular size/atomic radius.

Iodine has therefore the largest atomic radius and thus strongest intermolecular forces to make it a solid.

Iodine sublimes when heated to form (**caution:** highly toxic/poisonous) purple vapour.

This is because Iodine molecules are held together by weak van-derwaals/intermolecular forces which require little heat energy to break.

Electrical conductivity

All Halogens are poor conductors of electricity because they have no free delocalized electrons.

Solubility in polar and non-polar solvents

All halogens are soluble in water(polar solvent).

When a boiling tube containing either chlorine gas or bromine vapour is separately inverted in a beaker containing distilled water and tetrachloromethane (non-polar solvent), the level of solution in boiling tube rises in both water and tetrachloromethane.

This is because halogen are soluble in both polar and non-polar solvents. Solubility of halogens in water/polar solvents decrease down the group. Solubility of halogens in non-polar solvent increase down the group.

The level of water in chlorine is higher than in bromine and the level of tetrachloromethane in chlorine is lower than in bromine.

**Caution:** Tetrachloromethane , Bromine vapour and Chlorine gas are all **highly** toxic/poisonous.

**Table showing the physical properties of Halogens**

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Halogen** | **Formula of molecule** | **Electrical conductivity** | **Solubility in water** | **Melting point(oC)** | **Boiling point(oC)** |
| Fluorine | F2 | Poor | Insoluble/soluble in | -238 | -188 |
|  |  |  | tetrachloromethane |  |  |
| Chlorine | Cl2 | Poor | Insoluble/soluble in  tetrachloromethane | -101 | -35 |
| Bromine | Br2 | Poor | Insoluble/soluble in  tetrachloromethane | 7 | 59 |
| Iodine | I2 | Poor | Insoluble/soluble in  tetrachloromethane | 114 | sublimes |

Chemical properties

**(i)Displacement**

Experiment

Place separately in test tubes about 5cm3 of sodium chloride, Sodium bromide and Sodium iodide solutions.

Add 5 drops of chlorine water to each test tube:

Repeat with 5 drops of bromine water instead of chlorine water

Observation

**Using Chlorine water**

-Yellow colour of chlorine water fades in all test tubes except with sodium chloride.

-Coloured Solution formed.

**Using Bromine water**

Yellow colour of bromine water fades in test tubes containing sodium iodide.

-Coloured Solution formed.

Explanation

The halogens displace each other from their solution. The more electronegative displace the less electronegative from their solution.

Chlorine is more electronegative than bromine and iodine.

On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.

Bromine is more electronegative than iodide but less 6than chlorine.

On adding Bromine water, iodine is displaced from its solution but not chlorine.

**Table showing the displacement of the halogens**

(V) means there is displacement (x ) means there is no displacement

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Halogen ion in    solution  Halogen | F - | Cl - | Br - | I - |
| F2 | X |  |  |  |
| Cl2 | X | X |  |  |
| Br2 | X | X | X |  |
| I2 | X | X | X | X |

Chemical /ionic equations

With Fluorine

F2(g) + 2NaCl-(aq) -> 2NaF(aq) + Cl2(aq)

F2(g) + 2Cl-(aq) -> 2F-(aq) + Cl2(aq)

F2(g) + 2NaBr-(aq) -> 2NaF(aq) + Br2(aq)

F2(g) + 2Br-(aq) -> 2F-(aq) + Br2(aq)

F2(g) + 2NaI-(aq) -> 2NaF(aq) + I2(aq)

F2(g) + 2I-(aq) -> 2F-(aq) + I2(aq)

With chlorine

Cl2(g) + 2NaCl-(aq) -> 2NaCl(aq) + Br2(aq)

Cl2(g) + 2Br-(aq) -> 2Cl-(aq) + Br2(aq)

Cl2(g) + 2NaI-(aq) -> 2NaCl(aq) + I2(aq)

Cl2(g) + 2I-(aq) -> 2Cl-(aq) + I2(aq)

With Bromine

Br2(g) + 2NaI-(aq) -> 2NaBr(aq) + I2(aq)

Br2(g) + 2I-(aq) -> 2Br-(aq) + I2(aq)

Uses of halogens

1. Florine – manufacture of P.T.F.E (Poly tetra fluoroethene) synthetic fiber.
   * Reduce tooth decay when added in small amounts/quantities in tooth paste.

NB –large small quantities of fluorine /fluoride ions in water cause browning of teeth/flourosis.

* + Hydrogen fluoride is used to engrave words /pictures in glass.

1. Bromine - Silver bromide is used to make light sensitive photographic paper/films.

1. Iodide – Iodine dissolved in alcohol is used as medicine to kill bacteria in skin cuts. It is called tincture of iodine.

The table below to show some compounds of halogens.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Element  Halogen | H | Na | Mg | Al | Si | C | P |
| F | **HF** | **NaF** | **MgF2** | **AlF3** | **SiF4** | **CF4** | **PF3** |
| Cl | **HCl** | **NaCl** | **MgCl2** | **AlCl**  **3** | **SiCl 4** | **CCl4** | **PCl3** |
| Br | **HBr** | **NaBr** | **MgBr2** | **AlBr**  **3** | **SiBr4** | **CBr4** | **PBr3** |
| I | **Hl** | **Nal** | **Mgl2** | **All3** | **SiI4** | **C l 4** | **PBr3** |

(i) Below is the table showing the bond energy of four halogens.

Bond Bond energy k J mole-1

Cl-Cl 242

Br-Br 193

I-I 151

1. What do you understand by the term “bond energy”

**Bond energy is the energy required to break/ form one mole of chemical bond**

1. Explain the trend in bond Energy of the halogens above:

-**Decrease down the group from chlorine to Iodine**

**-Atomic radius increase down the group decreasing the energy required to break the covalent bonds between the larger atom with reduced effective nuclear @ charge an outer energy level that take part in bonding.**

(c)Group VIII elements: Noble gases

Group VIII elements are called **Noble gases.** They are all non metals. Noble gases occupy about 1.0% of the atmosphere as colourless gaseous mixture. Argon is the most abundant with 0.9%.

They exists as **monatomic** molecules with very weak van-der-waals /intermolecular forces holding the molecules. They include:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Symbol | Atomic number | Electron structure | State at room temperature |
| Helium | He | 2 | 2: | Colourless gas |
| Neon | Ne | 10 | 2:8 | Colourless gas |
| Argon | Ar | 18 | 2:8:8 | Colourless gas |
| Krypton | Kr | 36 | 2:8:18:8 | Colourless gas |
| Xenon | Xe | 54 | 2:8:18:18:8 | Colourless gas |
| Radon | Rn | 86 | 2:8:18:32:18:8 | Radioctive |

All noble gas atoms have a stable duplet(two electrons in the 1st energy level) or octet(eight electrons in other outer energy level)in the outer energy level. They therefore **do not** acquire/gain extra electron in the outer energy level or donate/lose. They therefore are therefore **zerovalent** .

The number of energy levels increases down the group from Helium to Randon.

The more the number of energy levels the bigger/larger the atomic size/radius. e.g. The atomic size/radius of Argon is bigger/larger than that of Neon because Argon has more/3 energy levels than Neon (2 energy levels).

Atomic radius noble gases increase down the group as the number of energy levels increases.

The effective nuclear attraction on the outer electrons thus decrease down the group.

The noble gases are generally unreactive because the outer energy level has the stable octet/duplet. The stable octet/duplet in noble gas atoms lead to a comparatively very high 1st ionization energy. This is because losing /donating an electron from the stable atom require a lot of energy to lose/donate and make it unstable.

As atomic radius increase down the group and the 1st ionization energy decrease, very electronegative elements like Oxygen and Fluorine are able to react and bond with lower members of the noble gases.e.g

Xenon reacts with Fluorine to form a covalent compound XeF6.This is because the outer electrons/energy level if Xenon is far from the nucleus and thus experience less effective nuclear attraction.

Noble gases have low melting and boiling points. This is because they exist as monatomic molecules joined by very weak intermolecular/van-der-waals forces that require very little energy to weaken and form liquid and break to form a gas. The intermolecular/van-der-waals forces increase down the group as the atomic radius/size increase from Helium to Radon. The melting and boiling points thus increase also down the group.

Noble gases are insoluble in water and are poor conductors of electricity.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Element | Formula of molecule | Electrical conductivity | Solubility in water | Atomic radius(nM) | 1st  ionization energy | Melting point(0C) | Boiling point(0C) |
| Helium | He | Poor | Insoluble | 0.128 | 2372 | -270 | -269 |
| Neon | Ne | Poor | Insoluble | 0.160 | 2080 | -249 | -246 |
| Argon | Ar | Poor | Insoluble | 0.192 | 1520 | -189 | -186 |
| Krypton | Kr | Poor | Insoluble | 0.197 | 1350 | -157 | -152 |
| Xenon | Xe | Poor | Insoluble | 0.217 | 1170 | -112 | -108 |
| Radon | Rn | Poor | Insoluble | 0.221 | 1134 | -104 | -93 |

Uses of noble gases

Argon is used in light bulbs to provide an inert environment to prevent oxidation of the bulb filament

Argon is used in arch welding as an insulator.

Neon is used in street and advertisement light

Helium is mixed with Oxygen during deep sea diving and mountaineering. Helium is used in weather balloon for meteorological research instead of Hydrogen because it is unreactive/inert. Hydrogen when impure can ignite with an explosion. Helium is used in making thermometers for measuring very low temperatures.

**CHEMICAL STRUCTURE**

Chemical structure is the pattern/arrangement of atoms **after** they have bonded. There are two main types of chemical structures:

**(i)simple molecular structure**

**(ii) giant structures**

(i)Simple molecular structure

Simple molecular structure is the pattern formed after atoms of non-metals have **covalently** bonded to form simple molecules.

Molecules are made of atoms joined together by weak intermolecular forces called **Van-der-waals forces.**The Van-der-waals forces hold the **molecules** together while the covalent bonds hold the **atoms** in the molecule.

**Illustration of simple molecular structure**

a)Hydrogen molecule(H2)

Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Illustration of simple molecular structure**

a)Hydrogen molecule(H2)

Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces

b)Oxygen molecule(O2)

Oxygen gas is made up of strong covalent bonds/**intra**molecular forces between each Oxygen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Strong** intramolecular forces/covalent bond

O=O:::: O=O:::: O=O:::: O=O

: : : : : : : : : : : : **weak** intermolecular O=O:::: O=O:::: O=O:::: O=O forces/van-der-waals forces

c)Iodine molecule(I2)

Iodine solid crystals are made up of strong covalent bonds/**intra**molecular forces between each iodine atom making the molecule.Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Strong** intramolecular forces/covalent bond

I--- I:::: I --- I:::: I --- I:::: I --- I

|  |  |
| --- | --- |
| : : : : : : : : : : : : : : | **weak** intermolecular |
| I --- I:::: I --- I:::: I --- I:::: I --- I | forces/van-der-waals |

forces

d)Carbon(IV) oxide molecule(CO2)

Carbon(IV) oxide gas molecule is made up of strong covalent bonds/**intra**molecular forces between each Carbon and oxygen atoms making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Strong** intramolecular forces/covalent bond

O=C=O:::: O=C=O:::: O=C=O

: : : : : : **weak** intermolecular

O=C=O:::: O=C=O:::: O=C=O forces/van-der-waals forces

The following are the main characteristic properties of simple molecular structured compounds:

a)**State**

Most simple molecular substances are gases, liquid or liquids or solid that sublimes or has low boiling/melting points at room temperature (25oC) and pressure (atmospheric pressure).

Examples of simple molecular substances include:

-all gases eg Hydrogen, oxygen, nitrogen, carbon (IV) oxide,

**-**Petroleum fractions eg Petrol, paraffin, diesel, wax,

-Solid non-metals eg Sulphur, Iodine

-Water

b) **Low melting/boiling points**

Melting is the process of **weakening** the intermolecular/ van-der-waal forces/ of attraction between the molecules that holding the substance/compound.

Note;

(i)Melting and boiling does not involve weakening/breaking the strong intramolecular force/covalent bonds holding the atoms **in** the molecule. (ii) Melting and boiling points increase with increase in atomic radius/size of the atoms **making** the molecule as the intermolecular forces / van-der-waal forces of attraction between the molecules increase. e.g.

Iodine has a higher melting/boiling point than chlorine because it has a higher /bigger atomic radius/size than chlorine, making the molecule to have stronger intermolecular force/ van-der-waal forces of attraction between the molecules than chlorine. Iodine is hence a solid and chlorine is a gas.

**(c)Insoluble in water/soluble in organic solvents**

Polar substances dissolve in polar solvents. Water is a polar solvent .Molecular substances do not thus dissolve in water because they are non-polar. They dissolve in non-polar solvents like methylbenzene, benzene, tetrachloromethane or propanone.

**d)Poor conductors of heat and electricity**

Substances with free mobile **ions** or free mobile/delocalized **electrons** conduct electricity. Molecular substances are poor conductors of heat/electricity because their molecules have no free mobile ions/electrons. This makes them very good **insulators**.

Hydrogen bonds

A hydrogen bond is an intermolecular force of attraction in which a very electronegative atom attracts hydrogen atom of another molecule.

The most electronegative elements are Fluorine, Oxygen and Nitrogen .Molecular compounds made up of these elements usually have hydrogen bonds.

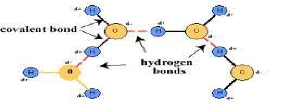
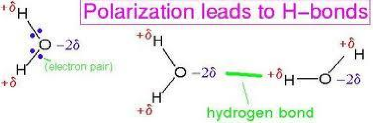
Hydrogen bonds are **stronger** than van-der-waals forces but **weaker** than covalent bonds. Molecular compounds with hydrogen bonds thus have higher melting/boiling points than those with van-der-waals forces.

**Illustration of Hydrogen bonding**

**a)Water molecule**

During formation of covalent bond, the oxygen atom attract/pull the shared electrons more to itself than Hydrogen creating partial negative charges(δ-)in Oxygen and partial positive charges(δ+)in Hydrogen.

Two molecules **attract** each other at the partial charges through Hydrogen bonding.



The hydrogen bonding in water makes it;

(i)a **liquid** with higher boiling and melting point than simple molecular substances with higher molecular mass. e.g. Hydrogen sulphide as in the table below;

Influence of H-bond in water (H2O) in comparison to H2S

|  |  |  |
| --- | --- | --- |
| **Substance** | **Water/ H2O** | **Hydrogen sulphide/ H2S** |
| Relative molecular mass | 18 | 34 |
| Melting point(oC) | 0 | -85 |
| Boiling point(oC) | 100 | -60 |

(ii)have higher **volume** in solid (ice) than liquid (water) and thus ice is less dense than water. Ice therefore floats above liquid water.

**b)Ethanol molecule**

Like in water, the oxygen atom attracts/pulls the shared electrons in the covalent bond more to itself than Hydrogen.

This creates a partial negative charge (δ-) on oxygen and partial positive charge(δ+) on hydrogen.

Two ethanol molecules attract each other at the partial charges through Hydrogen bonding forming a **dimme**r.

A dimmer is a molecule formed when two molecules join together as below:

Hydrogen bonds covalent bonds

**R1** O δ-…………………….…H δ+ O δ-

H δ+ **R2**

R1 and R2 are extensions of the molecule.

For ethanol it is made up of CH3CH2 – to make the structure:

|  |  |  |  |
| --- | --- | --- | --- |
|  |  |  |  |
|  |  |  | Hydrogen bonds covalent bonds |
|  | **CH CH** |  | O δ-………………………….…H δ+ O δ- |

**3 2**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  |  |  |  |  |  |  |
|  |  | Hδ+ |  |  |  | **CH2CH3** |

**b)Ethanoic acid molecule**

Like in water and ethanol above, the oxygen atom attracts/pulls the shared electrons in the covalent bond in ethanoic acid more to itself than Hydrogen.

This creates a partial negative charge (δ-)on oxygen and partial positive charge(δ+) on hydrogen.

Two ethanoic acid molecules attract each other at the partial charges through Hydrogen-bonding forming a **dimer.**

Hydrogen bonds covalent bonds

**R1**  **C** O δ-………………………….…H δ+ O δ-

O δ- H δ+………………..….O δ~~-~~  C **R2**

R1 and 2 are extensions of the molecule.

For ethanoic acid the extension is made up of CH3 – to make the structure;

Hydrogen bonds covalent bonds

|  |  |  |
| --- | --- | --- |
|  |  |  |
|  | O δ- H δ+…………………..……..………O δ- | C **CH3** |

**CH3**  O δ-…………………………………….…H δ+ O δ-

**C**

Ethanoic acid like ethanol exists as a dimer.

Ethanoic acid has a **higher** melting/boiling point than ethanol .This is because ethanoic acid has **two/more** hydrogen bond than ethanol.

**d) Proteins and sugars** in living things also have multiple/**complex** hydrogen bonds in their structures.

(ii) Giant structure

This is the pattern formed after substances /atoms /ions bond to form a long chain network.

Giant structures therefore extend in all directions to form a pattern that continues **repeating** itself.

There are **three** main giant structures.

a) giant covalent/atomic structure

b)giant ionic structure

c)giant metallic structure

a) giant covalent/atomic structure

Giant covalent/atomic structure is the pattern formed after atoms have covalently bonded to form long chain pattern consisting of indefinite number of atoms covalently bonded together.

The strong covalent bonds hold all the atoms together to form a very well packed structure. Examples of substances with giant covalent/atomic structure include:

1. carbon-**diamond**
2. carbon-**graphite**

(iii)silicon

(iv) silicon(IV) oxide/sand

Carbon-graphite and carbon-diamond are **allotropes** of carbon.

Allotropy is the existence of an element in more than one stable physical **form** at the same temperature and pressure.

Allotropes are atoms of the same element existing in more than one stable physical form at the same temperature and pressure.

Other elements that exhibit/show allotropy include;

-Sulphur as **monoclinic** sulphur and **rhombic** sulphur

-Phosphorus as **white** phosphorus and **red** phosphorus

**The structure of carbon-diamond**

Carbon has four valence electrons. The four valence electrons are used to form covalent bonds.

During the formation of diamond, one carbon atom covalently bond with four other carbon atoms.

C C

x x.

x **C**  x -----> C .x C x. C ------> CC C

x x.

C C

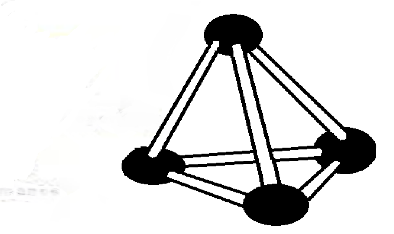
After the bonding, the atoms rearrange to form a regular **tetrahedral** in which **one** carbon is in the **centre** while **four** are at the **apex**/corners.

C

C

C C

C



This pattern repeats itself to form a long chain number of atoms covalently bonded together indefinitely. The pattern is therefore called **giant tetrahedral structure.**

It extends in all directions where one atom of carbon is always a centre of four others at the apex/corner of a regular tetrahedral.

C

C

C

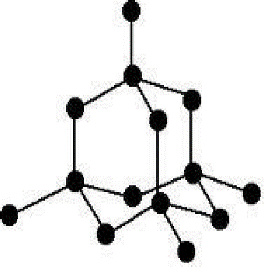
C

C

C

C

C



The giant tetrahedral structure of carbon-diamond is very well/closely packed and joined/bonded together by strong covalent bond.

This makes carbon-diamond to have the following properties:

1. **High melting/boiling point.**

The giant tetrahedral structure is very well packed and joined together by strong covalent bonds.

This requires a lot of energy/heat to weaken for the element to melt and break for the element to boil.

1. **High density.**

Carbon diamond is the **hardest** known **natural** substance.

This is because the giant tetrahedral structure is a very well packed pattern/structure and joined together by strong covalent bonds.

This makes Carbon diamond be used to make **drill** for drilling boreholes/oil wells.

The giant tetrahedral structure of carbon diamond is a very closely packed pattern /structure such that heat transfer by **conduction** is possible. This makes carbon diamond a **good thermal** conductor.

1. **Poor conductor of electricity.**

Carbon-diamond has **no free/delocalized electrons** within its structure and thus do not conduct electricity.

1. **Insoluble in water.**

Carbon-diamond is insoluble in water because it is non-polar and do not bond with water molecules.

1. **Is abrasive/Rough.**

The edges of the closely well packed pattern/structure of Carbon-diamond make its surface rough/abrasive and thus able to smoothen /cut metals and glass.

1. **Have characteristic luster.**

Carbon-diamond has a high **optical dispersion** and thus able to disperse light to different colours .This makes Carbon-diamond one of the most popular **gemstone** for making **jewellery**.

**The structure of carbon-graphite**

During the formation of graphite, one carbon atom covalently bond with **three** other carbon atoms leaving **one** free/delocalized electron.

C C

x x.

x **C**  x -----> C .x C x ------> CC x  free/delocalized electron

x x.

C C

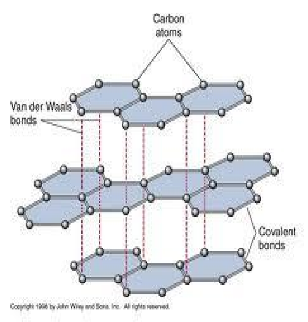
After the bonding, the atoms rearrange and join together to form a regular **hexagon** in which **six** carbon atoms are at the **apex/**corners**.**

The regular hexagon is joined to another in layers on the **same surface** by vander-waals forces.

Each layer extends to form a plane in all directions.

The fourth valence electron that does not form covalent bonding is free/mobile /delocalized within the layers.

This structure/pattern is called **giant hexagonal planar** structure.



The giant hexagonal planar structure of carbon-graphite is closely packed and joined/bonded together by strong covalent bonds. This makes carbon-graphite to have the following properties:

1. **High melting/boiling point.**

The giant hexagonal planar structure of carbon-graphite is well packed and joined together by strong covalent bonds.

This requires a lot of energy/heat to **weaken** for the element to melt and **break** for the element to boil.

1. **Good conductor of electricity.**

Carbon-graphite has **free/delocalized 4th valence electrons** within its structure and thus conducts electricity.

1. **Insoluble in water.**

Carbon-graphite is insoluble in water because it is **non-polar** and do not bond with water molecules.

1. **Soft.**

Layers of giant hexagonal planar structure of carbon graphite are held together by van-der-waals forces.

The van-der-waals forces easily break when pressed and reform back on releasing/reducing pressure/force thus making graphite soft.

1. **Smooth and slippery.**

When pressed at an **angle** the van-der-waals forces easily break and slide over each other making graphite soft and slippery.

It is thus used as a **dry** lubricant instead of oil.

**f)Some uses of carbon-graphite.**

1. As a dry lubricant- carbon graphite is smooth and slippery and thus better lubricant than oil.Oil **heat up** when reducing friction.
2. Making Lead-pencils- When pressed at an **angle** on paper the van-der-waals forces easily break and slide smoothly over contrasting background producing its characteristic black background.
3. As moderator in nuclear reactors to reduce the rate of decay/disintegration of radioactive nuclides/atoms/isotopes.
4. As electrode in dry/wet cells/battery- carbon graphite is inert and good conductor of electricity. Current is thus able to move from one electrode/terminal to the other in dry and wet cells/batteries. Carbon graphite is also very **cheap**.

b) giant ionic structure

Giant ionic structure is the pattern formed after **ions** have bonded through ionic/electrovalent bonding to form a long chain consisting of **indefinite** number of ions.

The strong ionic/electrovalent bond holds all the **cations** and **anions** together to form a very well packed structure.

Substances with giant ionic structure are mainly crystals of salts e.g. sodium chloride, Magnesium chloride, Sodium iodide, Potassium chloride, copper (II) sulphate(VI).

**The structure of sodium chloride**

Sodium chloride is made up of sodium (Na+) and chloride (Cl-)ions.

Sodium (Na+) ion is formed when a sodium atom donate /loose/donate an electron. Chloride (Cl-) ion is formed when a chlorine atom gain /acquire an extra electron from sodium atom.

Many Na+ and Cl- ions then rearrange such that **one** Na+ ion is surrounded by **six** Cl- ions and one Cl- ion is surrounded by six Na+ ions.

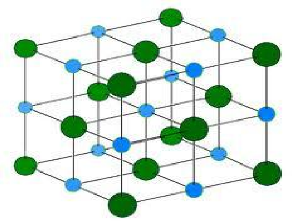
The pattern formed is a **giant cubic structure** where Cl- ion is sand witched between Na+ ions and the same to Na+ ions.

This pattern forms a **crystal**.

A crystal is a solid form of a substance in which particles are arranged in a definite pattern regularly repeated in three dimensions.

**The structure of sodium chloride**

The giant cubic structure/crystal of sodium chloride is as below;



The giant cubic structure/crystal of sodium chloride is very well packed and joined by strong ionic/electrovalent bonds. This makes sodium chloride and many ionic compounds to have the following properties:

1. **Have high melting /boiling points.**

The giant cubic lattice structure of sodium chloride is very **closely** packed into a crystal that requires a lot of energy/heat to **weaken** and melt/boil. This applies to all crystalline ionic compounds.

1. **Are good conductors of electricity in molten and aqueous state** but **poor conductor of electricity in solid.**

Ionic compounds have **fused** ions in solid crystalline state.

On heating and dissolving in water, the crystal is broken into free mobile ions (Na+ and Cl- ions).

The free mobile **ions** are responsible for **conduct**ing electricity in ionic compounds in molten and aqueous states.

**c)Soluble in water**

Ionic compounds are **polar** and dissolve in **polar** water molecules.

On dissolving, the crystal breaks to free the fused ions which are then surrounded by water molecules.

b) giant metallic structure

This is the pattern formed after metallic atoms have bonded through metallic bond.

The pattern formed is one where the metallic cations rearrange to form a cubic structure.

The cubic structure is bound together by the free delocalized electrons that move freely within.

The **more** delocalized electrons, the **stronger** the metallic bond.

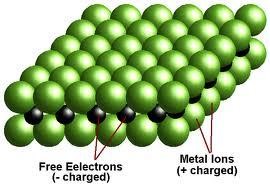
**The structure of sodium and aluminium.**

Sodium has one valence electrons.

Aluminium has three valence electrons.

After delocalizing the valence electrons ,the metal cations (Na+ and Al3+) rearrange to the apex /corners of a regular cube that extend in all directions.

The delocalized electrons remain free and mobile as shown below:



The giant cubic structure makes metals to have the following properties:

**a) Have high melting/boiling point**

The giant cubic structure is very well packed and joined/bonded together by the free delocalized electrons.

The more delocalized electrons the higher the melting/boiling point.

The larger/bigger the metallic cation ,the weaker the packing of the cations and thus the lower the melting/boiling point. e.g.

1. Sodium and potassium have both one valence delocalized electron.

Atomic radius of potassium is larger/bigger than that of sodium and hence less well packed in its metallic structure.

Sodium has therefore a higher melting/boiling point than potassium.

1. Sodium has one delocalized electron.

Aluminium has three delocalized electrons.

Atomic radius of sodium is larger/bigger than that of aluminium and hence less well packed in its metallic structure.

Aluminium has therefore a higher melting/boiling point than sodium because of the smaller well packed metallic (Al3+)ions and bonded/joined by more/three delocalized electrons.

The table below shows the **comparative** melting/boiling points of some metals:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Metal** | **Electronic structure** | **Atomic radius(nM)** | **Melting point(oC)** | **Boiling point(oC)** |
| Sodium | 2:8:1 | 0.155 | 98 | 890 |
| Potassium | 2:8:8:1 | 0.203 | 64 | 774 |
| Magnesium | 2:8:2 | 0.136 | 651 | 1110 |
| Aluminium | 2:8:3 | 0.125 | 1083 | 2382 |

**b) Good electrical and thermal conductor/electricity.**

All metals are good conductors of heat and electricity including Mercury which is a **liquid**.

The mobile delocalized electrons are free within the giant metallic structure to move from one end to the other transmitting heat/electric current.

The more delocalized electrons the better the thermal/electrical conductivity.

High temperatures/heating lowers the thermal/electrical conductivity of metals because the delocalized electrons vibrate and move randomly hindering transfer of heat

From the table above:

**Compare the electrical conductivity of;**

(i)**Magnesium and sodium**

Magnesium is a better conductor than sodium.

Magnesium has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

(ii)**Potassium and sodium**

Potassium is a better conductor than sodium.

Potassium has bigger/larger atomic radius than sodium. The delocalized electrons are less attracted to the nucleus of the atom and thus more free /mobile and thus better the electrical conductor.

1. **Insoluble in water**

All metals are insoluble in water because they are non polar and thus do not bond with water.

Metals higher in the reactivity/electrochemical series like; Potassium, sodium, Lithium and Calcium **reacts** with **cold** water producing hydrogen gas and forming an **alkaline** solution of their **hydroxides**.ie

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| 2K(s) | + | 2H2O(l) | -> | 2KOH(aq) + | H2(g) |
| 2Na(s) | + | 2H2O(l) | -> | 2NaOH(aq) + | H2(g) |
| 2Li(s) | + | 2H2O(l) | -> | 2LiOH(aq) + | H2(g) |
| Ca(s) | + | 2H2O(l) | -> | Ca(OH)2(aq)+ | H2(g) |

**Heavy** metal like Magnesium, Aluminium, Iron, Zinc and Lead **react** with **steam/water vapour** to produce hydrogen gas and form the corresponding **oxide**.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  | Mg(s) | + | H2O(g) | -> | MgO(s) | + | H2(g) |
|  | Fe(s) | + | H2O(g) | -> | FeO(s) | + | H2(g) |
|  | Zn(s) | + | H2O(g) | -> | ZnO(s) | + | H2(g) |
|  | Pb(s) | + | H2O(g) | -> | PbO(s) | + | H2(g) |
|  | 2Al(s) | + | 3H2O(g) | -> | Al2O3(s) | + | 3H2(g) |

Metals **lower** in the reactivity/electrochemical series than hydrogen like; copper, Mercury, Gold Silver and Platinum **do not** react with water/vapour.

1. **Shiny metallic-lustre**

All metals have a shiny **grey** metallic luster except copper which is **brown.**

When exposed to sunlight, the delocalized electrons gain energy, they vibrate on the metal surface scattering light to appear shiny.

With time, most metals corrode and are covered by a layer of the metal oxide.

The delocalized electrons are unable to gain and scatter light and the metal surface tarnishes/become dull.

1. **Ductile and malleable**

All metals are malleable (can be made into thin **sheet**) and ductile (can be made into **wire**.

When beaten/hit/pressed **lengthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a sheet.

When beaten/hit/pressed **lengthwise and bredthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a wire/thin strip.

1. **Have high tensile strength**

Metals are not brittle. The free delocalized electrons bind the metal together when it is bent /coiled at any angle.

The meta thus withstand stress/coiling

1. **Form alloys**

An alloy is a uniform mixture of two or more metals.

Some metals have spaces between their metallic cations which can be occupied by another metal cation with smaller atomic radius.

Common alloys include:

Brass(Zinc and Copper alloy)

Bronze(Copper and Tin alloy)

German silver

**Summary of Bonding and structure**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | **Simple molecular structure** | **Giant covalent /atomic structure** | **Giant ionic structure** | **Giant metallic structure** |
| **(i)Examples** | I2,S8,HCl,O2,CH4 | Graphite,diamond Si,SiO2 | NaCl, KCl, CaO,CuSO4 | Na,Fe,Cr,Hg,K |
| **Constituent particles making structure** | molecules | Atoms  (of non-metals) | Ions (cation and anions) | Atoms  (of metals) |
| **Type of substance** | Non-metal element/non-metal molecule/non-metal compound(electroneg ative elements) | Group IV nonmetals and some of their oxides | Metal-non metal compounds(co mpounds of electropositive and electronegative compounds) | Metallic elements  (with low electonegativity and high electropositivity) |
| **Bonding in solid state** | -Strong covalent bonds hold atoms together within separate molecules (intramolecular forces)  -Weak van-der-waals forces hold separate molecules together (intermolecular forces) | Atoms are linked through the whole structure by very strong covalent bonds. | Electrostatic attraction of cations and anions link the whole structure through strong ionic bond. | Electrostatic attraction of outer mobile electrons for positive nuclei binds atoms together though metallic bond |
| **Properties (i) Volatility** | -Highly volatile with low melting/boiling point | -Non volatile with very high melting/boiling | -Non volatile with very high melting/boiling | -Non volatile with very high |
|  | -Low latent heat of fusion/vaporization | points  -Low latent heat of fusion / vaporization | points -Low latent heat of fusion / vaporization | melting/boiling points  -Low latent heat of fusion /  vaporization |
| **(ii) State at**  **room**  **temperatur e /pressure** | Usually gases,volatile liquids or solids that sublimes | solids | solids | Solids except  Mercury(liquid) |
| **(iii)**  **Hardness** | Soft and brittle(low tensile strength) | Hard and brittle(low tensile strength) | Hard and brittle(low tensile strength) | Hard, malleable, ductile and have high tensile strength |
| **(iv)**  **Thermal**  **/electrical conductivity** | Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but some dissolve and react to form electrolytes e.g.  Hydrogen chloride and ammonia gases. | Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but  -Carbon-graphite  is a good electrical conductor while -Carbon-diamond is a good thermal conductor. | Poor thermal and electrical conductor when solid. Good thermal and electrical conductor in liquid/molten and aqueous states when the ions are not  fused | Good thermal and electrical conductor in solid and liquid/molten states due to the free mobile /delocalized  electrons |
| **(v)**  **Solubility** | Insoluble in polar solvents e.g. Water Soluble in non-polar solvents e.g.  tetrachloromethane, benzene, methylbenzene | Insoluble in all solvents | Soluble in polar solvents e.g.  Water  Insoluble in non-polar solvents e.g. tetrachlorometh ane, benzene, methylbenzene | Insoluble in polar/non-polar colvents.  -Some react with polar solvents -Some metal dissolve in other metals to form alloys e.g. Brass is formed when Zinc dissolve in copper. |

**C. PERIODICITY OF BONDING AND STRUCTURE**

The periodic table does not classify elements as metals and non-metals. The table arranges them in terms of atomic numbers.

However, based on structure and bonding of the elements in the periodic table;

(i)-the top right hand corner of about twenty elements are non-metals

(ii)-left of each non-metal is an element which shows characteristics of both metal and non-metal.

These elements are called **semi-metals/metalloids**. They include Boron, silicon, Germanium, Arsenic, and Terullium

(iii)-all other elements in the periodic table are metal.

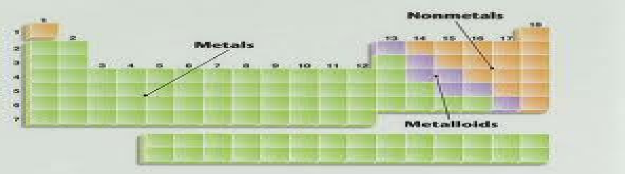
(iv)-Hydrogen is a non-metal with metallic characteristic/property of donating/losing outer electron to form cation/H+ ion.

1. –bromine is the only known natural liquid non-metal element at room temperature and pressure.
2. –mercury is only known natural liquid metal element at room temperature and pressure.
3. Carbon-graphite is a semi metals/metalloids. Carbon-diamond is a pure non-metal yet both are allotropes of carbon (same element)

**a) Sketch of the periodic table showing metals ,metalloid and non-metals**

Metals **Metalloids**  *Non-metals*

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| *H* |  |  |  |  |  |  |  | *He* |
| Li | Be |  | **B** | **C** | *N* | *O* | *F* | *Ne* |
| Na | Mg |  | Al | **Si** | *P* | *S* | *Cl* | *Ar* |
| K | Ca | Transition metals | Ga | Ge | **As** | *Se* | *Br* | *Kr* |
| Rb | Sr |  | In | Sn | Sb | **Te** | *I* | *Xe* |
| Cs | Ba |  | Tl | Pb | Bi | Po | **At** | *Rn* |
| Fr | Ra |  |  |  |  |  |  |  |



**b)Periodicity in the physical properties of elements across period 2 and 3**

Study table I and II below:

Table I(period 2)

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **Property** | **Li** | **Be** | **B** | **C** | **N** | **O** | **F** | **Ne** |
| **Melting point(oC)** | 180 | 1280 | 2030 | 3700  (graphite)  3550  (diamond) | -210 | -219 | -220 | -250 |
| **Boiling point(oC)** | 1330 | 2480 | 3930 | Graphite sublimes 4830  (diamond) | -200 | -180 | -190 | -245 |
| **Density at room**  **temperatu re (gcm-3)** | 0.50 | 1.85 | 2.55 | 2.25  (graphite)  3.53  (diamond) | 0.81 | 0.14 | 0.11 | 0.021 |
| **Type of element** | Metal | Metal | Metal | Metalloid | Non-  metal | Non-  metal | Non-  metal | Non-  metal |
| **Chemical structure** | Giant metallic | Giant metallic | Giant atomic/ covalent | Giant atomic/ covalent | Simple  molecul  a  or  molecul e/ N2 | Simple molecula  or  molecule  s /O2 | Simple  molecul  a  or  molecul e/F2 | Simple  molecul  a  or  molecul e/Ne |
| **State at room temperatu**  **re** | Solid | Solid | Solid | Solid | gas | gas | gas | gas |
| **Electron structure** | 2:1 | 2:2 | 2:3 | 2:4 | 2:5 | 2:6 | 2:7 | 2:8 |
| **Valency** | 1 | 2 | 3 | 4 | 3 | 2 | 1 | - |
| **Formular**  **of ion** | Li+ | Be2+ | B3+ | - | N3- | O2- | F- | - |

Table II (period 3)

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **Property** | **Na** | **Mg** | **Al** | **Si** | **P(white**  **)** | **S(Rhomb**  **ic)** | **Cl** | **Ar** |
| **Melting point(oC)** | 98 | 650 | 660 | 1410 | 44 | 114 | -101 | -189 |
| **Boiling point(oC)** | 890 | 1120 | 2450 | 2680 | 280 | 445 | -34 | -186 |
| **Density at room**  **temperatu re (gcm-3)** | 0.97 | 1.74 | 2.70 | 2.33  ( graphite  )  3.53  ( diamon  d) | 1.82 | 2.07 | 0.157 | 0.011 |
| **Type of element** | Metal | Metal | Metal | Metalloi d | Non-  metal | Non-metal | Non-  metal | Non-  metal |
| **Chemical structure** | Giant metallic | Giant metallic | Giant metallic | Giant atomic/ covalent | Simple  molecul  a  or  molecul  e/ P4 | Simple molecula  or  molecules  /S8 | Simple  molecul  a  or  molecul e/Cl2 | Simple molecula  or  molecule  /Ar |
| **State at room**  **temperatu**  **re** | Solid | Solid | Solid | Solid | Solid | Solid | gas | gas |
| **Electron structure** | 2:8:1 | 2:8:2 | 2:8:3 | 2:8:4 | 2:8:5 | 2:8:6 | 2:8:7 | 2:8:8 |
|  |  |  |  |  |  |  |  |  |
| **Valency** | 1 | 2 | 3 | 4 | 3 | 2 | 1 | - |
| **Formular**  **of ion** | Na+ | Mg2+ | Al3+ | - | P3- | S2- | Cl- | - |

From table I and II above:

**1. Explain the trend in atomic radius along /across a period in the periodic table**

Observation

Atomic radius of elements in the same period decrease successively across/along a period from left to right.

Explanation

Across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level.

Increase in nuclear charge increases the effective nuclear attraction on the outer energy level pulling it closer to the nucleus successively across the period .e.g.

(i)From the table 1and 2 above, atomic radius of Sodium (0.157nM) is higher than that of Magnesium(0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus than sodium.

(ii)The rate of decrease in the atomic radius become smaller as the atom become heavier e.g. Atomic radius of Magnesium from sodium falls by(0.157nM-

0.137nM) =**0.02**

Atomic radius of Chlorine from sulphur falls by(0.104nM- 0.099nM) =**0.005**

This is because gaining/adding one more proton to 11 already present cause greater proportional change in nuclear attraction power to magnesium than gaining/adding one more proton to 16 already present in sulphur to chlorine.

(iii)Period 3 elements have more energy levels than Period 2 elements. They have therefore bigger/larger atomic radius/size than corresponding period 2 elements in the same group.

**2.Explain the trend in ionic radius along/across a period in the periodic table**

Observation

Ionic radius of elements in the same period decrease successively across/along a period from left to right for the first three elements then increase drastically then slowly successively decrease.

Explanation

Across/along the period from left to right elements change form electron donors/losers (**reducing** agents) to electron acceptors (**oxidizing** agents).

(i)An atom form stable ion by either gaining/acquiring/ accepting extra electron or donating/losing outer electrons.

(ii)Metals form stable ions by donating/losing **all** the outer energy level electrons and thus **also** the outer energy level .i.e.

-Sodium ion has one less energy level than sodium atom. The ion is formed by sodium atom donating/losing (all) the outer energy level electron and thus also the outer energy level making the ion to have smaller ionic radius than atom.

(iii)Ionic radius therefore decrease across/along the period from Lithium to Boron in period 2 and from Sodium to Aluminium in period 3.This is because the number of electrons donated/lost causes increased effective nuclear attraction on remaining electrons /energy levels.

(iv)Non-metals form stable ion by gaining/acquiring/accepting extra electron in the outer energy level. The extra electron/s increases the repulsion among electrons and reduces the effective nuclear attraction on outer energy level. The outer energy level therefore expand/enlarge/increase in order to accommodate the extra repelled electrons .The more electrons gained/accepted/acquired the more repulsion and the more expansion to accommodate them and hence bigger/larger atomic radius. e.g.

**-**Nitrogen ion has three electrons more than Nitrogen atom. The outer energy level expand/enlarge/increase to accommodate the extra repelled electrons. Nitrogen atom thus has smaller atomic radius than the ionic radius of nitrogen ion.

(v) Ionic radius decrease from group IV onwards from left to right. This because the number of electrons gained to form ion decrease across/along the period from left to right. e.g. Nitrogen ion has bigger/larger ionic radius than Oxygen.

**3.Explain the trend in melting and boiling point of elements in a period in the periodic table.**

Observation

The melting and boiling point of elements rise up to the elements in Group

Explanation

Melting/boiling points depend on the packing of the structure making the element and the strength of the bond holding the atoms/molecules together.

Across/along the period (2 and 3) the structure changes from giant metallic, giant atomic/covalent to simple molecular.

(i)For metals, the number of delocalized electrons increases across/along the period and hence stronger metallic bond/structure thus requiring a lot of heat/energy to weaken.

The strength of a metallic bond also depends on the atomic radius/size. The melting /boiling point decrease as the atomic radius/size of metals increase due to decreased packing of larger atoms. e.g.

-The melting /boiling point of Lithium is lower than that of Beryllium because Beryllium has two/more delocalized electrons and hence stronger metallic structure/bond.

- The melting /boiling point of Lithium is higher than that of Sodium because the atomic radius/size Lithium is smaller and hence better packed and hence forms stronger metallic structure/bond.

(ii)Carbon-graphite/carbon-diamond in period 2 and Silicon in period 3 form very well packed giant atomic/covalent structures held together by strong covalent bonds. These elements have therefore very high melting/boiling points.

Both Carbon-graphite/ carbon-diamond have smaller atomic radius/size than Silicon in period 3 and thus higher melting/boiling points due to better/closer packing of smaller atoms in their well packed giant atomic/covalent structures.

(ii)Non-metals from group V along/across the period form simple molecules joined by weak intermolecular /van-der-waals force. The weak intermolecular /van-derwaals force require little energy/heat to weaken leading to low melting/boiling points. The strength of the intermolecular /van-der-waals forces decrease with decrease in atomic radius/ size lowering the melting/boiling points along/across the period (and raising the melting/boiling points down the group).e.g.

-The melting /boiling point of Nitrogen is higher than that of Oxygen. This is because the atomic radius/ size of Nitrogen is higher than that of Oxygen and hence stronger intermolecular /van-der-waals forces between Nitrogen molecules.

-The melting /boiling point of Chlorine is higher than that of Fluorine. This is because the atomic radius/ size of Chlorine is higher than that of Fluorine and hence stronger intermolecular /van-der-waals forces between Chlorine molecules.

(iii)Rhombic sulphur exists as a puckered ring of S8atoms which are well packed. Before melting the ring break and join to very long chains that entangle each other causing the **unusually** high melting/boiling point of Rhombic sulphur.

(iv)Both sulphur and phosphorus exists as allotropes.

Sulphur exists as **Rhombic**-sulphur and **monoclinic**-sulphur. Rhombic-sulphur is the stable form of sulphur at room temperature and pressure.

Phosphorus exists as **white**-phosphorus and **red**-phosphorus.

White-phosphorus is the stable form of Phosphorus at room temperature and pressure.

**4. State and explain the trend in density of elements in a period in the periodic table.**

Observation: Density increase upto the elements in group IV then falls across/along the period successively Explanation:

Density is the mass per unit volume occupied by matter/particles/atoms/molecules of element.

(i)For metals ,the stronger metallic bond and the more delocalized electrons ensure a very well packed giant metallic structure that occupy less volume and thus higher density.

The more the number of delocalized electrons along/across the period, the higher the density. e.g.

(i)Aluminium has a higher density than sodium. This is because aluminium has more /three delocalized electrons than /one sodium thus forms a very well packed giant metallic structure that occupy less volume per given mass/density.

(ii)Carbon-graphite ,carbon-diamond and silicon in group IV form a well packed giant atomic/covalent structure that is continuously joined by strong covalent bonds hence occupy less volume per given mass/density.

Carbon-graphite form a less well packed giant hexagonal planar structure joined by Van-der-waals forces. Its density (2.25gcm-3) is therefore less than that of Carbondiamond(3.53gcm-3) and silicon(2.33gcm-3).Both diamond and silicon have giant tetrahedral structure that is better packed. Carbon-diamond has smaller atomic radius/size than silicon. Its density is thus higher because of better packing and subsequently higher density. Carbon-diamond is the hardest known natural substance by having the highest density.

(iii)For non-metals, the strength of the intermolecular /van-der-waals forces decreases with decrease in atomic radius/size along/across the period. This decreases the mass occupied by given volume of atoms in a molecule from group VI onwards. e.g.

Phosphorus has a higher atomic radius/size than chlorine and Argon and thus stronger intermolecular/van-der-waals forces that ensure a given mass of phosphorus occupy less volume than chlorine and neon.

**5.State and explain the trend in thermal/electrical conductivity of elements in a period in the periodic table.**

Observation:

Increase along/across the period from group I, II, and III then decrease in Group IV to drastically decrease in group V to VIII (O).

Explanation

(i)Metals have free delocalized electrons that are responsible for thermal/electrical conductivity.Thermal/electrical conductivity increase with increase in number of delocalized electrons. The thermal conductivity decrease with increase in temperature/heating.

e.g.

Aluminium with three delocalized electrons from each atom in its metallic structure has the highest electrical /thermal conductivity in period 3.

(ii)Carbon-graphite has also free 4th valency electrons that are delocalized within its layers of giant hexagonal planar structure. They are responsible for the electrical conductivity of graphite.

(iii)Silicon and carbon diamond do not conduct electricity but conducts heat. With each atom too close to each other in their very well packed giant tetrahedral structure, heat transfer /radiate between the atoms. The thermal conductivity increase with increase in temperature/heating.

(iv)All other non-metals are poor /non-conductor of heat and electricity. They are made of molecules with no free /mobile delocalized electrons in their structure.

**Periodicity of the oxides of elements along/across period 3**

The table below summarizes some properties of the oxides of elements in period 3 of the periodic table.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **Formular of oxide/ Property** | **Na2O** | **MgO** | **Al2O3** | **SiO2** | **P2O5**  **P4O6** | **SO2** SO**3** | **Cl2O7 Cl2O** |
| **Melting**    **point(oC)** | 1193 | 3075 | 2045 | 1728 | 563 | -76 | -60 |
| **Boiling**    **point(oC)** | 1278 | 3601 | 2980 | 2231 | 301 | -10 | -9 |
| **Bond type** | Ionic | Ionic | Ionic | Covalent | Covale  nt | Covalent | Covalent |
| **Chemical structure** | Giant ionic structur e | Giant ionic structur e | Giant ionic structure | Giant atomic/ covalent | Simple molecul  a  or  molecul  e | Simple molecula  or  molecules | Simple molecula  or  molecule |
| **State at room**  **temperatu**  **re** | Solid | Solid | Solid | Solid | Solid | gas | Gas  (**Cl2O7** is a liquid) |
| **Nature of Oxide** | Basic/ alkaline | Basic/ alkaline | Amphotell ic oxide | 2:8:4 | 2:8:5 | 2:8:6 | 2:8:7 |

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **Reaction**  **with water** | React to form  NaOH  /alkalin e solution | React to form  MgOH)  2  /weakly alkaline solution | Don‘t react with water. | Don‘t react with water. | React to form  H2PO4  /weakly acidic solution | -SO2 react to form H2SO3 . H2SO3 is quickly oxidized to H2SO4 -SO2 react to form H2SO4/ strongly acidic | **-Cl2O7 reacts to form**  **HClO4**/weakly acidic solution |
| **Reaction with dilute acids** | Reacts to form salt and water | Reacts to form salt and water | Reacts to form salt and water | No  reaction | No  reaction | No  reaction | No  reaction |

1. All the oxides of elements in period 3 except those of sulphur and chlorine are **solids** at room temperature and pressure.
2. Across/along the period, bonding of the oxides changes from **ionic** in sodium oxide magnesium oxide and aluminium oxide (show both ionic and covalent properties) to **covalent** in the rest of the oxides.

1. Across/along the period, the structure of the oxides changes from giant **ionic** structure in sodium oxide, magnesium oxide and aluminium oxide to giant **atomic**/covalent structure in silicon (IV) oxide. The rest of the oxides form simple **molecules**/**molecular** structure.

1. Sodium oxide and magnesium oxide are **basic** **/alkaline** in nature. Aluminium oxide is **amphotellic** in nature (shows both acidic and basic characteristics). The rest of the oxides are **acidic** in nature.

1. Ionic compounds/oxides have very **high** melting/boiling points because of the strong **electrostatic attraction** joining the giant ionic crystal lattice.

The melting/boiling points increase from sodium oxide to aluminium oxide as the number of electrons involved in bonding increase, increasing the strength of the ionic bond/structure.

1. Silicon (IV) oxide is made of a well packed giant atomic/covalent structure joined by strong covalent bonds.

This results in a solid with very **high** melting/boiling point.

7.Phosphorus (V) oxide, sulphur(IV) oxide/ sulphur (VI) oxide and dichloride heptoxide exist as simple molecules/molecular structure joined by weak van-derwaals/intermolecular forces.

This results in them existing as **low** melting /boiling point solids/gases.

1. Ionic oxide conducts electricity in molten and aqueous states but not in solid.

In solid state the ions are fused/fixed but on heating to molten state and when dissolved in water, the ions are free / mobile.

Sodium oxide, magnesium oxide and aluminium oxide are therefore good conductors in molten and aqueous states.

1. Covalent bonded oxides do not conduct electricity in solid, molten or in aqueous states.

This is because they do not have free / mobile ion. Phosphorus (V) oxide, sulphur(IV) oxide/ sulphur (VI) oxide and dichloride heptoxide are thus nonconductors/insulators.

1. Silicon (IV) oxide is a poor/weak conductor of heat in solid state. This is because it has very closely packed structure for heat to radiate conduct along its structure.

1. Electopositivity decrease across the period while electronegativity increase across the period. The oxides thus become less ionic and more covalent along/across the period.

12.The steady change from giant ionic structure to giant atomic/ covalent structure then simple molecular structure lead to profound differences in the reaction of the oxides with water,acids and alkalis/bases:

(i) Reaction with water

a) Ionic oxides react with water to form alkaline solutions e.g.;

I.Sodium oxide reacts/dissolves in water forming an alkaline solution of sodium hydroxide.

Chemical equation: Na2O(s) + H2O (l) -> 2NaOH(aq)

1. Magnesium oxide slightly/ slowly reacts/dissolves in water forming an alkaline solution of magnesium hydroxide

Chemical equation: MgO(s) + 2H2O (l) -> Mg(OH) 2 (aq)

1. Aluminium oxide does reacts/dissolves in water.
2. Non-metallic oxides are acidic. They react with water to form weakly acidic solutions:
   1. Phosphorus (V) oxide readily reacts/dissolves in water forming a weak acidic solution of phosphoric (V) acid.

Chemical equation: P4O10 (s) + 6H2O (l) -> 4H3PO4 (aq)

Chemical equation: P2O5 (s) + 3H2O (l) -> 2H3PO4 (aq)

* 1. Sulphur (IV) oxide readily reacts/dissolves in water forming a weak acidic solution of sulphuric (IV) acid.

Chemical equation: SO2 (g) + H2O (l) -> H2SO3 (aq)

Sulphur (VI) oxide quickly fumes in water to form concentrated sulphuric (VI) acid which is a strong acid.

Chemical equation: SO3 (g) + H2O (l) -> H2SO4 (aq)

* 1. Dichlorine oxide reacts with water to form weak acidic solution of chloric(I) acid/hypochlorous acid.

Chemical equation: Cl2O (g) + H2O (l) -> 2HClO (aq)

* 1. Dichlorine heptoxide reacts with water to form weak acidic solution of chloric(VII) acid.

Chemical equation: Cl2O7 (l) + H2O (l) -> 2HClO4 (aq)

1. Silicon (IV) oxide **does not** react with water.

It reacts with hot concentrated alkalis forming silicate (IV) salts. e.g.

Silicon (IV) oxide react with hot concentrated sodium hydroxide to form sodium silicate (IV) salt.

Chemical equation: SiO2 (s) + 2NaOH (aq) -> Na2SiO3 (aq) + H2O (l)(ii) Reaction with dilute acids

1. Ionic oxides react with dilute acids to form salt and water only. This is a **neutralization** reaction. e.g.

Chemical equation: Na2O(s) + H2SO4 (aq) -> Na2SO4 (aq) + H2O(l)

Chemical equation: MgO(s) + 2HNO3(aq) -> Mg (NO3) 2 (aq) + H2O(l)

Chemical equation: Al2O3 (s) + 6HCl(aq) -> 2AlCl3 (aq) + 3H2O(l)

Aluminium oxide is amphotellic and reacts with **hot concentrated strong** alkalis sodium/potassium hydroxides to form **complex** sodium aluminate(III) and potassium aluminate(III) salt.

Chemical equation: Al2O3 (s) + 2NaOH(aq) + 3H2O(l) -> 2 NaAl(OH)4 (aq)

Chemical equation: Al2O3 (s) + 2KOH(aq) + 3H2O(l) -> 2 KAl(OH)4 (aq)

1. Acidic oxides **do not** react with dilute acids.

**c)Periodicity of the Chlorides of elements along/across period 3**

The table below summarizes some properties of the chlorides of elements in period 3 of the periodic table.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **Formular**  **of chloride/ Property** | **NaCl** | **MgCl2** | **AlCl3** | **SiCl4** | **PCl5**  **PCl3** | **SCl2** S**2**Cl**2** | **Cl2** |
| **Melting**    **point(oC)** | 801 | 714 | Sublimes  at  180 **oC** | -70 | PCl5 Sublimes  at  -94 **oC** | -78 | -101 |
| **Boiling**    **point(oC)** | 1465 | 1418 | 423(as  **Al2Cl6** vapour | 57 | 74(as  **P2Cl6**  Vapour  164 (as **PCl5)** | decompos es at 59 **oC** | -34 |
| **Bond type** | Ionic | Ionic | Ionic/ Covalent/ dative | Covalent | Covalent | Covalent | Covalent |
| **Chemical structure** | Giant ionic | Giant ionic | Molecular  / | Simple molecula | Simple molecula | Simple molecula | Simple molecula |

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  | structur e | structur e | dimerizes | or  molecule | or  molecule | or  molecules | or  molecule |
| **State at room temperatu**  **re** | Solid | Solid | Solid | liquid | Liquid **PCl5**  **is solid** | liquid | Gas |
| **Nature of Chloride** | Neutral | Neutral | Strongly acidic | Strongly acidic | Strongly acidic | Strongly acidic | Strongly acidic |
| **pH of solution** | 7.0 | 7.0 | 3.0 | 3.0 | 3.0 | 3.0 | 3.0 |
| **Reaction**  **with water** | Dissolv  e | Dissolv e | -  Hydrolyse d by water -Acidic hydrogen chloride fumes produced | -  Hydrolyse d by water -Acidic hydrogen chloride fumes produced | Hydrolyse d by water -Acidic hydrogen chloride fumes produced | Hydrolyse d by water -Acidic hydrogen chloride fumes produced | Forms  HCl and  HClO |
| **Electrical conductivit y in**  **molten/aq ueous state** | good | good | poor | nil | nil | nil | nil |

1. Sodium Chloride, Magnesium chloride and aluminium chloride are **solids** at room temperature and pressure.

Silicon(IV) chloride, phosphorus(III)chloride and disulphur dichloride are **liquids.** Phosphorus(V)chloride is a **solid**. Both chlorine and sulphur chloride are **gases.**

1. Across/along the period bonding changes from **ionic** in Sodium Chloride and Magnesium chloride to **covalent** in the rest of the chlorides.

1. Anhydrous aluminium chloride is also a molecular compound .Each aluminium atom is covalently bonded to three chlorine atoms.

In vapour/gaseous phase/state two molecules dimerizes to Al2O6 molecule through coordinate/dative bonding.

1. Across/along the period the structure changes from giant **ionic** in Sodium

Chloride and Magnesium chloride to **simple molecules/molecular structure** in the rest of the chlorides.

1. Ionic chlorides have very high melting /boiling points because of the strong ionic bond/electrostatic attraction between the ions in their crystal lattice.The rest of the chlorides have low melting /boiling points because of the weak van-der-waal /intermolecular forces.

1. Sodium Chloride and Magnesium chloride in molten and aqueous state have free/mobile ions and thus good electrical conductors. Aluminium chloride is a poor conductor. The rest of the chlorides do not conduct because they have no free/mobile **ions**.

1. Ionic chloride form **neutral** solutions with pH =7. These chlorides **ionize**/**dissociate** completely into free cations and anions.i.e;

Sodium Chloride and Magnesium chloride have pH=7 because they are fully/completely ionized/dissociated into free ions.

Chemical equation NaCl (s) -> Na+(aq) + Cl-(aq)

Chemical equation MgCl2 (s) -> Mg2+(aq) + 2Cl-(aq)

8 Across/along the period from aluminium chloride, **hydrolysis** of the chloride takes place when reacting/dissolved in water.

Hydrolysis is the reaction of a compound when dissolved in water.

a)Aluminium chloride is hydrolyzed by water to form aluminium **hydroxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation AlCl3 (s) + 3H2O(l)-> Al(OH)3( s) +

3HCl(g)

b)Silicon(IV) chloride is hydrolyzed by water to form silicon(IV)**oxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation SiCl4 (l) + 2H2O(l)-> SiO2(s) + 4HCl(g)

This reaction is highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

1. Both phosphoric (V) chloride and phosphoric (III) chloride are hydrolyzed by water to form phosphoric (V) **acid** and phosphoric (III) **acid** respectively. Fumes of hydrogen chloride gas are produced. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation PCl5 (s) + 4H2O(l)-> H3PO4( aq) + 5HCl(g)

Chemical equation PCl3 (s) + 3H2O(l)-> H3PO4( aq) +

3HCl(g)

This reaction is also highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

1. Disulphur dichloride similarly hydrolyzes in water to form **yellow** deposits of sulphur and produce a mixture of **sulphur (IV) oxide** and **hydrogen chloride** gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation 2S2Cl2 (l) + 2H2O(l)-> 3S(s) + SO2( g) +

4HCl(g)

**D. COMPREHENSIVE REVISION QUESTIONS**

1.The grid below represents periodic table. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
|  |  | | | | | | | A |
| B |  |  |  | G |  | H | E | C |
|  | J | I | L |  |  |  |  |
| D | ***N*** |  |  |  |  |  | M |  |

(a) (I) Indicate on the grid the position of an element represented by letter N whose electronic configuration of a divalent cation is 2:8:8 . ( 1 mark )

(II) Name the bond formed between D and H react. Explain your answer.(2 marks )

**Ionic/electrovalent**

**D is electropositive thus donates two electrons to electronegative H**

1. Write an equation for the reaction between B and water. ( 1 mark )

**Chemical equation 2B (s) + 2H2O(l) -> 2BOH(aq) + H2**

**(g)**

1. How do the atomic radii of I and L compare. Explain. ( 2 marks )

1. In terms of structure and bonding explain why the oxide of G has lower

melting point than oxide of L. ( 2 marks )

(b) Study the information given below and answer the question that follow.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Formula of compound N | NaCl | MgCl 2 | Al2Cl6 | SiCl4 | PCl3 | SCl2 |
| B.P(0C) | 1470 | 1420 | Sublimes | 60 | 75 | 60 |
| M.P(0C) | 800 | 710 | At  8000C | -70 | 90 | -80 |

( I)Why is the formula of aluminium chloride given as Al2Cl6 and not AlCl3 ?

( 1 mark )

1. Give two chlorides that are liquid at room temperature. Give a reason for the

answer. (2 marks )

1. Give a reason why Al2Cl6 has a lower melting point than MgCl2 although both

Al and Mg are metals. (1 mark )

1. Which of the chlorides would remain in liquid state for the highest

temperature range explain why ? (2 mark )

(Kakamega)

2. a) Study the information given below and answer the questions that follow.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Atomic radius (nm) | Ionic radius | Formula of oxide | Melting point of oxide (‗C) |
|  |  | (nm) |  |  |
| P  Q  R  S  T | 0.364  0.830  0.592  0.381  0.762 | 0.421  0.711  0.485  0.446  0.676 | A2O  BO2  E2O3  G2O5  JO | -119  837  1466  242  1054 |

1. Which elements are non-metals? Give a reason. (2mks)

1. Explain why the melting point of the oxide of R is higher than that of the

oxide of S. (2mks)

1. Give **two** elements that would react vigorously with each other. Explain

your answer. (2mks)

b) Study the information in the table below and answer the questions that follow (The letters do not represent the actual symbols of the elements)

|  |  |  |  |
| --- | --- | --- | --- |
|  | | Ionization Energy\_kJ/Mole | |
| Element | Electronic configuration | 1st ionization energy | 2nd ionization energy |
| A | 2.2 | 900 | 1800 |
| B | 2.8.2 | 736 | 1450 |
| C | 2.8.8.2 | 590 | 1150 |

1. What chemical family do the elements A, B and C belong? (1mk)

1. What is meant by the term ionization energy? (1mk) iii) The 2nd ionization energy is higher that the 1st ionization energy of each. Explain

(1mk)

(iv)When a piece of element C is placed in cold water, it sinks to the bottom and an effervescence of a colourless gas that burns explosively is produced. Use a simple diagram to illustrate how this gas can be collected during this experiment.

(3mks)

3. The grid below represents part of the periodic table. The letters do not represent the actual symbols.

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
|  |  | | | | | | | A |
| B |  |  | X | G |  | Z | E | V |
|  | J | I | L |  | T |  |  |
| D | ***N*** |  |  |  |  |  | M |  |

1. Select the most reactive non-metal.

(1mk)

1. Write the formula of the compound consisting of

I.D and Z only.

(2mk)

II. X and Z

1. Select an element that can form an ion of change +2 (1mk)

1. Which element has the least ionization energy? Explain

(2mks)

1. Suggest with reasons a likely pH value of an aqueous solution of the chlorine of:(3mks)

T.

B

X

1. To which chemical family do the following elements belong? (2mk)

J

V

1. An element K has relative atomic mass of 40.2.It has two isotopes of masses 39 and 42. Calculate the relative abundance of each isotope. (3mks)

4.The grid below shows part of the periodic table study it and answer the questions that follow. The letters do not represent the true symbols.

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
|  |  | | | | | | |  |
|  |  |  |  |  |  | **A** |  |  |
|  | **B** | **C** |  | **D** |  | **E** |  |
| **F** | **G** |  |  |  |  |  |  |  |
|  |  |  |  |  |  |  | **H** |  |
|  |  |  |  |  |  |  |  |  |

|  |  |
| --- | --- |
| \(a) Which element forms ions with charge of 2-. Explain | (2mks) |
| (b) What is the nature of the oxide formed by C. | (1mk) |
| (c) How does the reactivity of H compare with that of E. Explain? | (2mks) |
| (d)Write down a balanced equation between B and Chlorine. | (1mk) |
| (e) Explain how the atomic radii of F and G compare. | (1mk) |

(f) If the oxides of F and D are separately dissolved in water, state and explain the

effects of their aqueous solutions on litmus. (3mks)

1. (a) The grid below show part of the periodic table.(The letter do not represent the actual symbols).Use it to answer the questions that follow.

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **T** |  | | | | | | | **Q** |
|  |  |  |  | **S** |  | **R** | **K** |  |
| **A** | **J** |  | **Y** |  | **U** |  | **L** |  |
| **W** |  |  |  |  |  |  | **M** | **B** |
|  | **C** |  |  |  |  |  | **N** |  |
| **P** |  |  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |  |  |

(i)Select the most reactive non-metal. (1mk)

(ii)Select an element that forms a divalent cation. (1mk)

(iii)Element **Z** has atomic number 14.Show its position in the grid. (1mk)

(iv)How do the atomic radii of **U** and **J** compare? (2mks)

(v)How do electrical conductivity of **A** and **Y** compare? (2mks)

(vi)How does the boiling point of elements **K, L** and **M** vary? Explain (2mks (b) The table below gives information on four elements by letters K, L, M and N. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Electron arrangement** | **Atomic radius** | **Ionic radius** |
| **K** | **2:8:2** | **0.136** | **0.065** |
| **L** | **2:8:7** | **0.099** | **0.181** |
| **M** | **2:8:8:1** | **0.203** | **0.133** |
| **N** | **2:8:8:2** | **0.174** | **0.099** |

* + 1. Which two elements have similar properties? Explain. (2mks)

* + 1. Which element is a non-metal? Explain. (1mk)
    2. Which one of the elements is the strongest reducing agent. (1mk)

1. The grid given below represents part of the periodic table study it and answer the questions that follow. (The letters do not represent the actual symbols of the elements.)

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
|  |  | | | | | | | **A** |
|  |  |  |  | **B** |  |  |  |  |
|  | **C** | **D** |  |  | **E** |  |  |
|  | **F** |  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |  |  |

* + - 1. What name is given to the group of elements to which C and F belong?

(1mk)

* + - 1. Which letter represents the element that is the least reactive?

(1mk)

* + - 1. What type of bond is formed when B and E react? Explain (2mks)

(iv)Write formula of the compound formed where elements D and oxygen gas react. (1mk)

(v) On the grid indicate the a tick (√) the position of element G which is in the third period of the periodic table and forms G3- ions. (1mk)

(b) Study the information in the table below and answer the questions that follow. (The letter do not represents the actual symbols of the substance).

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Substance | Melting point oC | Boiling point oC | Solubility in water | Density at room. Temp/g/cm3 |
| H | -117 | 78.5 | Very soluble | 0.8 |
| J | -78 | -33 | Very soluble | 0.77x 1-3 |
| K | -23 | 77 | Insoluble | 1.6 |
| L | - 219 | -183 | Slightly Soluable | 1.33 x 10-3 |

I.(i) Which substance would dissolve in water and could be separated from the solution by fractional distillation. (1mk)

(ii) Which substances is a liquid at room temperature and when mixed with water two layers would be formed? (1mk)

II. Which letter represents a substance that is a gas at room temperature and which can be collected ;

* + - 1. Over water?

(1mk)

* + - 1. By downward displacement of air? Density of air at room temperature =

1.29 x 10-3 g/C (1mk)

**SALTS**

**INTRODUCTION TO SALTS**

1.(a) A salt is an ionic compound formed when the cation from a base combine with the anion derived from an acid.

A salt is therefore formed when the hydrogen ions in an acid are replaced wholly/fully or partially/partly ,directly or indirectly by a metal or ammonium radical.

1. The number of ionizable/replaceable hydrogen in an acid is called **basicity** of an acid.

Some acids are therefore:

(i)**mono**basic acids generally denoted **H**X e.g.

**H**Cl, **H**NO3,HCOO**H**,CH3COO**H**.

(ii)**di**basic acids ; generally denoted **H2**X e.g.

**H2**SO4, **H2**SO3, **H2**CO3,**H**OOCOO**H**.

(iii)**tri**basic acids ; generally denoted **H3**X e.g.

**H3**PO4.

1. Some salts are **normal** salts while other are **acid** salts.

(i)A normal salt is formed when all the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.

(ii)An acid salt is formed when part/portion the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.

Table showing normal and acid salts derived from common acids

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Acid name** | **Chemical formula** | **Basicity** | **Normal salt** | **Acid salt** |
| Hydrochloric acid | HCl | Monobasic | Chloride(Cl-) | None |
| Nitric(V)acid | HNO3 | Monobasic | Nitrate(V)(NO3-) | None |
| Nitric(III)acid | HNO2 | Monobasic | Nitrate(III)(NO2-) | None |
| Sulphuric(VI)acid | H2SO4 | Dibasic | Sulphate(VI) (SO42-) | Hydrogen sulphate(VI)  (HSO4-) |
| Sulphuric(IV)acid | H2SO3 | Dibasic | Sulphate(IV) (SO32-) | Hydrogen sulphate(IV)  (HSO3-) |
| Carbonic(IV)acid | H2CO3 | Dibasic | Carbonate(IV)(CO32-) | Hydrogen carbonate(IV)  (HCO3-) |
| Phosphoric(V) acid | H3PO4 | Tribasic | Phosphate(V)(PO43-) | Dihydrogen phosphate(V)  (H2PO42-)    Hydrogen diphosphate(V)  (HP2O42-) |

The table below show shows some examples of salts.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Base/alkali | Cation | Acid | Anion | Salt | Chemical name of salts |
| NaOH | +  Na | HCl | -  Cl | NaCl | Sodium(I)chloride |
| Mg(OH)  2 | 2+  Mg | H SO  2 4 | 2-  SO  4 | MgSO  4  Mg(HSO )  4 2 | Magnesium sulphate(VI) Magnesium hydrogen  sulphate(VI) |
| Pb(OH)  2 | 2+  Pb | HNO  3 | -  NO  3 | Pb(NO )  3 2 | Lead(II)nitrate(V) |
| Ba(OH)  2 | 2+  Ba | HNO  3 | -  NO  3 | Ba(NO )  3 2 | Barium(II)nitrate(V) |
| Ca(OH)  2 | 2+  Ba | H SO  2 4 | 2-  SO  4 | MgSO  4 | Calcium sulphate(VI) |
| NH OH  4 | +  NH  4 | H PO  3 4 | 3-  PO  4 | (NH ) PO  4 3 4  (NH ) HPO  4 2 4  NH H PO  4 2 4 | Ammonium phosphate(V)  Diammonium phosphate(V)  Ammonium diphosphate(V) |
| KOH | +  K | H PO  3 4 | 3-  PO  4 | K PO  3 4 | Potassium phosphate(V) |
| Al(OH)  3 | 3+  Al | H SO  2 4 | 2-  SO  4 | Al (SO )  2 4 2 | Aluminium(III)sulphate(VI) |
| Fe(OH)  2 | 2+  Fe | H SO  2 4 | 2-  SO  4 | FeSO  4 | Iron(II)sulphate(VI) |
| Fe(OH)  3 | 3+  Fe | H SO  2 4 | 2-  SO  4 | Fe (SO )  2 4 2 | Iron(III)sulphate(VI) |

(d) Some salts undergo **hygroscopy**, **deliquescence** and **efflorescence**.

(i) Hygroscopic salts /compounds are those that absorb water from the atmosphere but do not form a solution.

Some salts which are hygroscopic include anhydrous copper(II)sulphate(VI), anhydrous cobalt(II)chloride, potassium nitrate(V) common table salt.

(ii)Deliquescent salts /compounds are those that absorb water from the atmosphere and form a solution.

Some salts which are deliquescent include: Sodium nitrate(V),Calcium chloride, Sodium hydroxide, Iron(II)chloride, Magnesium chloride.

(iii)Efflorescent salts/compounds are those that lose their water of crystallization to the atmosphere.

Some salts which effloresces include: sodium carbonate decahydrate, Iron(II)sulphate(VI)heptahydrate, sodium sulphate (VI)decahydrate.

(e)Some salts contain water of crystallization.They are hydrated.Others do not contain water of crystallization. They are anhydrous.

Table showing some hydrated salts.

|  |  |
| --- | --- |
| Name of hydrated salt | Chemical formula |
| Copper(II)sulphate(VI)pentahydrate | CuSO4.5H2O |
| Aluminium(III)sulphate(VI)hexahydrate | Al2 (SO4) 3.6H2O |
| Zinc(II)sulphate(VI)heptahydrate | ZnSO4.7H2O |
| Iron(II)sulphate(VI)heptahydrate | FeSO4.7H2O |
| Calcium(II)sulphate(VI)heptahydrate | CaSO4.7H2O |
| Magnesium(II)sulphate(VI)heptahydrate | MgSO4.7H2O |
| Sodium sulphate(VI)decahydrate | Na2SO4.10H2O |
| Sodium carbonate(IV)decahydrate | Na2CO3.10H2O |
| Potassium carbonate(IV)decahydrate | K2CO3.10H2O |
| Potassium sulphate(VI)decahydrate | K2SO4.10H2O |

(f)Some salts exist as a simple salt while some as complex salts. Below are some complex salts.

Table of some complex salts

|  |  |  |
| --- | --- | --- |
| Name of complex salt | Chemical formula | Colour of the complex salt |
| Tetraamminecopper(II)sulphate(VI) | Cu(NH3) 4 SO4 H2O | Royal/deep blue solution |
| Tetraamminezinc(II)nitrate(V) | Zn(NH3) 4 (NO3 )2 | Colourless solution |
| Tetraamminecopper(II) nitrate(V) | Cu(NH3) 4 (NO3 )2 | Royal/deep blue solution |
| Tetraamminezinc(II)sulphate(VI) | Zn(NH3) 4 SO4 | Colourless solution |

(g)Some salts exist as two salts in one. They are called **double salts**.

Table of some double salts

|  |  |
| --- | --- |
| Name of double salts | Chemical formula |
| Trona(sodium sesquicarbonate) | Na2CO3 NaHCO3.2H2O |
| Ammonium iron(II)sulphate(VI) | FeSO4(NH4) 2SO4.2H2O |
| Ammonium aluminium(III)sulphate(VI) | Al2(SO4) 3(NH4) 2SO4.H2O |

(h)Some salts dissolve in water to form a **solution**. They are said to be soluble. Others do not dissolve in water. They form a **suspension/precipitate** in water.

Table of solubility of salts

Soluble salts Insoluble salts

All nitrate(V)salts

All sulphate(VI)/SO42- salts **except**  Barium(II) sulphate(VI)/BaSO4

Calcium(II) sulphate(VI)/CaSO4

Lead(II) sulphate(VI)/PbSO4

All sulphate(IV)/SO32- salts **except**  Barium(II) sulphate(IV)/BaSO3

Calcium(II) sulphate(IV)/CaSO3

Lead(II) sulphate(IV)/PbSO3

All chlorides/Cl- **except**  Silver chloride/AgCl

Lead(II)chloride/PbCl2(dissolves in hot water)

All phosphate(V)/PO43-

All sodium,potassium and ammonium salts

All hydrogen carbonates/HCO3-

All hydrogen sulphate(VI)/ HSO4-

Sodium carbonate/Na2CO3, **except**  All carbonates potassium carbonate/ K2CO3, ammonium carbonate (NH4) 2CO3

All alkalis(KOH,NaOH, NH4OH) **except** All bases

13 Salts can be prepared in a school laboratory by a method that uses its solubility in water.

(a) Soluble salts may be prepared by using any of the following methods: **(i)Direct displacement/reaction of a metal with an acid**.

By reacting a metal higher in the reactivity series than hydrogen with a dilute acid,a salt is formed and hydrogen gas is evolved.

Excess of the metal must be used to ensure all the acid has reacted.

When effervescence/bubbling /fizzing has stopped ,excess metal is filtered.

The filtrate is heated to concentrate then allowed to crystallize.

Washing with distilled water then drying between filter papers produces a sample crystal of the salt. i.e.

M(s) + H2X -> MX(aq) + H2(g)

|  |  |  |
| --- | --- | --- |
| Examples |  |  |
|  | Mg(s) + H2SO4(aq) | -> MgSO4 (aq) + H2(g) |
|  | Zn(s) + H2SO4(aq) | -> ZnSO4 (aq) + H2(g) |
|  | Pb(s) + 2HNO3(aq) | -> Pb(NO3) 2(aq) + H2(g) |
|  | Ca(s) + 2HNO3(aq) | -> Ca(NO3) 2(aq) + H2(g) |
|  | Mg(s) + 2HNO3(aq) | -> Mg(NO3) 2(aq) + H2(g) |
|  | Mg(s) + 2HCl(aq) | -> MgCl 2(aq) + H2(g) |
|  | Zn(s) + 2HCl(aq) | -> ZnCl 2(aq) + H2(g) |

**(ii)Reaction of an insoluble base with an acid**

By adding an insoluble base (oxide/hydroxide )to a dilute acid until **no more dissolves**, in the acid,a salt and water are formed. Excess of the base is filtered off. The filtrate is heated to concentrate ,allowed to crystallize then washed with distilled water before drying between filter papers e.g.

PbO(s) + 2HNO3(aq) -> Pb(NO3) 2(aq) + H2O (l)

Pb(OH)2(s) + 2HNO3(aq) -> Pb(NO3) 2(aq) + 2H2O (l)

CaO (s) + 2HNO3(aq) -> Ca(NO3) 2(aq) + H2O (l)

MgO (s) + 2HNO3(aq) -> Mg(NO3) 2(aq) + H2O (l)

MgO (s) + 2HCl(aq) -> MgCl 2(aq) + H2O (l)

ZnO (s) + 2HCl(aq) -> ZnCl 2(aq) + H2O (l)

Zn(OH)2(s) + 2HNO3(aq) -> Zn(NO3) 2(aq) + 2H2O (l)

CuO (s) + 2HCl(aq) -> CuCl 2(aq) + H2O (l)

CuO (s) + H2SO4(aq) -> CuSO4(aq) + H2O (l)

Ag2O(s) + 2HNO3(aq) -> 2AgNO3(aq) + H2O (l)

Na2O(s) + 2HNO3(aq) -> 2NaNO3(aq) + H2O (l)

**(iii)reaction of insoluble /soluble carbonate /hydrogen carbonate with an acid.** By adding an excess of a soluble /insoluble carbonate or hydrogen carbonate to adilute acid, effervescence /fizzing/bubbling out of carbon(IV)oxide gas shows the reaction is taking place. When effervescence /fizzing/bubbling out of the gas is over, excess of the insoluble carbonate is filtered off. The filtrate is heated to concentrate ,allowed to crystallize then washed with distilled water before drying between filter paper papers e.g.

PbCO3 (s) + 2HNO3(aq) -> Pb(NO3) 2(aq) + H2O (l)+ CO2(g)

ZnCO3 (s) + 2HNO3(aq) -> Zn(NO3) 2(aq) + H2O (l)+ CO2(g)

CaCO3 (s) + 2HNO3(aq) -> Ca(NO3) 2(aq) + H2O (l)+ CO2(g)

MgCO3 (s) + H2SO4(aq) -> MgSO4(aq) + H2O (l)+ CO2(g)

Cu CO3 (s) + H2SO4(aq) -> CuSO4(aq) + H2O (l) + CO2(g)

Ag2CO3 (s) + 2HNO3(aq) -> 2AgNO3(aq) + H2O (l) + CO2(g)

Na2CO3 (s) + 2HNO3(aq) -> 2NaNO3(aq) + H2O (l) + CO2(g)

K2CO3 (s) + 2HCl(aq) -> 2KCl(aq) + H2O (l) + CO2(g)

NaHCO3 (s) + HNO3(aq) -> NaNO3(aq) + H2O (l) + CO2(g)

KHCO3 (s) + HCl(aq) -> KCl(aq) + H2O (l) + CO2(g)

**(iv)neutralization/reaction of soluble base/alkali with dilute acid**

By adding an acid to a burette into a known volume of an alkali with 2-3 drops of an indicator, the colour of the indicator changes when the acid has completely reacted with an alkali at the **end point.** The procedure is then repeated without the indicator .The solution mixture is then heated to concentrate , allowed to crystallize ,washed with distilled water before drying with filter papers. e.g.

|  |  |  |
| --- | --- | --- |
|  | NaOH (aq) + HNO3(aq) | -> NaNO3(aq) + H2O (l) |
|  | KOH (aq) + HNO3(aq) | -> KNO3(aq) + H2O (l) |
|  | KOH (aq) + HCl(aq) | -> KCl(aq) + H2O (l) |
|  | 2KOH (aq) + H2SO4(aq) | -> K2SO4(aq) + 2H2O (l) |
|  | 2 NH4OH (aq) + H2SO4(aq) | -> (NH4)2SO4(aq) + 2H2O (l) |
|  | NH4OH (aq) + HNO3(aq) | -> NH4NO3(aq) + H2O (l) |

**(iv)direct synthesis/combination.**

When a metal **burn** in a gas jar containing a non metal , the two directly combine to form a salt. e.g.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| 2Na(s) | + | Cl2(g) | -> | 2NaCl(s) |
| 2K(s) | + | Cl2(g) | -> | 2KCl(s) |
| Mg(s) | + | Cl2(g) | -> | Mg Cl2 (s) |
| Ca(s) | + | Cl2(g) | -> | Ca Cl2 (s) |

Some salts once formed undergo **sublimation** and **hydrolysis**. Care should be taken to avoid water/moisture into the reaction flask during their preparation.Such salts include aluminium(III)chloride(AlCl3) and iron (III)chloride(FeCl3)

1. Heated aluminium foil reacts with chlorine to form aluminium(III)chloride that sublimes away from the source of heating then deposited as solid again

2Al(s) + 3Cl2(g) -> 2AlCl3 (**s/g**)

Once formed aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

AlCl3(s)+ 3H2 O(g) -> Al(OH)3 (**aq**) + 3HCl(g)

1. Heated iron filings reacts with chlorine to form iron(**III**)chloride that sublimes away from the source of heating then deposited as solid again

2Fe(s) + 3Cl2(g) -> 2FeCl3 (**s/g**)

Once formed , aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

FeCl3(s)+ 3H2 O(g) -> Fe(OH)3 (**aq**) + 3HCl(g)

(b)**Insoluble** salts can be prepared by reacting **two** suitable **soluble** salts to form **one soluble** and **one insoluble**. This is called **double decomposition** or **precipitation**. The mixture is filtered and the **residue** is washed with distilled water then dried.

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Cu**SO4(aq) + Na2**CO3** (aq) | -> | **CuCO3** (**s**) + Na2 SO4(aq) |
|  | **Ba**Cl2(aq) + K2**SO4** (aq) | -> | **BaSO4** (**s**) + 2KCl (aq) |
|  | **Pb(**NO3)2(aq) + K2**SO4** (aq) | -> | **PbSO4** (**s**) + 2KNO3 (aq) |
|  | 2**Ag**NO3(aq) + Mg**Cl**2 (aq) | -> | 2**AgCl**(**s**) + Mg(NO3)2 (aq) |
|  | **Pb(**NO3)2(aq) + (NH4**)** 2**SO4** (aq) | -> | **PbSO4** (**s**) + 2NH4NO 3(aq) |
|  | **Ba**Cl2(aq) + K2**SO3** (aq) | -> | **BaSO3** (**s**) + 2KCl (aq) |

14. Salts may lose their water of crystallization , decompose ,melt or sublime on heating on a Bunsen burner flame.

The following shows the behavior of some salts on heating gently /or strongly in a laboratory school burner:

**(a)effect of heat on chlorides**

All chlorides have very high melting and boiling points and therefore are not affected by laboratory heating except ammonium chloride. Ammonium chloride **sublimes** on **gentle** heating. It **dissociate** into the constituent **ammonia** and **hydrogen chloride** gases on strong heating.

NH4Cl(s) NH4Cl(g) NH3(g) + HCl(g)

(sublimation) (dissociation)

**(b)effect of heat on nitrate(V)**

(i) Potassium nitrate(V)/KNO3 and sodium nitrate(V)/NaNO3 decompose on heating to form Potassium nitrate(III)/KNO2 and sodium nitrate(III)/NaNO2 and producing Oxygen gas in each case.

2KNO**3** (s) -> 2KNO2(s) + O2(g)

2NaNO**3** (s) -> 2NaNO2(s) + O2(g)

(ii)Heavy metal nitrates(V) salts decompose on heating to form the oxide and a mixture of brown acidic nitrogen(IV)oxide and oxygen gases. e.g.

|  |  |
| --- | --- |
| 2Ca(NO3)2 (s) | -> 2CaO(s) + 4NO2(g) + O2(g) |
| 2Mg(NO3)2(s) | -> 2MgO(s) + 4NO2(g) + O2(g) |
| 2Zn(NO3)2(s) | -> 2ZnO(s) + 4NO2(g) + O2(g) |
| 2Pb(NO3)2(s) | -> 2PbO(s) + 4NO2(g) + O2(g) |
| 2Cu(NO3)2(s) | -> 2CuO(s) + 4NO2(g) + O2(g) |
| 2Fe(NO3)2(s) | -> 2FeO(s) + 4NO2(g) + O2(g) |

(iii)Silver(I)nitrate(V) and mercury(II) nitrate(V) are lowest in the reactivity series. They decompose on heating to form the **metal**(silver and mercury)and the Nitrogen(IV)oxide and oxygen gas. i.e.

2AgNO3(s) -> 2Ag (s) + 2NO2(g) + O2(g)

2Hg(NO3)2 (s) -> 2Hg (s) + 4NO2(g) + O2(g)

(iv)Ammonium nitrate(V) and Ammonium nitrate(III) decompose on heating to Nitrogen(I)oxide(relights/rekindles glowing splint) and nitrogen gas respectively.Water is also formed.i.e.

NH4NO3(s) -> N2O (g) + H2O(l)

NH4NO2(s) -> N2 (g) + H2O(l)

**(c) effect of heat on nitrate(V)**

Only Iron(II)sulphate(VI), Iron(III)sulphate(VI) and copper(II)sulphate(VI) decompose on heating. They form the **oxide**, and produce highly acidic fumes of acidic **sulphur(IV)oxide** gas.

2FeSO4 (s) -> Fe2O3(s) + SO3(g) + SO2(g)

Fe2(SO4) 3(s) -> Fe2O3(s) + SO3(g)

CuSO4 (s) -> CuO(s) + SO3(g)

**(d) effect of heat on carbonates(IV) and hydrogen carbonate(IV).** (i)Sodium carbonate(IV)and potassium carbonate(IV)**do not decompose** on heating.

(ii)Heavy metal nitrate(IV)salts decompose on heating to form the **oxide** and produce **carbon(IV)oxide** gas.Carbon (IV)oxide gas forms a white precipitate when bubbled in lime water. The white precipitate dissolves if the gas is in excess.

|  |  |
| --- | --- |
| e.g. CuCO3 (s) | -> CuO(s) + CO2(g) |
| CaCO3 (s) | -> CaO(s) + CO2(g) |
| PbCO3 (s) | -> PbO(s) + CO2(g) |
| FeCO3 (s) | -> FeO(s) + CO2(g) |
| ZnCO3 (s) | -> ZnO(s) + CO2(g) |

(iii)Sodium hydrogen carbonate(IV) and Potassium hydrogen carbonate(IV)decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i.e.

2NaHCO 3(s) -> Na2CO3(s) + CO2(g) + H2O(l)

2KHCO 3(s) -> K2CO3(s) + CO2(g) + H2O(l) (iii) Calcium hydrogen carbonate (IV) and Magnesium hydrogen carbonate(IV) decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i. e.

Ca(HCO3) 2(**aq**) -> CaCO3(s) + CO2(g) + H2O(l)

Mg(HCO3) 2(**aq**) -> MgCO3(s) + CO2(g) + H2O(l)

**CARBON**

Carbon is an element in Group IV(Group 4)of the Periodic table .It has atomic number 6 and electronic configuration 2:4 and thus has four valence electrons(tetravalent).It does not easily ionize but forms strong covalent bonds with other elements including itself.

**(a)Occurrence**

Carbon mainly naturally occurs as:

(i)allotropes of carbon i.e graphite, diamond and fullerenes.

(ii)amorphous carbon in coal, peat ,charcoal and coke.

(iii)carbon(IV)oxide gas accounting 0.03% by volume of normal air in the atmosphere.

**(b)Allotropes of Carbon**

Carbon naturally occur in two main crystalline allotropic forms, carbon-graphite and carbon-diamond

|  |  |
| --- | --- |
| Carbon-diamond | Carbon-graphite |
| Shiny crystalline solid | Black/dull crystalline solid |
| Has a very high melting/boiling point because it has a very closely packed giant tetrahedral structure joined by strong covalent bonds | Has a high melting/boiling point because it has a very closely packed giant hexagonal planar structure joined by strong covalent bonds |
| Has very high density(Hardest known natural substance) | Soft |
| Abrassive | Slippery |
| Poor electrical conductor because it has no free delocalized electrons | Good electrical conductor because it has free 4th valency delocalized electrons |
| Is used in making Jewels, drilling and cutting metals | Used in making Lead-pencils,electrodes in batteries and as a lubricant |
| Has giant tetrahedral structure | Has giant hexagonal planar structure |

c)Properties of Carbon (i)Physical properties of carbon

Carbon occur widely and naturally as a black solid

It is insoluble in water but soluble in carbon disulphide and organic solvents.

It is a poor electrical and thermal conductor.

(ii)Chemical properties of carbon

**I. Burning**

Experiment

Introduce a small piece of charcoal on a Bunsen flame then lower it into a gas jar containing Oxygen gas. Put three drops of water. Swirl. Test the solution with blue and red litmus papers.

Observation

-Carbon chars then burns with a blue flame

-Colourless and odourless gas produced

-Solution formed turn blue litmus paper faint red. Red litmus paper remains red.

Explanation

Carbon burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon burns in limited supply of air with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon (IV) oxide gas dissolve in water to form weak acidic solution of Carbonic (IV)acid.

Chemical Equation

C(s) + O2(g) -> CO2(g) (in excess air)

2C(s) + O2(g) -> 2CO(g) (in limited air)

CO2(g) + H2O (l) -> H2CO3 (aq) (very weak acid)

**II. Reducing agent**

Experiment

Mix thoroughly equal amounts of powdered charcoal and copper (II)oxide into a crucible. Heat strongly.

Observation

Colour change from black to brown

Explanation

Carbon is a reducing agent. For ages it has been used to reducing metal oxide ores to metal, itself oxidized to carbon(IV)oxide gas. Carbon reduces black copper(II)oxide to brown copper metal

Chemical Equation

|  |  |  |  |
| --- | --- | --- | --- |
| 2CuO(s) + C(s) -> | 2Cu(s) | + | CO2(g) |
| (black) | (brown) |  |  |
| 2PbO(s) + C(s) -> | 2Pb(s) | + | CO2(g) |
| (brown when hot/ yellow when cool) | (grey) |  |  |
| 2ZnO(s) + C(s) -> | 2Zn(s) | + | CO2(g) |
| (yellow when hot/ | (grey) |  |  |

white when cool)

|  |  |  |
| --- | --- | --- |
| Fe2O3(s) + 3C(s) | -> 2Fe(s) + | 3CO2(g) |
| (brown when hot/cool | (grey) |  |
| Fe3O4 (s) + 4C(s) | -> 3Fe(s) + | 4CO2(g) |
| (brown when hot/cool | (grey) |  |

**B: COMPOUNDS OF CARBON**

The following are the main compounds of Carbon

(i)Carbon(IV)Oxide(CO2)

(ii)Carbon(II)Oxide(CO)

(iii)Carbonate(IV) (CO32-)and hydrogen carbonate(IV(HCO3-)

(iv)Sodium carbonate(Na2CO3)

(i) Carbon(IV)Oxide (CO2)

**(a)Occurrence**

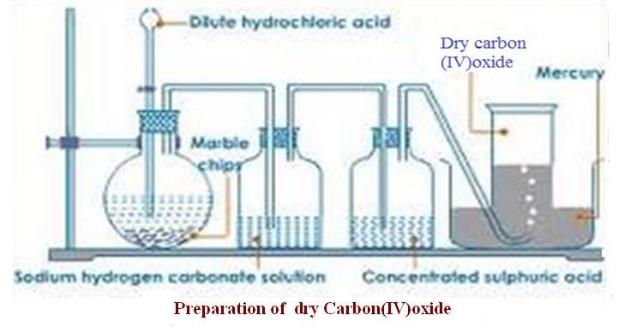
Carbon(IV)oxide is found:

-in the air /atmosphere as 0.03% by volume.

-a solid carbon(IV)oxide mineral in Esageri near Eldame Ravine and Kerita near Limuru in Kenya.

**(b)School Laboratory preparation**

In the school laboratory carbon(IV)oxide can be prepared in the school laboratory from the reaction of marble chips(CaCO3)or sodium hydrogen carbonate(NaHCO3) with dilute hydrochloric acid.



**(c)Properties of carbon(IV)oxide gas(Questions)**

**1.Write the equation for the reaction for the school laboratory preparation of carbon (IV)oxide gas.**

Any carbonate reacted with dilute hydrochloric acid should be able to generate carbon (IV)oxide gas.

Chemical equations

CaCO3(s) + 2HCl(aq) -> CaCO3 (aq) + H2O(l) + CO2 (g)

ZnCO3(s) + 2HCl(aq) -> ZnCO3 (aq) + H2O(l) + CO2 (g)

MgCO3(s) + 2HCl(aq) -> MgCO3 (aq) + H2O(l) + CO2 (g)

CuCO3(s) + 2HCl(aq) -> CuCO3 (aq) + H2O(l) + CO2 (g)

NaHCO3(s) + HCl(aq) -> Na2CO3 (aq) + H2O(l) + CO2 (g)

KHCO3(s) + HCl(aq) -> K2CO3 (aq) + H2O(l) + CO2 (g)

**2.What method of gas collection is used in preparation of Carbon(IV)oxide gas. Explain.**

Downward delivery /upward displacement of air/over mercury Carbon(IV)oxide gas is about 1½ times denser than air.

**3.What is the purpose of :**

**(a)water?**

To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction.

**(b)sodium hydrogen carbonate?**

To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction and by reacting with the acid to produce more carbon (IV)oxide gas .

Chemical equation

NaHCO3(s) + HCl(aq) -> Na2CO3 (aq) + H2O(l) + CO2 (g)

**(c)concentrated sulphuric(VI)acid?**

To dry the gas/as a drying agent

**4.Describe the smell of carbon(IV)oxide gas**

Colourless and odourless

**5. Effect on lime water.**

Experiment

Bubbled carbon(IV)oxide gas into a test tube containing lime water for about three minutes

Observation

White precipitate is formed.

White precipitate dissolved when excess carbon(IV)oxide gas is bubbled .

Explanation

Carbon(IV)oxide gas reacts with lime water(Ca(OH)2) to form an insoluble white precipitate of calcium carbonate. Calcium carbonate reacts with more Carbon(IV) oxide gas to form soluble Calcium hydrogen carbonate.

Chemical equation

Ca(OH)2(aq) + CO2 (g) -> CaCO3 (s) + H2O(l)

CaCO3 (aq) + H2O(l) + CO2 (g) -> Ca(HCO3) 2 (aq)

**6. Effects on burning Magnesium ribbon**

Experiment

Lower a piece of burning magnesium ribbon into a gas jar containing carbon (IV)oxide gas.

Observation

The ribbon continues to burn with difficulty White ash/solid is formed.

Black speck/solid/particles formed on the side of gas jar.

Explanation

Carbon(IV)oxide gas does not support combustion/burning.Magnesium burn to produce/release enough heat energy to decompose Carbon(IV) oxide gas to carbon and oxygen.Magnesium continues to burn in Oxygen forming white Magnesium Oxide solid/ash.Black speck/particle of carbon/charcoal residue forms on the sides of reaction flask. During the reaction Carbon(IV) oxide is reduced(Oxidizing agent)to carbon while Magnesium is Oxidized to Magnesium Oxide.

Chemical equation

2Mg(s) + CO2 (g) -> C (s) + 2MgO(l)

**7. Dry and wet litmus papers were separately put in a gas jar containing dry carbon (IV)oxide gas. State and explain the observations made.**

Observation

Blue dry litmus paper remain blue

Red dry litmus paper remain Red

Blue wet/damp/moist litmus paper turn red

Red wet/damp/moist litmus paper remain red

Explanation

Dry Carbon (IV) oxide gas is a molecular compound that does not dissociate/ionize to release **H+** and thus has no effect on litmus papers.

Wet/damp/moist litmus papers contains water that dissolves/react with dry carbon (IV) oxide gas to form the weak solution of carbonic (IV) acid(H2CO3).

Carbonic (IV) acid dissociate/ionizes to a few /little free **H+** and **CO32-.** The few H+ (aq) ions are responsible for turning blue litmus paper to faint red showing the gas is very weakly acidic.

Chemical equation

H2CO3(aq) -> 2H+ (aq) + CO32-(aq)

**8. Explain why Carbon (IV)oxide cannot be prepared from the reaction of: (i) marble chips with dilute sulphuric(VI)acid.**

Explanation

Reaction forms insoluble calcium sulphate(VI)that cover/coat unreacted marble chips stopping further reaction

Chemical equation

CaCO3(s) + H2SO4 (aq) -> CaSO4 (**s**) + H2O(l) + CO2 (g)

PbCO3(s) + H2SO4 (aq) -> PbSO4 (**s**) + H2O(l) + CO2 (g)

BaCO3(s) + H2SO4 (aq) -> BaSO4 (**s**) + H2O(l) + CO2 (g) **(ii) Lead(II)carbonate with dilute Hydrochloric acid.**

Reaction forms insoluble Lead(II)Chloride that cover/coat unreacted Lead(II) carbonate stopping further reaction unless the reaction mixture is heated.

Lead(II)Chloride is soluble in hot water.

Chemical equation

PbCO3(s) + 2HCl (aq) -> PbCl2 (**s**) + H2O(l) + CO2 (g)

**9. Describe the test for the presence of Carbon (IV)oxide.**

Using burning splint

Lower a burning splint into a gas jar suspected to contain Carbon (IV)oxide gas.The burning splint is extinguished.

Using Lime water.

Bubble the gas suspected to be Carbon (IV)oxide gas.A white precipitate that dissolve in excess bubbling is formed.

Chemical equation

Ca(OH)2(aq) + CO2 (g) -> CaCO3 (s) + H2O(l)

CaCO3 (aq) + H2O(l) + CO2 (g) -> Ca(HCO3) 2 (aq)

**10.State three main uses of** **Carbon (IV)oxide gas**

(i)In the Solvay process for the manufacture of soda ash/sodium carbonate

(ii)In preservation of aerated drinks

(iii)As fire extinguisher because it does not support combustion and is denser than air.

(iv)In manufacture of Baking powder.

**(ii) Carbon(II)Oxide (CO)**

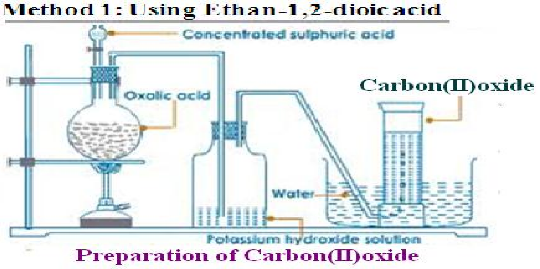
**(a)Occurrence**

Carbon(II)oxide is found is found from incomplete combustion of fuels like petrol charcoal, liquefied Petroleum Gas/LPG.

**(b)School Laboratory preparation**

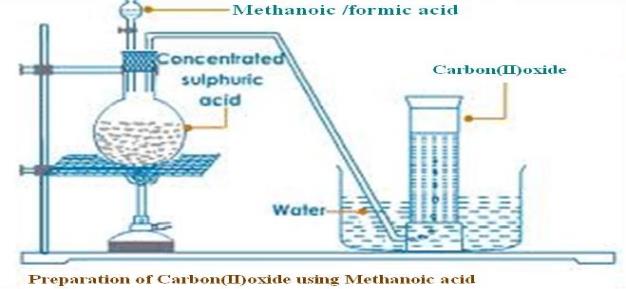
In the school laboratory carbon(II)oxide can be prepared from dehydration of methanoic acid/Formic acid(HCOOH) or Ethan-1,2-dioic acid/Oxalic acid(HOOCCOOH) using concentrated sulphuric(VI) acid. Heating is necessary.

METHOD 1:Preparation of Carbon (IV)Oxide from dehydration of Oxalic/ethan-1,2-dioic acid



METHOD 2:

Preparation of Carbon (IV)Oxide from dehydration of Formic/Methanoic acid



**(c)Properties of Carbon (II)Oxide(Questions)**

**1.Write the equation for the reaction for the preparation of carbon(II)oxide using;**

(i)Method 1;

Chemical equation

HOOCCOOH(s) –Conc.H2SO4--> CO(g) + CO2 (g) + H2O(l) H2C2O4(s) –Conc.H2SO4--> CO(g) + CO2 (g) + H2O(l)

(ii)Method 2;

Chemical equation

HCOOH(s) –Conc.H2SO4--> CO(g) + H2O(l)

H2CO2(s) –Conc.H2SO4--> CO(g) + H2O(l)

**2.What method of gas collection is used during the preparation of carbon (II) oxide.**

Over water because the gas is insoluble in water.

Downward delivery because the gas is 1 ½ times denser than air .

**3.What is the purpose of :**

1. **Potassium hydroxide/sodium hydroxide in Method 1**

To absorb/ remove carbon (II) oxide produced during the reaction.

2KOH (aq) + CO2 (g) -> K2CO3 (s) + H2O(l)

2NaOH (aq) + CO2 (g) -> Na2CO3 (s) + H2O(l)

1. **Concentrated sulphuric(VI)acid in Method 1 and 2.**

Dehydrating agent –removes the element of water (Hydrogen and Oxygen in ratio 2:1) present in both methanoic and ethan-1,2-dioic acid.

1. **Describe the smell of carbon(II)oxide.**

Colourless and odourless.

1. **State and explain the observation made when carbon(IV)oxide is bubbled in lime water for a long time.**

No white precipitate is formed.

1. **Dry and wet/moist/damp litmus papers were separately put in a gas jar containing dry carbon(IV)oxide gas. State and explain the observations made.**

Observation

-blue dry litmus paper remains blue

-red dry litmus paper remains red

* + wet/moist/damp blue litmus paper remains blue
  + wet/moist/damp red litmus paper remains red

Explanation

Carbon(II)oxide gas is a molecular compound that does not dissociate /ionize to release H+ ions and thus has no effect on litmus papers. Carbon(II)oxide gas is therefore a **neutral** gas.

1. **Carbon (II)oxide gas was ignited at the end of a generator as below.**

Flame K

Dry carbon(II)oxide

**(i)State the observations made in flame K.** Gas burns with a blue flame

**(ii)Write the equation for the reaction taking place at flame K.**

**2**CO(g) + O2 (g) -> 2CO2 (g)

1. **Carbon(II)oxide is a reducing agent. Explain**

Experiment

Pass carbon(II)oxide through glass tube containing copper (II)oxide. Ignite any excess poisonous carbon(II)oxide.

Observation

Colour change from black to brown. Excess carbon (II)oxide burn with a blue flame. Explanation

Carbon is a reducing agent. It is used to reduce metal oxide ores to metal, itself oxidized to carbon(IV)oxide gas. Carbon(II)Oxide reduces black copper(II)oxide to brown copper metal

Chemical Equation

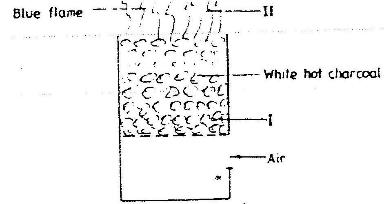
|  |  |  |
| --- | --- | --- |
| CuO(s) + CO(g) | -> Cu(s) + | CO2(g) |
| (black) | (brown) |  |
| PbO(s) + CO(g) | -> Pb(s) + | CO2(g) |
| (brown when hot/ yellow when cool) | (grey) |  |
| ZnO(s) + CO(g) | -> Zn(s) + | CO2(g) |
| (yellow when hot/ white when cool) | (grey) |  |
| Fe2O3(s) + 3CO(s) | -> 2Fe(s) + | 3CO2(g) |
| (brown when hot/cool | (grey) |  |
| Fe3O4 (s) + 4CO(g) | -> 3Fe(s) + | 4CO2(g) |
| (brown when hot/cool | (grey) |  |

These reaction are used during the extraction of many metals from their ore.

**9. Carbon (II) oxide is a pollutant. Explain.**

Carbon(II)oxide is highly poisonous/toxic.It preferentially combine with haemoglobin to form stable carboxyhaemoglobin in the blood instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood causing nausea , coma then death.

**10.The diagram below show a burning charcoal stove/burner/jiko. Use it to answer the questions that follow.**



**Explain the changes that take place in the burner**

Explanation

Charcoal stove has air holes through which air enters. Air oxidizes carbon to carbon(IV)oxide gas at region I. This reaction is exothermic(-∆H) producing more heat.

Chemical equation

C(s) + O2(g) -> CO2(g)

Carbon(IV)oxide gas formed rises up to meet more charcoal which reduces it to Carbon(II)oxide gas.

Chemical equation

2CO2 (g) + O2(g) -> 2CO (g)

At the top of burner in region II, Carbon (II)oxide gas is further oxidized to Carbon(IV)oxide gas if there is plenty of air but escape if the air is limited poisoning the living things around.

Chemical equation

2CO (g) + O2(g) -> 2CO2 (g)

(excess air)

**11.Describe the test for the presence of carbon(II)oxide gas.**

Experiment

Burn/Ignite the pure sample of the gas. Pass/Bubble the products into lime water/Calcium hydroxide . Observation

Colourless gas burns with a blue flame. A white precipitate is formed that dissolve on further bubbling of the products.

Chemical equation

2CO (g) + O2(g) -> 2CO2 (g) (gas burns with blue flame)

Chemical equation

Ca(OH) 2 (aq) + CO2 (g) -> CaCO3 (s) + H2O(l)

Chemical equation

CO2 (g) + CaCO3 (s) + H2O(l) -> Ca(HCO3) 2 (aq)

**12. State the main uses of carbon (II)oxide gas.**

(i) As a fuel /water gas

(ii)As a reducing agent in the blast furnace for extracting iron from iron ore(Magnetite/Haematite)

(iii)As a reducing agent in extraction of Zinc from Zinc ore/Zinc blende

(iv) As a reducing agent in extraction of Lead from Lead ore/Galena (v) As a reducing agent in extraction of Copper from Copper iron sulphide/Copper pyrites.

**(iii)Carbonate(IV) (CO32-)and hydrogen carbonate(IV(HCO3-)**

1.Carbonate (IV) (CO32-) are normal salts derived from carbonic(IV)acid (H2CO3) and hydrogen carbonate (IV) (HCO3-) are acid salts derived from carbonic(IV)acid. Carbonic(IV)acid(H2CO3) is formed when carbon(IV)oxide gas is bubbled in water. It is a dibasic acid with two ionizable hydrogens.

H2CO3(aq) ->2H+(aq) + CO32-(aq)

H2CO3(aq) -> H+(aq) + HCO3 - (aq)

2.Carbonate (IV) (CO32-) are insoluble in water **except** Na2CO3 , K2CO3 and

(NH4)2CO3

3.Hydrogen carbonate (IV) (HCO3-) are soluble in water. Only five hydrogen carbonates exist. **NaHCO3 , KHCO3 ,NH4HCO3 Ca(HCO3)2 and Mg(HCO3)2** Ca(HCO3)2 and Mg(HCO3)2 exist only in aqueous solutions.

3.The following experiments show the effect of heat on Carbonate (IV) (CO32-) and Hydrogen carbonate (IV) (HCO3-) salts:Experiment

In a clean dry test tube place separately about 1.0 of the following:

Zinc(II)carbonate(IV), sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) ammonium carbonate(IV), potassium hydrogen carbonate(IV), Lead(II)carbonate(IV), Iron(II)carbonate(IV), and copper(II)carbonate(IV). Heat each portion gently the strongly. Test any gases produced with lime water.

Observation

(i)Colorless droplets form on the cooler parts of test tube in case of sodium carbonate(IV) and Potassium carbonate(IV).

(ii)White residue/solid left in case of sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) and potassium hydrogen carbonate(IV).

(iii)Colour changes from blue/green to black in case of copper(II)carbonate(IV).

(iv) Colour changes from green to brown/yellow in case of Iron (II)carbonate(IV). (v) Colour changes from white when cool to yellow when hot in case of Zinc (II) carbonate(IV).

(vi) Colour changes from yellow when cool to brown when hot in case of Lead (II) carbonate(IV).

(vii)Colourless gas produced that forms a white precipitate with lime water in all cases.

Explanation

1. Sodium carbonate(IV) and Potassium carbonate(IV) exist as hydrated salts with 10 molecules of water of crystallization that condenses and collects on cooler parts of test tube as a colourless liquid.

Chemical equation

Na2CO3 .10H2O(s) -> Na2CO3 (s) + 10H2O(l)

K2CO3 .10H2O(s) -> K2CO3 (s) + 10H2O(l)

1. Carbonate (IV) (CO32-) and Hydrogen carbonate (IV) (HCO3-) salts decompose on heating except Sodium carbonate(IV) and Potassium carbonate(IV). (a) Sodium hydrogen carbonate(IV) and Potassium hydrogen carbonate(IV) decompose on heating to form sodium carbonate(IV) and Potassium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

2NaHCO3 (s) -> Na2CO3 (s) + H2O(l) + CO2 (g)

(white) (white)

2KHCO3 (s) -> K2CO3 (s) + H2O(l) + CO2 (g)

(white) (white)

1. Calcium hydrogen carbonate(IV) and Magnesium hydrogen carbonate(IV) decompose on heating to form insoluble Calcium carbonate(IV) and Magnesium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

Ca(HCO3)2 (aq) -> CaCO3 (s) + H2O(l) + CO2 (g)

(Colourless solution) (white)

Mg(HCO3)2 (aq) -> MgCO3 (s) + H2O(l) + CO2 (g) (Colourless solution) (white)

1. Ammonium hydrogen carbonate(IV) decompose on heating to form ammonium carbonate(IV) .Water and carbon(IV)oxide gas are also produced.

Chemical equation

2NH4HCO3 (s) -> (NH4)2CO3 (s) + H2O(l) + CO2 (g)

(white) (white)

(d)All other carbonates decompose on heating to form the metal oxide and produce carbon(IV)oxide gas e.g.

Chemical equation

MgCO3 (s) -> MgO (s) + CO2 (g)

(white solid) (white solid)

Chemical equation

BaCO3 (s) -> BaO (s) + CO2 (g)

(white solid) (white solid)

Chemical equation

CaCO3 (s) -> CaO (s) + CO2 (g)

(white solid) (white solid)

Chemical equation

CuCO3 (s) -> CuO (s) + CO2 (g)

(blue/green solid) (black solid)

Chemical equation

ZnCO3 (s) -> ZnO (s) + CO2 (g) (white solid) (white solid when cool/

Yellow solid when hot)

Chemical equation

PbCO3 (s) -> PbO (s) + CO2 (g)

(white solid) (yellow solid when cool/

brown solid when hot)

4.The following experiments show the presence of Carbonate (IV) (CO32-) and Hydrogen carbonate (IV) (HCO3-) ions in sample of a salt:

(a)Using Lead(II) nitrate(V)

1. Using a portion of salt solution in a test tube .add four drops of Lead(II)nitrate(V)solution.Preserve.

|  |  |
| --- | --- |
| Observation | inference |
| White precipitate/ppt | CO32- ,SO32- ,SO42- ,Cl |

1. To the preserved solution ,add six drops of dilutte nitric(V)acid. Preserve.

|  |  |
| --- | --- |
| Observation | inference |
| White precipitate/ppt persists    White precipitate/ppt dissolves | SO42- ,Cl -    CO32- ,SO32 |

II. To the preserved sample( that forms a precipitate ),heat to boil.

|  |  |
| --- | --- |
| Observation | inference |
| White precipitate/ppt persists    White precipitate/ppt dissolves | SO42    Cl |

II. To the preserved sample( that do not form a precipitate ),add three drops of acidified potassium manganate(VII)/lime water

|  |  |
| --- | --- |
| Observation | inference |
| Effervescence/bubbles/fizzing colourless gas produced  Acidified KMnO4 decolorized/no white  precipitate on lime water    Effervescence/bubbles/fizzing colourless gas produced  Acidified KMnO4 not decolorized/ white precipitate on lime water | SO32-          CO32- |

Experiments/Observations:

**(b)Using Barium(II)nitrate(V)/ Barium(II)chloride**

I. To about 5cm3 of a salt solution in a test tube add four drops of Barium(II) nitrate (V) / Barium(II)chloride. Preserve.

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt | SO42- , SO32- , CO32- ions |
| II. To the preserved sample in (I) abov Preserve.    Observation 1 | e, add six drops of 2M nitric(V) acid . |
| Observation | Inference |
| White precipitate/ppt persists | SO42- , ions |
| Observation 2 |  |
| Observation | Inference |
| White precipitate/ppt dissolves | SO32- , CO32- , ions |

III.To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

|  |  |
| --- | --- |
| Observation | Inference |
| (i)acidified potassium manganate(VII)decolorized  (ii)Orange colour of acidified potassium dichromate(VI) turns to green | SO32- ions |

Observation 2

|  |  |
| --- | --- |
| Observation | Inference |
| (i)acidified potassium manganate(VII) not decolorized  (ii)Orange colour of acidified potassium dichromate(VI) does not turns to green | CO32- ions |

Explanations

**Using Lead(II)nitrate(V)**

(i)Lead(II)nitrate(V) solution reacts with chlorides(Cl-), Sulphate (VI) salts (SO42- ), Sulphate (IV)salts (SO32-) and carbonates(CO32-) to form the insoluble white precipitate of Lead(II)chloride, Lead(II)sulphate(VI), Lead(II) sulphate (IV) and Lead(II)carbonate(IV).

Chemical/ionic equation:

Pb2+(aq) + Cl- (aq) -> PbCl2(s)

Pb2+(aq) + SO42+ (aq) -> PbSO4 (s)

Pb2+(aq) + SO32+ (aq) -> PbSO3 (s)

Pb2+(aq) + CO32+ (aq) -> PbCO3 (s)

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

* Lead(II)chloride and Lead(II)sulphate(VI) do not react with the acid and thus their white precipitates remain/ persists.
* Lead(II) sulphate (IV) and Lead(II)carbonate(IV) reacts with the acid to form **soluble** Lead(II) nitrate (V) and produce/effervesces/fizzes/bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively. . Chemical/ionic equation:

|  |  |  |  |
| --- | --- | --- | --- |
| PbSO3 (s) + 2H+(aq) -> | H2 O (l) + Pb2+(aq) | + | SO2 (g) |
| PbCO3 (s) + 2H+(aq) -> | H2 O (l) + Pb2+(aq) | + | CO2 (g) |

(iii)When Lead(II)chloride and Lead(II)sulphate(VI) are heated/warmed; - Lead(II)chloride dissolves in hot water/on boiling(recrystallizes on cooling)

* Lead(II)sulphate(VI) do not dissolve in hot water thus its white precipitate persists/remains on heating/boiling.

(iv)When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

* **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green.

**Carbon(IV)oxide will not.**

Chemical equation:

5SO32-(aq)+ 2MnO4- (aq) +6H+(aq) -> 5SO42-(aq) + 2Mn2+(aq) + 3H2O(l)

(purple) (colourless)

3SO32-(aq)+ Cr2O72-(aq) +8H+(aq) -> 3SO42-(aq) + 2Cr3+(aq) + 4H2O(l)

(Orange) (green)

* **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.** Chemical equation:

Ca(OH)2(aq)+ CO2 (g) -> CaCO3(s) + H2O(l)

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

**Using Barium(II)nitrate(V)/ Barium(II)Chloride**

(i)Barium(II)nitrate(V) and/ or Barium(II)chlo**ride** solution reacts with Sulphate (VI) salts (SO42- ), Sulphate (IV)salts (SO32-) and carbonates(CO32-) to form the insoluble white precipitate of Barium(II)sulphate(VI), Barium(II) sulphate (IV) and Barium(II)carbonate(IV).

Chemical/ionic equation:

Ba2+(aq) + SO42+ (aq) -> BaSO4 (s)

Ba2+(aq) + SO32+ (aq) -> BaSO3 (s)

Ba2+(aq) + CO32+ (aq) -> BaCO3 (s)

(ii)When the insoluble precipitates are acidified with nitric(V) acid, - Barium (II)sulphate(VI) do not react with the acid and thus its white precipitates remain/ persists.

- Barium(II) sulphate (IV) and Barium(II)carbonate(IV) reacts with the acid to form **soluble** Barium(II) nitrate (V) and produce /effervesces /fizzes/ bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

BaSO3 (s) + 2H+(aq) -> H2 O (l) + Ba2+(aq) + SO2 (g)

BaCO3 (s) + 2H+(aq) -> H2 O (l) + Ba2+(aq) + CO2 (g)

(iii) When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

* **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green. **Carbon(IV)oxide will not.**

Chemical equation:

5SO32-(aq)+ 2MnO4- (aq) +6H+(aq) -> 5SO42-(aq) + 2Mn2+(aq) + 3H2O(l)

(purple) (colourless)

3SO32-(aq)+ Cr2O72-(aq) +8H+(aq) -> 3SO42-(aq) + 2Cr3+(aq) + 4H2O(l)

(Orange) (green)

* **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

Ca(OH)2(aq)+ CO2 (g) -> CaCO3(s) + H2O(l)

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

**(iii) Sodium carbonate(IV) (Na2CO3)**

**(a)Extraction of sodium carbonate from soda ash**

Sodium carbonate naturally occurs in Lake Magadi in Kenya as Trona.trona is the double salt ; sodium sesquicarbonate.NaHCO3 .Na2CO3 .H2O.It is formed from the volcanic activity that takes place in Lake Naivasha, Nakuru ,Bogoria and Elementeita .All these lakes drain into Lake Magadi through underground rivers.

Lake Magadi has no outlet.

Solubility of Trona decrease with increase in temperature.High temperature during

the day causes trona to naturally crystallize .It is mechanically scooped/dredged/dug and put in a furnace.

Inside the furnace, trona decompose into soda ash/sodium carbonate.

Chemical equation

2NaHCO3 .Na2CO3 .H2O (s) -> 3Na2CO3 (s) + 5H2O(l) + CO2 (g)

(trona) (soda ash)

Soda ash is then bagged and sold as Magadi soda.It is mainly used:

(i)in making glass to lower the melting point of raw materials (sand/SiO2 from 1650oC and CaO from 2500oC to around 1500oC)

(ii)in softening hard water

(iii)in the manufacture of soapless detergents.

(iv)Swimming pool “pH increaser”

Sodium chloride is also found dissolved in the lake. Solubility of sodium chloride decrease with decreases in temperature/ sodium chloride has lower solubility at lower temperatures. When temperatures decrease at night it crystallize out .The crystals are then mechanically dug/dredged /scooped then packed for sale as animal/cattle feeds and seasoning food.

**Summary flow diagram showing the extraction of Soda ash from Trona**

Sodium chloride and Trona

dissolved in the se

a

Natural

fractional crystallization

Crystals of Trona

(

**Day time**

)

Crystals of sodium

chloride(

**At night**

)

Dredging

/scooping/

digging

Crushing

Furnace

(

Heating

)

Carbon(IV) oxide

Soda ash

Bagging

NaCl(s)

Bagging

Na

2

CO

3

s

)

(

**b)The Solvay process for industrial manufacture of sodium carbonate(IV) (i)Raw materials.**

-**Brine** /Concentrated Sodium chloride from salty seas/lakes.

-**Ammonia** gas from Haber.

-**Limestone** /Calcium carbonate from chalk /limestone rich rocks. -**Water** from rivers/lakes.

**(ii)Chemical processes**

Ammonia gas is passed **up** to meet a **downward** flow of sodium chloride solution / brine to form **ammoniated** brine/**ammoniacal** brine **mixture** in the **ammoniated brine chamber**

The ammoniated brine mixture is then pumped up, atop the carbonator/ solvay tower.

In the carbonator/ solvay tower,ammoniated brine/ammoniacal brine mixture slowly trickle down to meet an upward flow of carbon(IV)oxide gas.

The carbonator is shelved /packed with quartz/broken glass to

(i) reduce the rate of flow of ammoniated brine/ammoniacal brine mixture.

(ii)increase surface area of the liquid mixture to ensure a lot of ammoniated brine/ammoniacal brine mixture react with carbon(IV)oxide gas.

Insoluble sodium hydrogen carbonate and soluble ammonium chloride are formed from the reaction. Chemical equation

CO2(g) + H2O(l) + NaCl (aq) + NH3(g) -> NaHCO3(s) + NH4Cl(aq)

The products are then filtered. **Insoluble** sodium hydrogen carbonate forms the **residue** while soluble ammonium chloride forms the **filtrate**.

Sodium hydrogen carbonate itself can be used:

1. as baking powder and preservation of some soft drinks.
2. as a buffer agent and antacid in animal feeds to improve fibre digestion. (iii) making dry chemical fire extinguishers.

In the Solvay process Sodium hydrogen carbonate is then heated to form Sodium carbonate/soda ash, water and carbon (IV) oxide gas.

Chemical equation

2NaHCO3 (s) -> Na2CO3(s) + CO2(g) + H2O(l)

Sodium carbonate is stored ready for use in:

1. during making glass/lowering the melting point of mixture of sand/SiO2

from 1650oC and CaO from 2500oC to around 1500oC

1. in softening hard water
2. in the manufacture of soapless detergents.
3. swimming pool “pH increaser”.

Water and carbon(IV)oxide gas are recycled back to the ammoniated brine/ammoniacal brine chamber.

More carbon(IV)oxide is produced in the kiln/furnace. Limestone is heated to decompose into Calcium oxide and carbon(IV)oxide.

Chemical equation

CaCO3 (s) -> CaO(s) + CO2(g)

Carbon(IV)oxide is recycled to the carbonator/solvay tower. Carbon (IV)oxide is added water in the **slaker** to form Calcium hydroxide. This process is called **slaking**.

Chemical equation

CaO(s) + H2O (l) -> Ca(OH)2 (aq)

Calcium hydroxide is mixed with ammonium chloride from the carbonator/solvay tower in the **ammonia regeneration chamber** to form Calcium chloride , water and more ammonia gas.

Chemical equation

Ca(OH)2 (aq) +2NH4Cl (aq) -> CaCl2(s) + 2NH3(g) + H2O(l)

NH3(g) and H2O(l) are recycled.

Calcium chloride may be used:

(i)as drying agent in the school laboratory during gas preparation

(except ammonia gas)

(ii)to lower the melting point of solid sodium chloride / rock salt salts during the Downs process for industrial extraction of sodium metal.

**Detailed Summary flow diagram of Solvay Process**

Ammoniated

brine

Brine

Ammonia

regeneration

chamber

Solvay Tower/

Carbonator

Kiln/

Furnace

Ammonium

chloride

Carbon(IV)

Oxide

Haber

process

Slaker

Calcium hydroxide

Calcium

oxide

Water

Sodium

hydrogen

Carbonate

Roaster

**Sodium**

**carbonate**

Calcium chloride

Brine saturated with ammonia

Coke & Limestone

**Practice**

**1. The diagram below shows part of the Solvay process used in manufacturing sodium carbonate. Use it to answer the questions that follow.**

Carbon (IV)oxide Ammonium chloride

Saturated

Ammoniasodium chloride Process I Sodium hydrogen carbonate

solution

Sodium carbonate Process II**(a)Explain how Sodium Chloride required for this process is obtained from the sea.**

Sea water is pumped /scooped into shallow pods. Evaporation of most of the water takes place leaving a very concentrated solution. **(b)(i) Name process: I.** Filtration

**II.** Decomposition

**(ii) Write the equation for the reaction in process:**

**Process I**

Chemical equation

CO2(g) + H2O(l) + NaCl (aq) + NH3(g) -> NaHCO3(s) + NH4Cl(aq)

**Process II**

Chemical equation

2NaHCO3 (s) -> Na2CO3(s) + CO2(g) + H2O(l)

**(c)(i) Name two substances recycled in the solvay process** Ammonia gas , Carbon(IV)Oxide and Water. **(ii)Which is the by-product of this process?** Calcium(II)Chloride /CaCl2

**(iii)State two uses that the by-product can be used for:**

1. As a drying agent in the school laboratory preparation of gases.
2. In the Downs cell/process for extraction of Sodium to lower the melting point of rock salt.

**(iv)Write the chemical equation for the formation of the byproducts in the Solvay process.**

Chemical equation

Ca(OH)2 (aq) +2NH4Cl (aq) -> CaCl2(s) + 2NH3(g) + H2O(l)

**(d)In an experiment to determine the % purity of Sodium carbonate produced in the Solvay process ,2.15g of the sample reacted with exactly 40.0cm3 of 0.5M Sulphuric(VI)acid.**

**(i)Calculate the number of moles of sodium carbonate that reacted.**

Chemical equation

Na2CO3 (aq) +H2SO4 (aq) -> Na2SO4 (aq)+ CO2(g) + H2O(l)

Mole ratio Na2CO3 :H2SO4  => 1:1

Moles H2SO4 = Molarity x Volume => 0.5 x 40.0 = **0.02 Moles**

1000 1000

Moles of Na2CO3 = **0.02 Moles**

**(ii)Determine the % of sodium carbonate in the sample.**

Molar mass of Na2CO3 = **106g**

Mass of Na2CO3 = moles x Molar mass => 0.02 x 106 = **2.12 g**

**%** of Na2CO3 =(2.12 g x 100) = **98.6047%**

2.15

**(e) State two uses of soda ash.**

1. during making glass/lowering the melting point of mixture of sand/SiO2

from 1650oC and CaO from 2500oC to around 1500oC

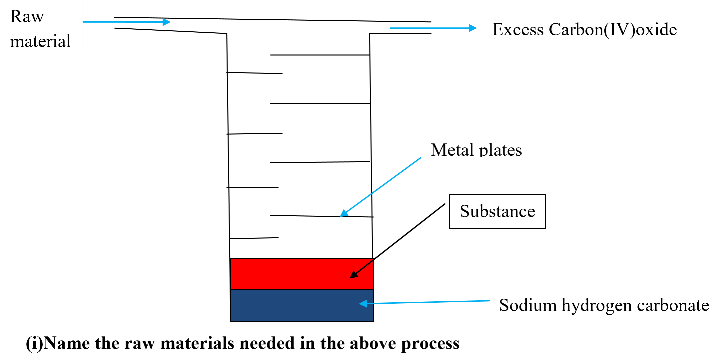
1. in softening hard water
2. in the manufacture of soapless detergents. (iv) swimming pool “pH increaser”.

**(f)The diagram below shows a simple ammonia soda tower used in manufacturing sodium carbonate .Use it to answer the questions that follow:** -Ammonia

-Water

-Carbon(IV)oxide

-Limestone

-Brine/ Concentrated sodium chloride 

**(ii)Identify substance A**

Ammonium chloride /NH4Cl

**(iii) Write the equation for the reaction taking place in: I.Tower.**

Chemical equation

CO2(g) + NaCl (aq) + H2O(l) + NH3(g) -> NaHCO3(s) + NH4Cl(aq)

1. **Production of excess carbon (IV)oxide.**

Chemical equation

CaCO3 (s) -> CaO(s) + CO2(g)

1. **The regeneration of ammonia**

Chemical equation

Ca(OH)2 (aq) +2NH4Cl (aq) -> CaCl2(s) + 2NH3(g) + H2O(l)

**(iv)Give a reason for having the circular metal plates in the tower.**

-To slow the downward flow of brine.

-To increase the rate of dissolving of ammonia.

-To increase the surface area for dissolution

**(v)Name the gases recycled in the process illustrated above.** Ammonia gas , Carbon(IV)Oxide and Water.

**2. Describe how you would differentiate between carbon (IV)oxide and carbon(II)oxide using chemical method.**

Method I

-Bubble both gases in lime water/Ca(OH)2

-white precipitate is formed if the gas is carbon (IV) oxide

* **No** white precipitate is formed if the gas is carbon (II) oxide

Method II

-ignite both gases

* Carbon (IV) oxide does **not** burn/ignite
* Carbon (II) oxide burn with a blue non-sooty flame.

Method III

-Lower a burning splint into a gas containing each gas separately. -burning splint is extinguished if the gas is carbon (IV) oxide -burning splint is **not** extinguished if the gas is carbon (II) oxide.

**3.Using Magnesium sulphate(VI)solution ,describe how you can differentiate between a solution of sodium carbonate from a solution of sodium hydrogen carbonate**

-Add Magnesium sulphate(VI) solution to separate portions of a solution of sodium carbonate and sodium hydrogen carbonate in separate test tubes

-White precipitate is formed in test tube containing sodium carbonate

-No white precipitate is formed in test tube containing sodium hydrogen carbonate.

Chemical equation

Na2CO3 (aq) +MgSO4 (aq) -> Na2SO4 (aq) + MgCO3(s)

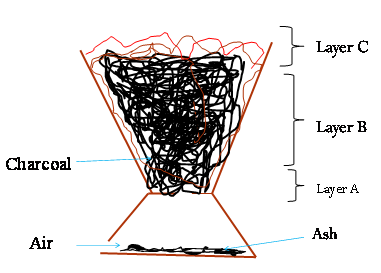
|  |  |  |
| --- | --- | --- |
| Ionic equation |  | (white ppt) |
| CO32- (aq) | + Mg2+ (aq) -> | MgCO3(s) |
|  |  | (white ppt) |

Chemical equation

2NaHCO3 (aq) +MgSO4 (aq) -> Na2SO4 (aq) + Mg(HCO3)2 (aq)

(colourless solution)

**4. The diagram below shows a common charcoal burner .Assume the burning take place in a room with sufficient supply of air.**



**(a)Explain what happens around:**

**(i)Layer A**

Sufficient/excess air /oxygen enter through the air holes into the burner .It reacts with/oxidizes Carbon to carbon(IV)oxide

Chemical equation

C(s) + O2(g) -> CO2 (g)

**(ii)Layer B**

**Hot** carbon(IV)oxide rises up and is reduced by more carbon/charcoal to carbon (II)oxide.

Chemical equation

C(s) + CO2(g) -> 2CO (g)

**(ii)Layer C**

**Hot** carbon(II)oxide rises up and burns with a blue flame to be oxidized by the excess air to form carbon(IV)oxide.

2CO (g) + O2(g) -> 2CO2(g)

**(b)State and explain what would happen if the burner is put in an enclosed room.**

The hot poisonous /toxic carbon(II)oxide rising up will not be oxidized to Carbon(IV)oxide.

**(c)Using a chemical test , describe how you would differentiate two unlabelled black solids suspected to be charcoal and copper(II)oxide.**

Method I

-Burn/Ignite the two substances separately.

-Charcoal burns with a blue flame - Copper(II)oxide does not burn

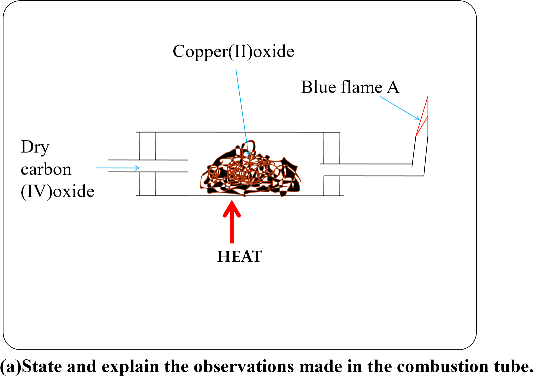
Method II

-Add dilute sulphuric(VI)acid/Nitric(V)acid/Hydrochloric acid separately.

-Charcoal does not dissolve.

- Copper(II)oxide dissolves to form a colourless solution.

**5. Excess Carbon(II)oxide was passed over heated copper(II)oxide as in the set up shown below for five minutes.**



Observation

Colour change from black to brown

Explanation

Carbon (II)oxide reduces black copper(II)oxide to brown copper metal itself oxidized to Carbon(IV)oxide.

Chemical equation

CO(g) + CuO (s) -> Cu(s) + CO2(g)

(black) (brown)

**(b) (i)Name the gas producing flame A**

Carbon(II)oxide

**(ii)Why should the gas be burnt?**

It is toxic/poisonous

**(iii)Write the chemical equation for the production of flame A**

**2**CO(g) + O2(g) -> 2CO2(g)

**(c)State and explain what happens when carbon(IV)oxide is prepared using Barium carbonate and dilute sulphuric(VI)acid.**

Reaction starts then stops after sometime producing small/little quantity of carbon(IV)oxide gas.

Barium carbonate react withdilute sulphuric(VI)acid to form insoluble Barium sulphate(VI) that cover/coat unreacted Barium carbonate stopping further reaction to produce more Carbon(IV)oxide.

1. **Using dot () and cross(x) to represent electrons show the bonding in a molecule of :**

**i) Carbon(I**

**(**

**I**

**)**

**oxide**

jgthungu@gmail.com

56

**x**

**●**

**xx**

O

**xx**

C

**x**

**●**

**●●**

lone pairs of electrons in carbon and oxygen

atoms

bonded pairs of

6

electrons

2

covalent bonds

1

dative bond

(

ii) Carbon(IV)Oxide.

jgthungu@gmail.com

30

**●**

**x**

O

**●●**

C

**●**

**x**

**●**

x

**●**

x

lone pairs of electrons in oxygen atom

4

bonded pairs of

electrons

O

**●●**

**●●**

**●●**

CO

2

**(**

**triatomic**

**molecule)**

1. **Carbon (IV)oxide is an environmental pollutant of global concern.**

**Explain.**

-It is a green house gas thus causes global warming.

-It dissolves in water to form acidic carbonic acid which causes “acid rain”

**(f)Explain using chemical equation why lime water is used to test for the presence of Carbon (IV) oxide instead of sodium hydroxide.**

Using lime water/calcium hydroxide:

* + a visible white precipitate of calcium carbonate is formed that dissolves on bubbling excess Carbon (IV) oxide gas

Chemical equation

Ca(OH)2(aq) + CO2 (g) -> CaCO3 (s) + H2O(l)

(white precipitate)

CaCO3 (aq) + H2O(l) + CO2 (g) -> Ca(HCO3) 2 (aq) Using sodium hydroxide:

* + **No** precipitate of sodium carbonate is formed Both sodium carbonate and sodium hydrogen carbonate are soluble salts/dissolves.

Chemical equation

2NaOH (aq) + CO2 (g) -> Na2CO3 (s) + H2O(l)

(No white precipitate)

Na2CO3 (s) + H2O(l) + CO2 (g) -> 2NaHCO3 (s)

**(g)Ethan-1,2-dioic acid and methanoic acid may be used to prepare small amount of carbon(II)oxide in a school laboratory.**

**(i) Explain the modification in the set up when using one over the other.**

Before carbon(II)oxide is collected:

-when using methanoic acid, **no** concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

-when using ethan-1,2-dioic acid, concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

**(ii)Write the equation for the reaction for the formation of carbon(II)oxide from: I.Methanoic acid.**

Chemical equation HCOOH(aq) -> CO(g) + H2O(l)

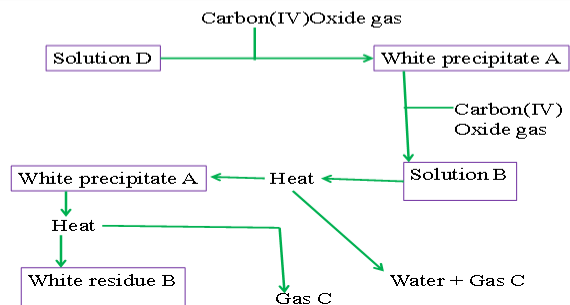
**II.** **Ethan-1,2-dioic acid**

Chemical equation HOOCCOOH(aq) -> CO2(g)+CO(g)+H2O(l) **(h)Both carbon(II)oxide and carbon(IV)oxide affect the environment. Explain why carbon(II)oxide is more toxic/poisonous.**

-Both gases are colourless,denser than water and odourless.

-Carbon(II)oxide is preferentially absorbed by human/mammalian haemoglobin when inhaled forming stable carboxyhaemoglobin instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood leading to suffocation and quick death. **-**Carbon(IV)oxide is a green house gas that increases global warming. -Carbon(II)oxide is readily oxidized to carbon(IV)oxide

**6.Study the flow chart below and use it to answer the questions that follow.**



**(a)Name:**

**(i)the white precipitate A**

Calcium carbonate

1. **solution B**

Calcium hydrogen carbonate

1. **gas C**

Carbon(IV)oxide

1. **white residue B**

Calcium oxide

1. **solution D**

**Calcium hydroxide/lime water**

**(b)Write a balanced chemical equation for the reaction for the formation of:**

1. **the white precipitate A from solution D**

Chemical equation

Ca(OH)2(aq) + CO2 (g) -> CaCO3 (s) + H2O(l)

1. **the white precipitate A from solution B**

Chemical equation

Ca(HCO3)2(aq) -> CO2 (g) + CaCO3 (s) + H2O(l)

1. **solution B from the white precipitate A**

Chemical equation

CO2 (g) + CaCO3 (s) + H2O(l) -> Ca(HCO3)2(aq)

1. **white residue B from the white precipitate A**

Chemical equation

CaCO3(s) -> CO2 (g) + CaO (s)

**(iv) reaction of white residue B with water**

Chemical equation

CaO (s) + H2O(l) -> Ca(OH)2(aq)

**CHLORINE AND ITS COMPOUND**

Chlorine is a non-metallic element in group VII (Group 17) of the periodic table. It has electronic configuration 2:8:7. It gains one valence election to form stable Cl-ion, it belongs to the chemical family of halogens.

Occurrence

-As Brine-concentration sodium chloride solution dissolved in salty seas water, oceans and lakes e.g. Lake Magadi in Kenya is very salty.

-As rock-salt solid sodium chloride crystals in the earth’s crust all over the world.

B) Preparation

Chlorine gas may be prepared in the school laboratory from the following:

a) Heating solid Manganese (iv) Oxide and Concentrated Hydrochloric acid.

b) Heating Lead (IV) Oxide and concentrated hydrochloric acid.

c) Reacting Potassium Manganate (VII) with concentrated Hydrochloric acid

d) Reacting Potassium /sodium Dichromate (VI) Acid with Concentrated Hydrochloric acid.

Set up of school laboratory preparation of chlorine.

c) Properties of chlorine. (Questions)

1. What is the colour of chlorine?

**Pale green.**

2. Describe the smell of chlorine.

**Pungent irritating smell.**

3. What method is used in collection of chlorine gas explain.

**-Downward delivery.**

**-Chlorine is 11/2 denser than air.**

4. (i) What is the purpose of concentrated sulphuric (VI) acid.

**-To dry the gas.**

(ii) Name two other substances that can be used in place of concentrated sulphuric (VI) acid.

**-Calcium chloride.**

(iii) Name a substance that cannot be used in place of concentrated sulphuric (VI) acid explain.

**-Calcium oxide reacts with chlorine.**

5.(a)Write three possible reactions between solution A and solid B:

**1. 2KMnO +16HCl → 2KCl + 2MnCl2 + 8H2O + 5Cl2**

**2. K2Cr2O7 +14HCl → 2KCl + 2Cr2Cl3 + 7H­2O + 3Cl2**

**3. Na2­Cr2O7 + 14HCl → 2NaCl + CrCl3 + 7H2O + 3Cl2**

**4. PbO2 + 4HCl → PbCl2 + Cl2 + 2H2O**

**5. MnO2 + 4HCl → MnCl2 + Cl­2 + 2H2O**

(b) Why is Hydrochloric acid used in all the above cases?

**Oxidizing agents KMnO4/PbO2/MnO2/K2Cr2O/Na2Cr­2O7 readily oxidize hydrochloric acid to chlorine themselves reduced to their chlorides.**

**Generally:**

**2HCl (aq) + [O] → CL2 (g) + H2O (l)**

**(From oxidizing agent)**

6. State and explain the observation the observation made when chlorine is bubbled in water.

**-Pale yellow colour of chlorine fades.**

**-yellow solution formed.**

**Chlorine dissolves then reacts with water to form yellow chlorine water. Chlorine water is chemically a mixture of hydrochloric acid and chloric(I)acid (hypochlorous acid)**

**Chemical equation:Cl2(g) + H2O(l) → HCl(aq) + HCl(aq)**

7. Chlorine water was exposed to sunlight for two hours as in the set up below. Next to the diagram, show and explain the observations made.

**Chlorine (I) acid is an unstable compound.**

**After two hours the chloric (I) acid in chlorine water decomposes to hydrochloric acid and releases oxygen gas. This reaction takes place in sunlight.**

8. State and explain the observation made when chlorine gas is bubbled in gas jar containing damp/wet/moist litmus papers.

**O – The blue litmus turns red then both the red/blue litmus papers are bleached/decolourized.**

**E-Chlorine reacts with water in the litmus papers to form acidic hydrochloric acid and chloric (l) acid that turns blue litmus papers red.**

**Eq.- Cl2(g) + H2O(l) → HCl(aq) + HClO(aq)**

**E.-Unstable chloric (I) acid oxidizes the dye/Coloured litmus paper to colourless material**

**NB; Chlorine does not therefore bleach/decolourize dry litmus paper/dye because chloric(I) acid cannot be formed in absence of water.**

9. Blue litmus papers were put in a flask containing cold dilute sodium hydroxide. Chlorine gas was bubbled into the solution. State and explain the observations made.

**O-blue litmus papers were bleached /decolorized.**

**Pale green colour of chlorine fades.**

**E-Sodium hydroxide reacts with chlorine to form sodium chloride and sodium hypochlorite. Sodium hypochlorite bleaches dyes by oxidation.**

**CE – Cl2 + 2NaOH → NaCl + NaClO + H2O**

**NaClO + dye →NaCl + (dye + O)**

**(coloured) (Colourless)**

**NaClO + (dye-O) → NaCl + dye**

**(Coloured) (Colourless)**

10. Blue litmus papers were put in flask containing hot concentrated sodium hydroxide. Chlorine gas was bubbled into the solution. State and explain the observations made.

**O. blue litmus papers were bleached.**

**Pale green colour of chlorine fades.**

**E-Hot concentrated sodium hydroxide reacts with chlorine to form sodium chloride and sodium chloride (V)**

**Sodium chlorate (V) bleaches by oxidation.**

**2Cl2 + 4NaOH → 3NaCl + NaClO3 + H2O**

**NaClO3 + 3(dyes) → NaCl + 3(dye + O)**

**NaClO3 is also a weed killer**

11. State three main use of chlorine gas.

-Manufacture of polyvinyl chloride (P.V.C) // polychloroethene pipes.

-Manufacture of hydrochloric acid used in “Pickling” of metals.

-Manufacture of bleaching agents

-Chlorination of water to kill germs.

12. The diagram below shows the effect of chlorine on heated iron wool.

**Method I**

**Method II**

a) Identify a suitable drying agent to dry chlorine gas.

**-Conc. H2SO4 / Concentrated sulphuric (VI) acid.**

**-Anhydrous Calcium (II) Chloride.**

**-Silica gel**

b) State and explain the observations made in Tube A in method I and II

**O - Iron glows red hot**

**Brown crystals are formed**

**E – Iron reacts with chlorine to form dark brown crystals of iron (III) Chloride.**

**This reaction is exothermic and requires no farther heating once started.**

**Iron (II) Chloride sublimes away ensuring the unreacted Iron completely reacts with chlorine gas.**

**C.Eq : 2Fe(s) + 3Cl2(g) → FeCl3(g)**

c) (i) Why is the brown solid collected at the point as shown in method I and II.

**-Heated iron (III) Chloride crystals sublime to gas and solidify on the cooler parts.**

(ii) Name another metal that can be used in place of iron to react with chlorine and collet at similar point on heating explain.

**Metal- Aluminum**

**E-Aluminum reacts with chlorine to form a while subtunate of aluminum (III) chloride at the cooler parts**

**CE: 2Al(s) + 3Cl2(g) → AlCl3(s/g)**

d) What is the purpose of suction pump?

**To pull the gaseous products into the set up.**

e) What is the function of:

(i) Sodium hydroxide in method I. Explain.

**To absorb poisonous/toxic excess unreacted chlorine gas.**

**Sodium hydroxide reacts with chlorine to form sodium chloride, Sodium hypochlorite and water.**

**CE: 2NaOH(aq) + Cl2(g) → NaCl(aq) + NaClO(aq) + H2O(l)**

(ii) Anhydrous calcium chloride in method II. Explain.

**-To absorb moisture/water in the set-up from hydrolyzing iron (III) chloride.**

**E-Iron (III) chloride easily fumes and reacts with small traces of water to form a brown solution of iron (III) hydroxide and hydrogen chloride gas.**

**CE:FeCl3(s) + 3HCl(aq) → Fe(OH)3(aq) + 3HCl(g)**

**NB: AlCl3(s) + 3HCl(aq) → Al(OH)3(aq) + 3HCl(g)**

f) Based on e (i) and (ii) above what precaution should be made in (i) method I to ensure correct results.

**-Tube B should be completely dry to prevent hydrolysis of iron (III) Chloride to iron (III) hydroxide.**

(ii) Carrying out method II

**-Should be done in a fume chamber as in the open because chlorine gas is poisonous/toxic.**

(g) Name another substance that can be used place of:

(i) Sodium hydroxide in method I

**Potassium hydroxide**

(ii) Calcium chloride in method II

(h) Calcium oxide cannot be used in place of calcium chloride during prop of chlorine. Explain.

**Calcium oxide is a base. It reacts/absorbs water to form calcium hydroxide solution.**

**Calcium hydroxide reacts with chlorine to form a mixture of calcium chloride and calcium hypochlorite.**

**2Ca (OH)2(aq) + 2Cl2(g) → CaCl2(aq) + CaOCL2(aq) + H2O(l)**

13. (a)State and explain the observation made when a piece of burning magnesium ribbon is lowered in a gas jar containing chlorine gas.

**-Magnesium ribbon contains burning with a bright flame.**

**-White solid formed.**

**-Pale yellow colour of chlorine fades**

**E: Magnesium reacts with chlorine forming a white solid of magnesium chloride.**

**CE: Mg(s) + Cl(g) → MgCl2(s)**

(b) Write the equation for the reaction that takes place if zinc is used.

**Zn(s) + Cl(g) → ZnCl2(s)**

14. The set up below shows burning phosphorus lowered in a gas jar containing chlorine gas.

a) State the observations made.

**-Phosphorus continues to burn.**

**-Dense white fumes formed.**

**-Pale green colour of chlorine fades.**

b) Write two possible equations that take place.

**P4(s) + 6Cl2(g) → 4 PCl3(s)**

**P4(s) + 10Cl2(g) → 4 PCl3(s)**

(c) State two reasons why the defragmenting spoon has a rid/cover.

**-Chlorine in the gas jar is poisonous/toxic.**

**-Burning phosphorus produces poisonous/toxic phosphorus (III) chloride // phosphorus (V) chloride.**

**-Ensure the reaction is not affected by air/oxygen from the atmosphere.**

(d) After the reaction is complete, 2cm3 of distilled water were added. The solution formed was tested with both blue and red litmus papers.

(i) State the observations made.

**-Blue litmus paper turns red**

**-Red litmus paper remain red**

(ii) Explain the observation made in d (i) above

**-Phosphoric (V) Chloride Hydrolyse in water to phosphoric (V) acid and produce hydrogen chloride gas. Both are acidic.**

**PCl5 + 4H2O → H3PO4 + 5HCl**

15. State and explain the observations made when gas jar containing chlorine is inverted over another containing hydrogen sulphide gas.

**O – Yellow solid formed.**

* **Pale colour of chlorine fades**

**E - Chlorine oxidizes hydrogen sulphide to sulphur**

**It’s self-reduced to hydrogen chloride gas. A little water catalyzes the reaction.**

**C. Eq: H2S(g) + Cl2(g) → S(s) + HCl(g)**

**(yellow solid) (White Fume)**

16. Chlorine was bubbled in aqueous ammonia solution in a beaker state and explain the observation made.

**O – White fumes evolved.**

* **Pale green colour of chlorine fades.**

**E – Chlorine reacts with ammonia gas to form a dense white fume of ammonia chloride and Nitrogen gas is produced.**

**CE – 8NH3(g) + 3Cl2(g) → 6Nh4Cl(s) + N2(g)**

17. (a) Dry gas was bubbled in cold dilute sodium hydroxide solution. Explain the observations made:

**O – Pale green colour of chlorine fades.**

**Pale yellow solution is formed.**

**E – Chlorine reacts without concentrated sodium hydroxide / Potassium hydroxide solution to form pale yellow solution of metal chlorate (V) and chlorides of the metal**

**CE: Cl2(g) + 2NaOH → NaClO(aq) + NaCl(aq) + H2O(l)**

**(Sodium hydroxide) (Sodium Chlorate (I))**

**Cl2(g) + 2KOH → KClO(aq) + NaCl(aq) + H2O(l)**

**(Potassium hydroxide) (Potassium Chlorate (I))**

(b)The experiment in 17(a) was repeated with hot concentrated sodium hydroxide solution. Explain the observation made.

**O – Pale green colour of chlorine fades.**

**- Pale yellow solution is formed.**

**E-Chlorine reacts with hot concentrated sodium hydroxide/Potassium hydroxide solution to form pale yellow solution of metal chlorate (v) and chlorides of metals.**

**CE 3Cl2(g) + 6NaOH → NaClO3 (aq) + 5NaCl(aq) + 3H2O(l)**

**(Sodium hydroxide) (Sodium Chlorate (V))**

**The products formed when chlorine reacts with alkalis depend thus on temperature and the concentration of alkalis.**

(c) (i) Write the equation for the formation of calcium chlorite (I) and calcium chlorate (V).

**2Ca (OH)2(aq) + 2Cl2(g) → CaCl2(aq) + CaOCL2(aq) + H2O(l)**

**(Calcium hydroxide) (Calcium Chlorate)**

**(Cold/dilute)**

**Ca (OH)2(aq) + Cl2(g) → CaCl2(aq) + Ca(ClO3)2(aq) + H2O(l)**

**(Calcium Chlorate)**

**B: THE HOLOGENS**

a) What are halogens?

**These are elements in group VII of the periodic table. They include**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Element** | **Symbol** | **Atomic number** | **Electric configuration** | **Charge of ion** | **Valency** | **State at Room Temperature** |
| **Fluorine**  **Chlorine**  **Bromine**  **Iodine**  **Astatine** | **F**  **Cl**  **Br**  **I**  **At** | **9**  **17**  **35**  **53**  **85** | **2:7**  **2:8:7**  **2:8:18:7**  **2:8:18:18:7**  **2:8:18:32:18:7** | **F-**  **Cl-**  **Br-**  **I-**  **At-** | **1**  **1**  **1**  **1**  **1** | **Pale yellow gas**  **Pale green gas**  **Red liquid**  **Grey Solid**  **Radioactive** |

b) Compare the atomic radius and ionic radius of chloride ion and chlorine. Explain.

**The radius of chlorine is smaller than the ionic radius of the chloride ion.**

**Effective nucleus attraction on outer energy level in chloride ion is less than chlorine atom because of extra gained electron gained electron that repelled thus causes the outer energy level to expand/increase.**

c) Compare the atomic radius of chlorine and fluorine Explain.

**Atomic radius of Fluorine is smaller than that of chlorine. Chlorine has more energy levels than fluorine occupied by more electrons.**

d) Chlorine is a gas, Bromine is a liquid, and Iodine is a solid. Explain the above observations.

-**Bromine, Chlorine and iodine exists as diatomic molecules bonded by strong covalent bond. Each molecule is joined to the other by weak intermolecular forces/ Van-der-Waal’s forces.**

**-The strength of intermolecular/Van-der-Waal’s forces of attraction increase with increase in molecular size/atomic radius Iodine has therefore the largest atomic radius and thus strongest intermolecular forces to make it a solid.**

e) (i) What is electronegativity?

**Electronegativity is the tendency/ease of acquiring /gaining electrons by an element during chemical reaction.**

**It is measured using Pauling’s scale. Where fluorine with Pauling scale 4.0 is the most electronegative element in the periodic table and thus the highest tendency to acquire/gain extra electron.**

(ii) The table below shows the electronegativity of the halogens.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Halogen** | **F** | **Cl** | **Br** | **I** | **At** |
| **Electronegativity (Pauley’s scale)** | **4.0** | **3.0** | **2.8** | **2.5** | **2.2** |

Explain the trend in electronegativity of the halogens.

**Decrease down the group from fluorine to Astatine**

**-Atomic radius increase down the group decreasing electron – attracting power down the group from fluorine to astatine.**

(f) (i)What is electron affinity

**Electron affinity is the energy required to gain an electron in an atom of an element in its gaseous state**.

(ii) Study the table below showing the election affinity of halogens for the process x+e→x-

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Halogen | F | Cl | Br | I |
| Electron affinity kJmole-1 | -333 | -364 | -342 | -295 |

(iii) Explain the trend in electron affinity of the halogens.

**-Decrease down the group**

**-Atomic radius of halogens increase down the group thus incoming/gained electron is attracted less strongly by the progressively larger atoms with a decreasing effective nuclear charge on outer energy level**

(ivWhich is a move stable ion Cl- or Br-explain?

**-Cl- ion.**

**-Has a more negative/exothermic electron affinity than Br-**

(v) Differentiate electron affinity and:

I. Ionization energy.

**Ionization energy is the energy required to loose or donate an electron in an atom of an element in its gaseous state while electron affinity is the energy required to gain/acquire extra electron by an atom of an element in its gaseous state.**

II. Electronegativity.

**-Electron affinity is the energy required to gain an electron in an atom of an element in gaseous state. It involves the process:**

**X + E → X-**

(g) (g)

**Electronegativity is the ease/tendency of gaining/ acquiring electrons by an element during chemical reactions.**

**It does not involve use of energy but theoretical arbitrary Pauling’s’ scale of measurements.**

**(g) (i)** 5cm3 of sodium chloride, Sodium bromide and Sodium iodide solutions were put separately in test tubes. 5 drops of chlorine water was added to each test tube: state and explain the observation made.

**O: - Yellow colour of chlorine water fades in all test tubes expect with sodium chloride.**

**-Coloured Solution formed.**

**E: Chlorine is more electronegative than bromine and iodine. On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.**

(ii) The experiment in g (i) was repeated with 5 drops of bromine water instead of chlorine water .explain the observation made.

**O – Yellow colour of bromine water fades in test tube containing sodium iodide.**

* **Brown solution formed in test tube containing sodium iodide**

**E - Bromine is more electronegative than iodide but less than chlorine. On adding Bromine water, iodide displaced from its solution but not chlorine.**

(iii) Using the knowledge in g (i) and (ii) above,

1. Complete the table below using (V) to show no reaction to complete the table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Halogen ion in  solution  Halogen | F- | Cl- | Br- | I- |
| F2 | X |  |  |  |
| Cl2 | X | X |  |  |
| Br2 | X | X | X |  |
| I2 | X | X | X | X |

1. Write an ionic equation for the reaction where there is (V)

(h) State one uses of

1. Fluorine – **manufacture of P.T.F.E (Poly tetra fluoroethene) synthetic fiber.**

* **Reduce tooth decay when added in small amounts/equations in tooth paste.**

**NB –large small quantities of fluorine /fluoride ions in water cause browning of teeth/fluorosis.**

* **Hydrogen fluoride is used to engrave words pictures in glass.**

1. Bromine - **Silver bromide is used to make light sensitive photographic paper/films.**
2. Iodide – Iodine dissolved in alcohol is used as medicine to kill bacteria in skin cuts. It is called tincture of iodine.
3. The table below to show some compounds of halogens.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Element  Halogen | H | Na | Mg | Al | Si | C | P |
| F | **HF** | **NaF** | **MgH2** | **AlF3** | **SiF4** | **CF4** | **PF3** |
| Cl | **HCl** | **NaCl** | **MgCl** | **AlCl3** | **SiCl3** | **CCl4** | **PCl3** |
| Br | **HBr** | **NaBr** | **MgBr2** | **AlBo3** | **SiBr4** | **CBr4** | **PBr3** |
| I | **Hl** | **Nal** | **Mgl2** | **All3** | **SiL4** | **Cl2** | **Pb3** |

(j) (i) Using dot (.) and Cross (x) to represent electrons, show the bonding in chlorine molecule.

1. Name the type of bond formed.

**Covalent.**

1. Below is the table showing the bond energy of four halogens.

Bond Bond energy k J mole-1

Cl-Cl 242

Br-Br 193

I-I 151

1. What do you understand by the term “bond energy”

**Bond energy is the energy required to break/ form one mole of chemical bond**

1. Explain the trend in bond Energy of the halogens above:

-**Decrease down the group from chlorine to Iodine**

**-Atomic radius increase down the group decreasing the energy required to break the covalent bonds between the larger atom with reduced effective nuclear @ charge an outer energy level that take part in bonding.**

(k) Some compounds of chlorine are in the table below the oxidation state of chlorine in each compound.

Compound Oxidation state Name of compound

NaClO3 **+5** **Sodium chlorate (V)**

ClO2 **+4** **Chloric (IV) oxide**

KClO2 **+3** **Potassium chlorate (III)**

NaClO **+1** **Sodium Chlorite (I)**

Cl2 **0****Chlorine Molecule**

NaCl **-1** **Sodium Chloride (I)**

MgCl2 **-1** **Magnesium Chloride (I)**

**C. HYDROGEN CHLORIDE**

1. Occurrence

Hydrogen Chloride does not occur free in the atmosphere or in nature

1. Preparation

Hydrogen chloride may be prepared in the school laboratory by reacting solid sodium/potassium chloride crystals with concentrated sulphuric (Vi) acid as in the set up below.

1. Properties of hydrogen chloride gas(questions)
2. What precautions should be taken when handling concentrated sulphuric acid? Explain.

**-Wear protective clothing/gloves to avoid accidental contact with skin.**

**-Concentrated sulphuric (VI) acid is highly corrosive-it causes painful wounds when in contact with skin.**

1. What method of gas collection is used? Explain.

**-Downward delivery// upward displacement of water**

**-Hydrogen chloride is denser than air.**

1. a) Write the equation for the reaction that takes place.

H2SO4 + NaCl NaHSO4 HCl

b) What property of concentrated sulphuric (VI) acid is used during the above reaction **-is the least volatile mineral acid, thus displace the more volatile hydrogen chloride from its salt (KCl/NaCl)**

1. i, What is the purpose of concentrated sulphuric (VI) acid in flask B.

**-Drying agent / to dry the gas.**

ii, What property of concentrated sulphuric (VI) acid is used during the above use

**-Is hygroscopic – absorbs water but do not form solution.**

iii, Using a chemical equation, explain why anhydrous calcium oxide cannot be used in flask B

**-anhydrous Calcium chloride and silica gel**

iv, Using a chemical equation, explain why anhydrous calcium oxide cannot be used in flask B

**-Calcium oxide reacts with water /moisture to form calcium hydroxide. The calcium hydroxide formed reacts with chlorine to form calcium hypochlorite.**

**This reduces the amount of @ chlorine produced.**

d)Blue and red litmus papers were dipped in the hydrogen chloride prepared above. The Procedure was repeated with damp/wet/moist litmus papers. Explain the differences in observations made.

-**Dry blue litmus papers remain blue**

**-Dry red litmus papers remain red**

**-Damp/moist/wet blue litmus papers turn red**

**-Damp/moist/wet red litmus paper remained red.**

**-Dry hydrogen chloride is a molecular compound that is joined by covalent bonds between the atoms. The gas is polar thus dissolves in water and ionize completely to free H+ that are responsible to turning blue litmus paper red.**

1. Dry hydrogen chloride gas was bubbled in two separately beakers containing water and in methylbenzene as below.
2. Classify the two solvents as either “polar” or “non-polar”

**Water – polar**

**Methylbenzene – non-polar**

(ii) State and explain the observations made in the beaker containing:

1. Methylbenzene – **Colour of litmus solution remain hydrogen chloride is a molecular substance. When dissolved in non-polar solvent, it does not dissociate / ionize H+ that changes the colour of litmus solution.**
2. Water – **colour of litmus solution change to red hydrogen chloride is a molecular substance. When dissolved in polar solvent like water, it dissociate ionize to H+ that changes litmus solution to red.**

(iii)Why should an inverted filter funnel be used?

**- The filter funnel is dipped just below the water surface to increase the surface area of dissolving the gas and prevent suck back.**

(iv)Name the solution Formed when hydrogen chloride dissolves in water.

**Hydrochloric acid**

(f) The test for hydrogen chloride gas.

**-Dip a glass rod in ammonia. Bring it to the mouth of a gas jar containing a gas suspected to be hydrogen chloride**

**-White fumes of ammonia chloride are formed.**

(g) Place 5cm3 of dilute hydrochloric acid into a four separate test tubes. To separate test tube add zinc, magnesium iron and copper metals. State and explain the observations made.

**O – Effervescence/bubbles/fizzing in all cases except copper**

* **Colourless solution formed with zinc and magnesium.**
* **Green solution formed with ion.**
* **Gas produced that extinguishes splint with explosion.**

**E. Metals above hydrogen in reactivity series react with hydrochloric and liberating hydrogen gas.**

**CE:**

**Concentrated hydrochloric acid is a weak oxidizing agent than other concentrated acids i.e Sulphuric (VI) acid and nitric (V) acid that react with all metals even those lower in the reactivity series.**

(h) Place 5cm3 of dilute hydrochloric acid into five separate test tubes. To separate test tubes, add calcium carbonate, silver carbonate, copper carbonate, iron (II) carbonate and Sodium hydrogen carbonate. Explain the observations made.

**O - Effervescence/bubbles/fizzing vigorously except in silver carbonate and lead (II) carbonate that stop later.**

* **Colourless solution formed except with iron (II) carbonate and copper (II) carbonate**
* **Green solution formed with iron (II) carbonate**
* **Blue solution formed with copper (II) carbonate**

**E. Carbonates and hydrogen carbonate react with dilute hydrochloric acid to produce carbon (IV) oxide, water and form chlorides.**

**All chlorides formed are soluble Except Lead (II) Chloride (soluble on heating/warming) and silver chloride.**

1. **Chemical equation: CaCO3 (s) + 2HCl(aq) → CaCl2(aq) + H2O(l) + CO2(g)**

**IE: (Colourless solution)**

1. **Chemical equation: Ag2CO3 (s) + 2HCl(aq) → AgCl(s) + H2O(l) + CO2(g)**

**IE: (Coats/Cover Ag2CO3)**

1. **Chemical equation: CuCO3 (s) + 2HCl(aq) → CuCl2(aq) + H2O(l) + CO2(g)**

**IE: (Blue Solution)**

1. **Chemical equation: FeCO3 (s) + 2HCl(aq) → FeCl2(aq) + H2O(l) + CO2(g)**

**IE:**

1. **Chemical equation: NaHCO3 (s) + HCl(aq) → NaCl(aq) + H2O(l) + CO2(g)**

**IE:**

1. Place 5cm3 of dilute sodium hydroxide, Potassium hydroxide and aqueous ammonia solution into three separate test tubes. Add one drop of phenolphthalein indicator drop wise, add dilute hydrochloric acid. Explain the observations made.

**O – Colour of Phenolphthalein indicator change from pink to colourless.**

**E – Hydrochloric acid neutralizes alkalis to salt and water**

**When all the alkali has reacted with the acid, An extra slight excess acid turns the indicator used to colourless.**

**Chemical equation: NaOH(aq) + HCl(aq) → NaCl(aq) + H2O(l)**

**IE:**

**Chemical equation: KOH(aq) + HCl(aq) → KCl(aq) + H2O(l)**

**IE:**

**Chemical equation: NH4OH(aq) + HCl(aq) → NHaCl(aq) + H2O(l)**

**IE:**

**(J)** Place 5cm3 of hydrochloric acid into four separate test tube tubes Separately add about 1g of each of copper (II) Oxide, Zinc (II) Oxide, Lead (II) Oxide< Calcium (II) Oxide. What happens to each test tube? Explain.

**O: All Solid dissolves except Lead (II) Oxide**

**Colourless solution formed with zinc Oxide and calcium (II) Oxide blue solution formed with copper (II) Oxide.**

**E: Metal oxides dissolves in dilute hydrochloric acid to form water and chloride salt Insoluble Lead (II) chloride and silver chloride once formed cover/coat unreacted oxides stopping further reaction.**

**Chemical equation: CuO(s) + HCl (aq) → CaCl2(aq) + H2O(l)**

**IE:**

**Chemical equation: CaO(s) + HCl (aq) → CaCl2 (aq) + H2O (l)**

**IE:**

**Chemical equation: PbO(s) + 2HCl (aq) → PbCl2 (aq) + H2O (l)**

**IE: None**

**Chemical equation: ZnO(s) + HCl (aq) → ZCl2 (aq) + H2O (l)**

**(K)** Manufacture of Hydrochloric acid.

**(i) Raw Materials**

1. Hydrogen – From electrolysis of Brine from the flowing mercury-cathode cell during the manufacture of sodium hydroxide solution.

- From water gas – passing steam in heated charcoal.

C(s) + H2O → CO(g) + H2(g)

- From Partial oxidation of natural gas methane

CH4(g) + O2(g) → CO(g) + 3H2(g)

2. Chlorine - From electrolysis of fused/solid sodium chloride in the downs process during extraction of sodium

- From electrolysis of brine/concentrated sodium chloride solution in the flowing mercury-cathode during the manufacture of sodium hydroxide solution.

1. **Chemical processes.**

* Hydrogen and chlorine gases are passed through concentrated sulphuric acid to as a drying agent.
* Small amount of pure hydrogen is continuously ignited in a chamber with continuous supply of pure dry chlorine. Large amount of hydrogen explodes.
* Hydrogen burns in chlorine to form hydrogen chloride gas.

CE: H2(g) + Cl(g) → 2HCl(g)

* The hydrogen chloride produced is then passed up to meet a downward flow of water in absorption chambers. Hydrogen chloride is very soluble in water and dissolves to form 35% concentrated hydrochloric acid.

CE: HCl(g) + (aq) → HCl(aq)

The absorption chamber is shelved and packed with broken glass beads to:

* Slow down the downward flow of water.
* Increase surface area over which the water dissolves.
* The hydrochloric acid is then transported in steel tanks lined with rubber for market

1. **Uses of Hydrochloric Acid**

* To standardize the pH of (alcohol and wines)
* Regenerating ion-exchange resin during removal of hardness of water.
* Pickling of metals to remove side layers on their surfaces.
* In the manufacture of dyes and drugs.
* Making zinc chloride for making dry cells.

1. **Diagram Showing Industrial manufacture.**

1. **Environmental effects of manufacturing HCl.**

* Hydrochloric acid is acidic. Any leakage from a manufacturing plant to nearby rivers/lake causes excess acidity that lowers pH of water killing marine life.
* Hydrogen chloride leakage into atmosphere dissolves to form “acidic rain” that accelerate corrosion in buildings, Breathing problems to human beings and kill fauna and flora around the paint.
* Chlorine leakage causes breathing and sight problems to human being. It accelerates bleaching of dyed metals.
* Hydrogen leakage can cause an explosion because impure hydrogen explodes on ignition.

1. **Factors considered in setting hydrochloric acid manufacturing plant.**
2. Nearness to the manufacturing of sodium hydroxide because the byproducts of electrolysis of brine are the raw materials for hydrochloric acid plant.
3. Availability of natural gas for extraction of hydrogen.
4. Nearness/Availability of water to dissolve the hydrogen chloride gas.
5. Availability of labour, market, capital and good means of transport.

**D CHLORIDE (Cl-) SALTS**

1. **Occurrence.**
2. Chlorides are salts derived from hydrochloric acid. Hydrochloric acid is a monobasic (HX) salt with only one ionazable/replaceable H in its molecule. All chlorides are therefore normal salts.
3. All metals exist as chloride salt except platinum and gold as below

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Metal | K | Na | Li | Mg | Ca | Al | Zn | Fe | Pb | H. | Cu | Ag | Hg |
| Formula of chloride | KCl | NaCl | LiCl | MgCl2 | CaCl2 | AlCl3 | ZnCl2 | FeCl2  FeCl3 | PbCl  PbCl4 | HCL | CuCl  CuCl2 | AgCl | Hg2Cl2  HgCl2 |

Both FeCl2 and FeCl3 exists but FeCl2 is readily oxidized to FeCl3 because it is more stable

-PbCl2 and PbCl4 exist but PbCl4 is only oxidized to form PbCl2 by using excess chlorine. It is less stable.

-CuCl and CuCl2 exists but CuCl­2 is (thermodynamically) more stable than CuCl. CuCl disproportionate to Cu and CuCl2.

* HgCl and HgCl2exists as molecular compounds.

1. All chlorides are soluble/dissolves in water except silver chloride(AgCl), Copper (I) chloride CuCl, mercury (I) Chloride Hg2Cl2 and lead (II) Chloride PbCl2 that dissolves in warm water.
2. Most chlorides are very stable compounds. They do not decompose on gentle or strong Bunsen burner heating in a school laboratory except Ammonium Chloride.
3. **Heating ammonium chloride**

Place about 2g of solid ammonium chloride crystals in a clean dry boiling tube as in the set below.

**Observation**

**-red litmus paper turn blue**

**-blue litmus paper remains blue**

Then later:

**-both blue litmus papers turn red**

**Explanation:**

Ammonium chloride on heating decomposes through **chemical sublimation** to ammonia and hydrogen chloride gas. Ammonia gas is less dense than hydrogen chloride. It is a basic gas and diffuses out faster to turn red litmus paper to blue. Hydrogen chloride is an acidic gas .It is denser than ammonia gas and thus diffuses slower than ammonia gas to turn the already both blue litmus paper to red.

Chemical equation

NH4Cl(s) -> HCl(g) + NH3 (g)

(acidic gas) (basic/alkaline gas)

1. **Test for Cl- ions**
2. The following experiment shows the test for the presence of Cl- ions in solids chloride salts.
3. **Procedure:**

Place about 1g of sodium chloride, Zinc chloride and copper (II) chloride in separate boiling tubes. Place moist blue and red litmus papers on the mouth of the test tube. Carefully, add three drops of concentrated sulphuric (VI) acid. Dip a glass rod in aqueous ammonia solution then bring it to the mouth of the boiling tube.

|  |  |
| --- | --- |
| **Observation** | **inference** |
| **-red litmus paper remain red**  **-blue litmus paper turn red**  **-vigorous effervescence/fizzing**  **/bubbling**  **-white fumes produced on** | **H+ ions**  **Cl- ions**  **HCl gas suspected** |

**(b)Explanation:**

Concentrated sulphuric (VI) acid is the less volatile mineral acid. It vigorously displaces chlorine in metal chlorides to evolve acidic hydrogen chloride gas fumes.

Chemical equation

NaCl(s) + H2SO4(l) -> NaHSO4(aq) + HCl(g)

CuCl2(s) + H2SO4(l) -> CuSO4(aq) + 2HCl(g)

ZnCl2(s) + H2SO4(l) -> ZnSO4(aq) + 2HCl(g)

Hydrogen chloride and ammonia gases react and form white fumes of ammonium chloride that confirms presence of Cl- ions in the solid substance.

Chemical equation

NH3(g) + HCl(g) -> NH4Cl(s)

1. The following experiment shows the test for the presence of Cl- ions in solution/aqueous chloride salts.

**(i)Using aqueous Lead (II) nitrate.**

**(a)Procedure:**

I.Place about 5cm3 of sodium chloride, Iron (III) chloride and copper (II) chloride in separate boiling tubes. Add four drops of Lead (II) nitrate solution to each. Preserve.

|  |  |
| --- | --- |
| **Observation** | **Inference** |
| **White precipitate/ppt** | **SO42-, SO32-, Cl-,CO32-** |

II.To the preserved sample, add six drops of nitric (V) acid. Preserve.

|  |  |
| --- | --- |
| **Observation** | **Inference** |
| **White precipitate/ppt persist** | **SO42-, Cl-** |

III. To the preserved sample, heat the mixture to boil

|  |  |
| --- | --- |
| **Observation** | **Inference** |
| **White precipitate/ppt dissolves on boiling/warming** | **Cl-** |

**Explanation:**

I. When Lead(II) nitrate(V) solution is added to an unknown salt , a white precipitate/ppt of Lead(II) sulphate(VI) Lead(II) carbonate(IV) Lead(II) sulphate(IV) Lead(II) chloride(I) are formed.

Ionic equation:

Pb2+ (aq) + SO42-(aq) -> PbSO4(s)

Pb2+ (aq) + SO32-(aq) -> PbSO3(s)

Pb2+ (aq) + CO32-(aq) -> PbCO3(s)

Pb2+ (aq) + Cl-(aq) -> PbCl2(s)

II. When the white precipitate/ppt formed is acidified with dilute nitric(V) acid, the white precipitate of Lead(II) sulphate and Lead(II) chloride(I) persist/remain while that of Lead(II) carbonate and Lead(II) sulphate(IV) dissolves.

III.On heating /warming Lead (II) chloride dissolves but on cooling it recrystallizes. This shows the presence of Cl-ions in aqueous solutions

.

**(ii)Using aqueous silver nitrate.**

**Procedure**

I. Place about 5cm3 of sodium chloride, Iron (III) chloride and copper (II) chloride in separate boiling tubes. Add four drops of silver nitrate solution to each. Preserve.

|  |  |
| --- | --- |
| **Observation** | **Inference** |
| **White precipitate/ppt** | **Cl-,CO32-** |

II. To the preserved sample, add six drops of nitric (V) acid. Preserve.

|  |  |
| --- | --- |
| **Observation** | **Inference** |
| **White precipitate/ppt persist** | **Cl-** |

**Explanation:**

I. When silver nitrate solution is added to an unknown salt , a white precipitate/ppt of silver(I) carbonate(IV) and silver(I) chloride(I) are formed.

Ionic equation:

2Ag+ (aq) + CO32-(aq) -> Ag2CO3(s)

Ag+ (aq) + Cl-(aq) ->AgCl(s)

II.When the white precipitate/ppt formed is acidified with dilute nitric (V) acid, the white precipitate of silver (I) chloride (I) persist/remain. This shows the presence of Cl-ions in aqueous solutions.

Silver (I) carbonate (IV) dissolves when reacted with nitric (V) acid.

**COMPREHENSIVE REVISION QUESTIONS**

1In an experiment, dry hydrogen chloride gas was passed through heated zinc turnings as in the set up below. The gas produced was the passed through copper(II) oxide

Tube S Zinc turnings Tube V copper(II) oxide

Dry gas K

HCl

HEAT HEAT

Flask T Water

1. Write the equation for the reaction :

(i)For the school laboratory preparation of hydrogen chloride gas.

**NaCl(s) + H2SO4(l) -> NaHSO4(aq) + HCl(g)**

(ii)in tube S

**Zn(s) + 2HCl(aq) -> ZnCl2(aq) + H2(g)**

b)State and explain the observation made in tube V.

Observations-**colour of solid changes from black to brown**

**-colourless liquid forms on the cooler parts of tube V**

Explanation-**Hydrogen produced in tube S reduces black copper(II) oxide to brown copper metal and the gas oxidized to water vapour that condense on cooler parts..**

Chemical equation.

**CuO(s) +H2(g) ->Cu(s) + H2O(l)**

(c)How would the total mass of tube S and tube V and their contents compare before and after the experiment.

Tube S- **Mass increase/rise because Zinc combine with chlorine to form heavier Zinc Chloride.**

Tube V- **Mass decrease/falls/lowers because copper (II) oxide is reduced to lighter copper and oxygen combine with hydrogen to form water vapour that escape.**

2. Chlorine is prepared by using solid sodium chloride, concentrated sulphuric(VI) acid and potassium manganate(VII)

a)What is the role of the following in the reaction;

(i) concentrated sulphuric(VI) acid

**To produce hydrogen chloride gas by reacting with the solid sodium chloride.**

(ii) Potassium manganate(VII)

**To oxidize hydrogen chloride gas to chlorine**

3. Use the flow chart below to answer the questions that follow.

Water KMnO4 Hot concentrated

sodium hydroxide

KCl(s) + gas X solution W gas Q bleaching agent Z

H2SO4

a)(i) Name:

gas X **Hydrogen chloride**

solution W **hydrochloric acid**

gas Q **chlorine**

bleaching agent Z **sodium chlorate(V)**

b) Write the chemical equation for the formation of:

(i) gas X

**NaCl(s) + H2SO4(l) -> NaHSO4(aq) + HCl(g)**

(ii)solution W

**HCl(g) + (aq) -> HCl(aq)**

(iii)gas Q

**2KMnO4 + 16HCl(aq) -> 2KCl(aq) + 2MnCl2(aq) + 8H2O(l) + 5Cl2(g)**

(iv)bleaching agent Z

**6NaOH(aq) +** **3Cl2(g) ->NaCl(aq) + NaClO3(aq) + 3H2O(l)**

c) State and explain the following observations;

(i) a glass rod dipped in aqueous ammonia is brought near gas X

**Observation: Dense white fumes**

**Explanation: Ammonia gas reacts with hydrogen chloride gas to form dense white fumes of ammonium chloride.**

**Chemical equation: NH3(g) +HCl(g) -> NH4Cl(s)**

(ii)Wet blue and red litmus papers were dipped into gas Q

**Observations: Blue litmus paper turned red the both are bleached**

**/decolorized.**

**Explanations: chlorine reacts with water to form both acidic hydrochloric and chloric (I) acids that turn blue litmus paper red. Unstable chloric (I) acid oxidizes the dye in the papers to colourless.**

**Chemical equations**

**Cl2(g) + HCl(aq) ->HCl(aq) + HClO(aq)**

**Coloured dye + HClO(aq) ->HCl(aq) + (Colourless dye +O)//**

**(Coloured dye-O) + HClO(aq) ->HCl(aq) + Colourless dye**

4. Use the flow chart below to answer the questions that follow

Liquid A water

Rock salt hydrogen B process NaCl(aq) +

Chloride Z H2O

NH3(g) AgNO3(aq)

White solid X White precipitate C

1. Name

Liquid A **Concentrated sulphuric(VI) acid**

Process Z **Neutralization**

White solid X **Ammonium chloride**

b) Write the equation for the formation of:

(i) Hydrogen chloride

**NaCl(s) + H2SO4(l) -> NaHSO4(aq) + HCl(g)**

(ii) B

**HCl(g) + (aq) -> HCl(aq)**

(iii)process Z (using ionic equation)

**H+ (aq) + OH-(aq) -> H2O(l)**

(iv)C (using ionic equation)

**Ag+ (aq) + Cl-(aq) -> AgCl(s)**

c)Describe how solution B is obtained.

**Bubbling hydrogen chloride gas through inverted funnel into distilled water until no more dissolve.**

5 The results obtained when halogens are bubbled into test tubes containing solutions of halide A,B and C is as in the table below. Tick(v) means a reaction took place. Cross(x) means no reaction took place.

|  |  |  |  |
| --- | --- | --- | --- |
| Halogens | Halide ions in solution | | |
| A | B | C |
| I2 | x | - | x |
| Br2 | x | v | - |
| Cl2 | - | v | v |

a)Identify the halide ions represented by letter

A **Cl-**

B **I-**

C **Br-**

b)Write the ionic equation for the reaction that take place with halide:

(i) C

**Cl2(g) + 2Br-(aq) -> 2Cl-(aq) + Br2(aq)**

(ii) B

**Cl2(g) + 2Br-(aq) -> 2Cl-(aq) + Br2(aq)**

**Cl2(g) + 2I-(aq) -> 2Cl-(aq) + I2(aq)**

6. The diagram below shows a set-up of apparatus for the school laboratory collection of dry chlorine gas.

Substance Q

Dry chlorine

MnO2 Drying agent L

a)Name:

(i) substance Q **Concentrated hydrochloric acid**

(ii)suitable drying agent L

**-Concentrated sulphuric(VI) acid**

**-anhydrous calcium chloride**

**-silica gel**

b) State a missing condition for the reaction to take place faster.

**-Heat/Heating**

c)Red and blue litmus papers were dipped into the chlorine gas from the above set up .State and explain the observations made.

**Observation: Blue litmus paper remain blue. Red litmus paper remain red.**

**Explanation: Dry chlorine has no effect on dry litmus papers.**

d)Write the equation for the reaction taking place in the conical flask

**MnO4 (s) + 4HCl(aq) -> MnCl2(aq) + 2H2O(l) + Cl2(g)**

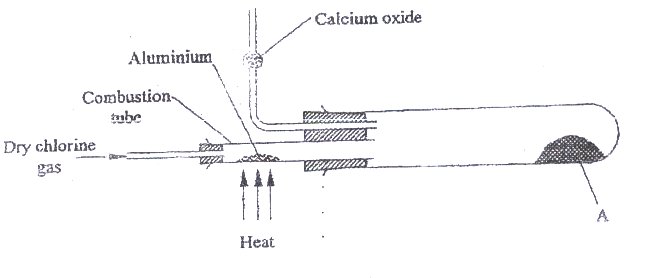
e)Name two other substances that can be used in place of MnO2

**Lead(IV) oxide (PbO2)**

**Potassium manganate(VI)(KMnO4)**

**Potassium dichromate(K2Cr2O4)**

**Bleaching powder(CaOCl2)**

7.The set up below shows the apparatus used to prepare and collect anhydrous iron(II) chloride.

Iron

liquid P cold water

HEAT tube Z solid S

solid Q Liquid R

a)Name suitable substance

**P concentrated sulphuric(VI) acid.**

**Q sodium /potassium chloride**

**R concentrated sulphuric(VI) acid.**

1. Write the equation for the reaction:

I.between P and Q

**NaCl(s) + H2SO4(l) -> NaHSO4(aq) + HCl(g)**

**KCl(s) + H2SO4(l) -> KHSO4(aq) + HCl(g)**

II.For the formation of S

**Fe(s) + 2HCl(g) -> FeCl2(s) + H2(g)**

1. (i) Name two gases that come out of tube Z

**-excess hydrogen chloride**

**-hydrogen**

(ii) The two gases were passed through sodium hydroxide solution.

I.Write an ionic equation for the reaction that take place.

**HCl(g) + OH-(aq) ->Cl-(aq) + H2O(l)**

II a burning splint was introduced to the gas that comes out of the sodium hydroxide solution. State the observation made.

**The splint is extinguished with an explosion/pop sound**

1. State and explain the following

(i)Small amount of water is added to iron (II) chloride in a test tube then shaken

**Solid dissolves to form a green solution. Iron(II) chloride is soluble in water**

(ii)I. Three drops of aqueous sodium hydroxide is added to aqueous iron(II) chloride and then added excess of the alkali.

Observation: **Green precipitate is formed that persist/remain /insoluble in excess alkali.**

Explanation**: Iron(II) chloride reacts with aqueous sodium hydroxide to form a green precipitate of iron(II) hydroxide.**

Ionic equation:**Fe2+(aq) + OH-(aq) -> Fe(OH)2(s)**

II. Six drops of hydrogen peroxide is added to the mixture in d(ii) above.

Observation: **Effervescence/bubbling/fizzing take place and the green precipitate dissolve to form a yellow/brown solution.**

Explanation**: hydrogen peroxide oxidizes green Fe2+to yellow/ brown Fe3+solution.**

8. Use the flow chart below to answer the questions that follow:

Iron water

Hydrogen chloride solid Y aqueous Y

Water aqueous ammonia

Hydrochloric acid

Potassium manganate (VII) Green precipitate

Gas Z Bleaching agent A

Dilute sodium hydroxide

a)Write an equation for the school laboratory formation of hydrogen chloride gas

**NaCl(s) + H2SO4(l) -> NaHSO4(aq) + HCl(g)**

**KCl(s) + H2SO4(l) -> KHSO4(aq) + HCl(g)**

b)Name:

I. solid Y **Iron (II) chloride (FeCl2)**

II green precipitate **Iron (II) hydroxide (Fe (OH)2**

III Gas Y **Chlorine (Cl2)**

IV. Bleaching agent A **Sodium hypochlorite (NaOCl)**

c)Blue and red litmus papers were dipped into bleaching agent A. Write the equation for the reaction that takes place.

**Coloured dye + NaOCl(aq) ->NaCl(aq) + (Colourless dye + O)//**

**(Coloured dye-O) + NaOCl(aq) ->NaCl(aq) + Colourless dye**

d)State four uses of gas Z

1. **Bleaching agent**
2. **Manufacture of hydrochloric acid**
3. **Chlorination of water to kill germs**
4. **Manufacture of PVC pipes**
5. Calculate the mass of solid Y formed when 60cm3 of hydrogen chloride at r.t.p is completely reacted. (1 mole of a gas =24dm3 at r.t.p, Fe = 56.O, Cl= 35.5)

**Chemical equation; Fe(s) + 2HCl(g) -> FeCl2(s) + Cl2(g)**

**Mole ratio HCl: FeCl2 = 1:1 molar mass Cl2 = 127g**

**Moles of HCl used = 60cm3 /24000cm3 = 2.5 x 10 -3 moles**

**Moles of FeCl2 = Moles of HCl => 2.5 x 10 -3 moles**

**Mass of FeCl2 = moles x molar mass => 2.5 x 10 -3 x 127 =0.3175g**

9.Use the flow chart below to answer the questions that follow.

Pb(NO3)2

Rock salt conc.H2SO4  Gas A white precipitate B

HNO3

White precipitate D Heat white precipitate C

No white precipitate No white precipitate

a)Write the chemical equation for the formation of gas A

**NaCl(s) + H2SO4(l) -> NaHSO4(aq) + HCl(g)**

b)Identify:

(i) four possible ions that can produce white precipitate B

**SO42-,SO32-, CO32-, Cl-**

(ii)two possible ions that can produce;

I.White precipitate C

**SO42-,Cl-**

II.colourless solution D

**SO32-, CO32-**

(iii)possible ions present in

I.White precipitate E

**SO42-**

II.colourless solution F

**Cl-**

10. Study the flow chart below and use it to answer the questions that follow.

Iron water NaOH(aq)

HCl(g) heat solid P solution of P green ppt

HCl(aq) Gas Q Solid R Solution of R brown ppt

KMnO4Iron Water NaOH(aq)

a)Identify substance:

P **Iron(II) chloride//FeCl2**

Q **Chlorine // Cl2**

R **Iron(III) chloride//FeCl3**

b)Write the equation for the reaction for the formation of:

(i) gas Q

**2KMnO4 (s) + 16HCl(aq) -> 2KCl(aq) + 2MnCl2(aq) + 8H2O(l) + 5Cl2(g)**

(ii) the green precipitate (using ionic equation)

Ionic equation:**Fe2+(aq) + 2OH-(aq) -> Fe(OH)2(s)**

(ii) the brown precipitate (using ionic equation)

Ionic equation:**Fe3+(aq) + 3OH-(aq) -> Fe(OH)3(s)**

c)A glass rod was dipped in aqueous ammonia. The rod was then brought near hydrogen chloride. State and explain the observation made.

**Observation: White fumes**

**Explanation: Ammonia gas reacts with hydrogen chloride gas to form white fumes of ammonium chloride.**

**Chemical equation: NH3(g) + HCl(g) -> NH4Cl(s)**

11. Below is a set up in the preparation of a particular salt. Study it and answer the questions that follow.

Iron wool. Anhydrous calcium chloride

Dry chlorine

HEAT salt K

1. State the observation made when iron wool is heated.

**Iron wool glows red hot. Colour changes from grey to dark grey/black.**

b)(i) Identify salt K

**iron(III) chloride// FeCl3**

(ii)Write the equation for the formation of salt K

2Fe**(s) + 3Cl2(g) -> 2FeCl3(s/g)**

(iii)What property of salt K is exhibited as shown in the experiment.

**It sublimes//sublimation.**

(iv)Calculate the minimum volume of chlorine required to form 700kg of iron(III) chloride at room temperature.(Fe= 56.0, Cl=35.5, 1 mole of a gas =24000cm3, 1000g = 1kg)

**Mole ratio Fe : Cl2 = 2: 3 molar mass** Fe**Cl3 = 162.5g**

**Method 1**

**2 x 162.5 g FeCl3 -> 3x 22400 cm3 Cl2**

**700 x1000 gFeCl3 -> (700 x1000 x3 x22400)/(2 x 162.5)**

**=1.4474 x 10-8 cm3**

**Method 2**

**Moles of FeCl3= mass/ molar mass**

**=> (700 x 1000) / 162.5 = 4307.6923 moles**

**Moles of Cl2= 3/2 moles of FeCl3**

**=>3/2 x 4307.6923 = 6461.5385 moles**

**Volume of chlorine= moles x molar gas volume**

**=>6461.5385 x 24000 = 1.5508 x 10-8 cm3**

c) Name another metal that can produce similar results as salt K.

**Aluminium// Al**

d)(i) What is the purpose of anhydrous calcium chloride.

**-ensure the apparatus are water free.**

**-prevent water from the atmosphere from entering and altering//hydrolysing salt K**

(ii) Write the equation for the reaction that take place if anhydrous calcium chloride is not used in the above set up.

**FeCl3(s) + 3H2O(l) -> Fe(OH)3(aq)** + 3**HCl(g)**

**SULPHUR (S)**

Sulphur is an element in Group VI(Group 16)of the Periodic table . It has atomic number 16 and electronic configuration 16 and valency 2 /divalent and thus forms the ion S2-

A. Occurrence.

Sulphur mainly occurs :

(i) as **free** element in Texas and Louisiana in USA and Sicily in Italy.

(ii)**Hydrogen sulphide** gas in active volcanic areas e.g. Olkaria near Naivasha in Kenya

(iii)as **copper pyrites(**CuFeS2**)** and **iron pyrites**(FeS2) in other parts of the world.

B. Extraction of Sulphur from **Fraschs** process

Suphur occurs about 200 metres underground. The soil structure in these areas is usually **weak** and can easily **cave** in.

Digging of tunnels is thus discouraged in trying to extract the mineral.

Sulphur is extracted by drilling three concentric /round pipes of diameter of ratios 2:8: 18 centimeters.

Superheated water at 170oC and 10atmosphere pressure is forced through the outermost pipe.

The high pressures ensure the water remains as liquid at high temperatures instead of vapour of vapour /gas.

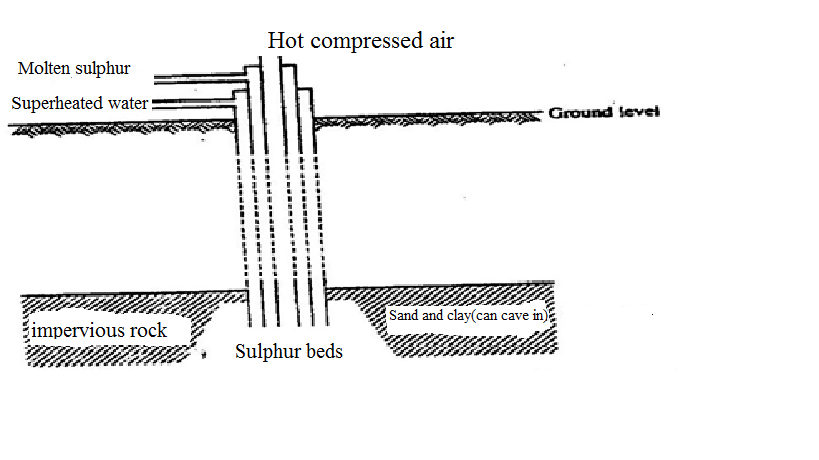
The superheated water melts the sulphur because the melting point of sulphur is lower at about at about 115oC.

A compressed air at 15 atmospheres is forced /pumped through the innermost pipe.

The hot air forces the molten sulphur up the middle pipe where it is collected and solidifies in a large tank.

It is about 99% pure.

Diagram showing extraction of Sulphur from Fraschs Process



C. Allotropes of Sulphur.

1. Sulphur exist as two crystalline allotropic forms:

(i)Rhombic sulphur

(ii)Monoclinic sulphur

|  |  |
| --- | --- |
| Rhombic sulphur | Monoclinic sulphur |
| Bright yellow crystalline solid  Has a melting point of 113oC  Has a density of 2.06gcm-3  Stable below 96oC  Has octahedral structure | Pale yellow crystalline solid  Has a melting point of 119oC  Has a density of 1.96gcm-3  Stable above 96oC  Has a needle-like structure |

Rhombic sulphur and Monoclinic sulphur have a **transition** temperature of 96oC.This is the temperature at which one allotrope changes to the other.

2. Sulphur exists in non-crystalline forms as:

(i)Plastic sulphur-

Plastic sulphur is prepared from heating powdered sulphur to boil then pouring a thin continuous stream in a beaker with cold water. A long thin elastic yellow thread of plastic sulphur is formed .If left for long it turn to bright yellow crystalline rhombic sulphur.

(ii)Colloidal sulphur-

Colloidal sulphur is formed when sodium thiosulphate (Na2S2O3) is added hydrochloric acid to form a yellow precipitate.

D. Heating Sulphur.

A molecule of sulphur exists as puckered ring of eight atoms joined by covalent bonds as S8.

● ● ●

Sulphur atoms

Covalent bonds

● ●

● ●

●

On heating the yellow sulphur powder melts at 113oC to clear amber liquid with low viscosity and thus flows easily.

On further heating to 160oC the molten liquid darkens to a brown very viscous liquid that does not flow easily.

This is because the S8 rings break into S8 chain that join together to form very long chains made of over 100000 atoms of Sulphur.

The long chains **entangle** each other reducing their mobility /flow and hence increases their viscosity.

On continued further heating to above 160oC, the viscous liquid darkens but becomes more mobile/flows easily and thus less viscous.

This is because the long chains break to smaller/shorter chains.

At 444oC, the liquid boils and forms brown vapour of a mixture of S8 ,S6 ,S2 molecules that solidifies to S8 ring of “flowers of sulphur” on the cooler parts.

Summary of changes on heating sulphur

|  |  |
| --- | --- |
| Observation on heating | Explanation/structure of Sulphur |
| Solid sulphur  Heat to 113oC  Amber yellow liquid  Heat to 160oC  Liquid darkens  Heat to 444oC  Liquid boils to brown vapour  Cool to room temperature  Yellow sublimate  (Flowers of Sulphur) | Puckered S8 ring  Puckered S8 ring in liquid form (low viscosity/flow easily)  Puckered S8 ring break/opens then join to form long chains that entangle (very high viscosity/very low rate of flow)  Mixture of S8 ,S6 ,S2 vapour  Puckered S8 ring |

E. Physical and Chemical properties of Sulphur.(**Questions**)

**1. State three physical properties unique to Sulphur**

Sulphur is a yellow solid, insoluble in water, soluble in carbon disulphide/tetrachloromethane/benzene, poor conductor of heat and electricity. It has a melting point of 115oC and a boiling point of 444oC.

**2. Moist/damp/wet blue and red litmus papers were put in a gas jar containing air/oxygen. Burning sulphur was then lowered into the gas jar. State and explain the observation made.**

Observations

-Sulphur melts then burns with a blue flame

Colourless gas produced that has a pungent smell

Red litmus paper remains red. Blue litmus paper turns red.

Explanation

Sulphur burns in air and faster in Oxygen to form Sulphur(IV)Oxide gas and traces/small amount of Sulphur(VI)Oxide gas. Both oxides react with water to form the corresponding acidic solution i.e

(i) Sulphur(IV)Oxide gas reacts with water to form sulphuric(IV)acid

(ii) Sulphur(VI)Oxide gas reacts with water to form sulphuric(VI)acid

Chemical equation

S(s) + O2(g) -> SO2(g) (Sulphur(IV)Oxide gas)

2S(s) + 3O2(g) -> 2SO3(g) (Traces of Sulphur(VI)Oxide gas)

SO2(g) + H2O(l) -> H2 SO3 (aq) ( sulphuric(IV)acid) SO3(g) + H2O(l) -> H2 SO4 (aq) ( sulphuric(VI)acid).

**3. Iron filings were put in a test tube containing powdered sulphur then heated on a Bunsen flame. Stop heating when reaction starts. State and explain the observations made. Test the effects of a magnet on the mixture before and after heating. Explain.**

Observations

Before heating, the magnet attracts iron filings leaving sulphur

After heating, the magnet does not attract the mixture.

After heating, a red glow is observed that continues even when heating is stopped..

Black solid is formed.

Explanation Iron is attracted to a magnet because it is ferromagnetic. When a mixture of iron and sulphur is heated, the reaction is exothermic giving out heat energy that makes the mixture to continue glowing even after stopping heating. Black Iron(II)sulphide is formed which is a compound and thus not ferromagnetic.

Chemical equation Fe(s) + S(s) -> FeS(s) (Exothermic reaction/ **-**∆H)

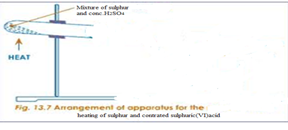
Heated powdered heavy metals combine with sulphur to form black sulphides.

Cu(s) + S(s) -> CuS(s)

Zn(s) + S(s) -> ZnS(s)

Pb(s) + S(s) -> PbS(s)

**4.The set up below show the reaction of sulphur on heated concentrated sulphuric(VI)acid.**



**(i)State and explain the observation made.**

Observation

Yellow colour of sulphur fades

Orange colour of potassium dichromate(VI)paper turns to green.

Explanation

Hot concentrated sulphuric(VI)acid oxidizes sulphur to sulphur (IV)oxide gas. The oxide is also reduced to water. Traces of sulphur (VI)oxide is formed.

Chemical equation

S(s) + 3H2 SO4 (l) -> 3SO2(g) + 3H2O(l) +SO3(g)

Sulphur (IV)oxide gas turns Orange potassium dichromate(VI)paper to green.

**(ii)State and explain the observation made if concentrated sulphuric (VI) acid is replaced with concentrated Nitric (V) acid in the above set up.**

Observation

Yellow colour of sulphur fades

Colurless solution formed

Brown fumes/gas produced.

Explanation

Hot concentrated Nitric(V)acid oxidizes sulphur to sulphuric (VI)acid. The Nitric (V) acid is reduced to brown nitrogen(IV)oxide gas.

Chemical equation

S(s) + 6HNO3 (l) -> 6NO2(g) + 2H2O(l) +H2SO4 (l)

NB:

Hydrochloric acid is a weaker oxidizing agent and thus cannot oxidize sulphur like the other mineral acids.

**5.State three main uses of sulphur .**

Sulphur is mainly used in:

(i)Contact process for the manufacture/industrial/large scale production of concentrated sulphuric(VI)acid.

(ii)Vulcanization of rubber to make it harder, tougher, stronger, and more durable.

(iii)Making gun powder and match stick heads

(iv) As ointments to treat fungal infections

**6. Revision Practice**

**The diagram below represents the extraction of sulphur by Fraschs process. Use it to answer the questions that follow.**

N

M

L

**(a)Name the substances that passes through:**

**M** Superheated water at 170oC and 10 atmosphere pressure

**L** Hot compressed air

N Molten sulphur

**(b)What is the purpose of the substance that passes through L and M?**

M- Superheated water at 170oC and 10 atmosphere pressure is used to melt the sulphur

L- Hot compressed air is used to force up the molten sulphur.

**(c) The properties of the two main allotropes of sulphur represented by letters A and B are given in the table below. Use it to answer the questions that follow.**

|  |  |  |
| --- | --- | --- |
|  | **A** | **B** |
| **Appearance** | **Bright yellow** | **Pale yellow** |
| **Density(gcm-3)** | **1.93** | **2.08** |
| **Melting point(oC)** | **119** | **113** |
| **Stability** | **Above 96oC** | **Below 96oC** |

**I.What are allotropes?**

Different forms of the same element existing at the same temperature and pressure without change of state.

**II. Identify allotrope:**

1. Monoclinic sulphur

B . Rhombic sulphur

**III. State two main uses of sulphur.**

-Manufacture of sulphuric(VI)acid

-as fungicide

-in vulcanization of rubber to make it harder/tougher/ stronger

-manufacture of dyes /fibres

**(d)Calculate the volume of sulphur (IV)oxide produced when 0.4 g of sulphur is completely burnt in excess air.(S = 32.0 ,I mole of a gas occupies 24 dm3 at room temperature)**

Chemical equation

S(s) + O2(g) -> SO2(g)

Mole ratio S: SO2 = 1:1

**Method 1**

32.0 g of sulphur -> 24 dm3 of SO2(g)

0.4 g of sulphur -> 0.4 g x 24 dm3 = **0.3 dm3**

32.0 g

**Method 2**

Moles of sulphur used= Mass of sulphur => 0.4 = 0.0125 moles

Molar mass of sulphur 32

Moles of sulphur used = Moles of sulphur(IV)oxide used=>0.0125 moles

Volume of sulphur(IV)oxide used = Moles of sulphur(IV)oxide x volume of one mole of gas =>0.0125 moles x 24 dm3 = **0.3 dm3**

**B. COMPOUNDS OF SULPHUR**

The following are the main compounds of sulphur:

(i) Sulphur(IV)oxide

(ii) Sulphur(VI)oxide **.**

(iii) Sulphuric(VI)acid

(iv) Hydrogen Sulphide

(v) Sulphate(IV)/SO32- and Sulphate(VI)/ SO42- salts

**(i) Sulphur(IV)oxide(SO2)**

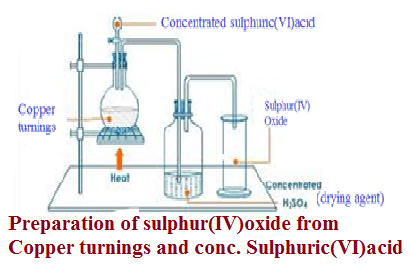
**(a) Occurrence**

Sulphur (IV)oxide is found in volcanic areas as a gas or dissolved in water from geysersand hot springs in active volcanic areas of the world e.g. Olkaria and Hells gate near Naivasha in Kenya.

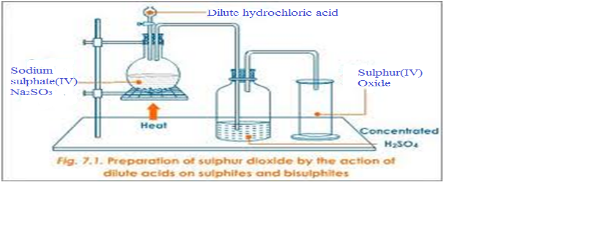
**(b) School laboratory preparation**

In a Chemistry school laboratory Sulphur (IV)oxide is prepared from the reaction of

Method 1:Copper and Sulphuric(VI)acid.



Method 2:Sodium Sulphate(IV) and hydrochloric acid.



**(c)Properties of Sulphur(IV)oxide(Questions)**

**1. Write the equations for the reaction for the formation of sulphur (IV)oxide using:**

**(i)Method 1**

Cu(s) + 2H2SO4(l) -> CuSO4(aq) + SO2(g) + 2H2O(l)

Zn(s) + 2H2SO4(l) -> ZnSO4(aq) + SO2(g) + 2H2O(l)

Mg(s) + 2H2SO4(l) -> MgSO4(aq) + SO2(g) + 2H2O(l)

Fe(s) + 2H2SO4(l) -> FeSO4(aq) + SO2(g) + 2H2O(l)

Calcium ,Lead and Barium will form insoluble sulphate(VI)salts that will cover unreacted metals stopping further reaction thus producing very small amount/quantity of sulphur (IV)oxide gas.

**(ii)Method 2**

Na2SO3(aq) + HCl(aq) -> NaCl(aq ) + SO2(g) + 2H2O(l)

K2SO3(aq) + HCl(aq) -> KCl(aq ) + SO2(g) + 2H2O(l)

BaSO3(s) + 2HCl(aq) -> BaCl2(aq ) + SO2(g) + H2O(l)

CaSO3(s) + 2HCl(aq) -> CaCl2(aq ) + SO2(g) + H2O(l)

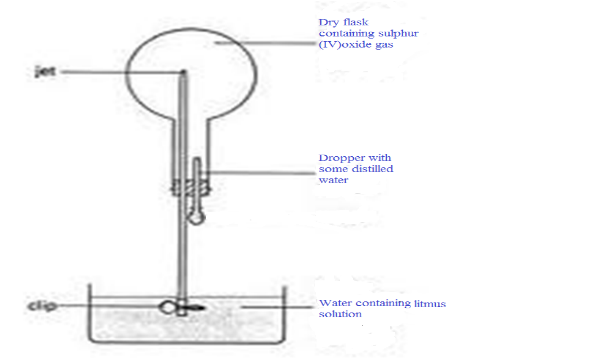
PbSO3(s) + 2HCl(aq) -> PbCl2(s ) + SO2(g) + H2O(l)

Lead(II)chloride is soluble on heating thus reactants should be heated to prevent it coating/covering unreacted PbSO3(s)

**2.State the physical properties unique to sulphur (IV)oxide gas.**

Sulphur (IV)oxide gas is a colourless gas with a pungent irritating smell and choking smell which liquidifies easily. It is about two times denser than air.

**3. The diagram below show the solubility of** sulphur (IV)oxide gas. Explain.



Sulphur(IV) oxide is very soluble in water.

One drop of water dissolves all the Sulphur (IV) oxide in the flask leaving a vacuum.

If the clip is removed, atmospheric pressure forces the water up through the narrow tube to form a fountain to occupy the vacuum.

An acidic solution of sulphuric (IV)acid is formed which turns litmus solution red.

Chemical equation

SO2(g) + H2O(l) -> H2 SO3 (aq) ( sulphuric(IV)acid turn litmus red)

**4.Dry litmus papers and wet/damp/moist litmus papers were put in a gas jar containing sulphur(IV) oxide gas. State and explain the observations made.**

Observations

(i)Dry Blue litmus paper remains blue.

Dry red litmus paper remains red.

(ii) Wet/damp/moist blue litmus paper turns red.

Moist/damp/wet red litmus paper remains red.

Both litmus papers are then bleached /decolorized.

Explanation

Dry sulphur(IV) oxide gas is a molecular compound that does not dissociate/ionize to release H+(aq)ions and thus has no effect on dry blue/red litmus papers.

Wet/damp/moist litmus papers contain water that dissolves /react with dry sulphur(IV) oxide gas to form a solution of weak sulphuric(IV)acid (H2 SO3 (aq)).

Weak sulphuric(IV)acid(H2 SO3 (aq)) dissociates /ionizes into free H+(aq)ions:

H2 SO3 (aq) -> 2H+(aq) + SO32- (aq)

The free H+(aq)ions are responsible for turning blue litmus paper turns red showing the gas is acidic.

The SO32- (aq) ions in wet/damp/moist sulphur(IV) oxide gas is responsible for many reactions of the gas.

It is easily/readily oxidized to sulphate(VI) SO42- (aq) ions making sulphur(IV) oxide gas act as a reducing agent as in the following examples:

**(a)Bleaching agent**

Wet/damp/moist coloured flowers/litmus papers are bleached/decolorized when put in sulphur(IV) oxide gas.

This is because sulphur(IV) oxide removes atomic oxygen from the coloured dye/ material to form sulphuric(VI)acid.

Chemical equations

(i)Formation of sulphuric(IV)acid

SO2(g) + H2O(l) -> H2 SO3 (aq)

(ii)Decolorization/bleaching of the dye/removal of atomic oxygen.

Method I. H2 SO3 (aq) + (dye + O) -> H2 SO4 (aq) + dye

(coloured) (colourless)

Method II. H2 SO3 (aq) + (dye) -> H2 SO4 (aq) + (dye - O)

(coloured) (colourless)

Sulphur(IV) oxide gas therefore bleaches by reduction /removing oxygen from a dye unlike chlorine that bleaches by oxidation /adding oxygen.

The bleaching by removing oxygen from Sulphur(IV) oxide gas is temporary. This is because the bleached dye regains the atomic oxygen from the atmosphere/air in presence of sunlight as catalyst thus regaining/restoring its original colour. e.g.

Old newspapers turn brown on exposure to air on regaining the atomic oxygen.

The bleaching through adding oxygen by chlorine gas is permanent.

**(b)Turns Orange acidified potassium dichromate(VI) to green**

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containingacidified potassium dichromate(VI) solution. or;

(ii)Dip a filter paper soaked in acidified potassium dichromate(VI) into a gas jar containing Sulphur(IV) oxide gas.

Observation:

Orange acidified potassium dichromate(VI) turns to green.

Explanation:

Sulphur(IV) oxide gas reduces acidified potassium dichromate(VI) from orange Cr2O72- ions to green Cr3+ ions without leaving a residue itself oxidized from SO32- ions in sulphuric(IV) acid to SO42- ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i)Reaction of Sulphur(IV) oxide gas with water

SO2(g) + H2O(l) -> H2 SO3 (aq)

(ii)Dissociation /ionization of Sulphuric(IV)acid.

H2 SO3 (aq) -> 2H+(aq) + SO32- (aq)

(iii)Oxidation of SO32- (aq)and reduction of Cr2O72-(aq)

3SO32-(aq)+ Cr2O72-(aq) +8H+(aq) -> 3SO42-(aq) + 2Cr3+(aq) + 4H2O(l)

This is a confirmatory test for the presence of Sulphur(IV) oxide gas.

Hydrogen sulphide also reduces acidified potassium dichromate(VI) from orange Cr2O72- ions to green Cr3+ ions leaving a **yellow** residue.

**(c)Decolorizes acidified potassium manganate(VII)**

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containingacidified potassium manganate(VII) solution. or;

(ii)Dip a filter paper soaked in acidified potassium manganate(VII) into a gas jar containing Sulphur(IV) oxide gas.

Observation:

Purple acidified potassium manganate(VII) turns to colourless/ acidified potassium manganate(VII) is decolorized.

Explanation:

Sulphur(IV) oxide gas reduces acidified potassium manganate(VII) from purple MnO4- ions to green Mn2+ ions without leaving a residue itself oxidized from SO32- ions in sulphuric(IV) acid to SO42- ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i)Reaction of Sulphur(IV) oxide gas with water

SO2(g) + H2O(l) -> H2 SO3 (aq)

(ii)Dissociation /ionization of Sulphuric(IV)acid.

H2 SO3 (aq) -> 2H+(aq) + SO32- (aq)

(iii)Oxidation of SO32- (aq)and reduction of MnO4- (aq)

5SO32-(aq)+ 2MnO4- (aq) +6H+(aq) -> 5SO42-(aq) + 2Mn2+(aq) + 3H2O(l)

(purple) (colourless)

This is another test for the presence of Sulphur(IV) oxide gas.

Hydrogen sulphide also decolorizes acidified potassium manganate(VII) from purple MnO4- ions to colourless Mn2+ ions leaving a **yellow** residue.

**(d)Decolorizes bromine water**

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containingbromine water . or;

(ii)Put three drops of bromine water into a gas jar containing Sulphur(IV) oxide gas. Swirl.

Observation:

Yellow bromine water turns to colourless/ bromine water is decolorized.

Explanation:

Sulphur(IV) oxide gas reduces yellow bromine water to colourless hydrobromic acid (HBr) without leaving a residue itself oxidized from SO32- ions in sulphuric(IV) acid to SO42- ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i)Reaction of Sulphur(IV) oxide gas with water

SO2(g) + H2O(l) -> H2 SO3 (aq)

(ii)Dissociation /ionization of Sulphuric(IV)acid.

H2 SO3 (aq) -> 2H+(aq) + SO32- (aq)

(iii)Oxidation of SO32- (aq)and reduction of MnO4- (aq)

SO32-(aq)+ Br2 (aq) + H2O(l) -> SO42-(aq) + 2HBr(aq)

(yellow) (colourless)

This can also be used as another test for the presence of Sulphur(IV) oxide gas.

Hydrogen sulphide also decolorizes yellow bromine water to colourless leaving a **yellow** residue.

**(e)Reduces Iron(III)** Fe3+ **salts to Iron(II) salts** Fe2+

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containingabout 3 cm3 of Iron (III)chloride solution. or;

(ii)Place about 3cm3 of Iron (III)chloride solution into a gas jar containing Sulphur(IV) oxide gas.Swirl.

Observation:

Yellow/brown Iron (III)chloride solution turns to green

Explanation:

Sulphur(IV) oxide gas reduces Iron (III)chloride solution from yellow/brown Fe3+ ions to green Fe2+ ions without leaving a residue itself oxidized from SO32- ions in sulphuric(IV) acid to SO42- ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i)Reaction of Sulphur(IV) oxide gas with water

SO2(g) + H2O(l) -> H2 SO3 (aq)

(ii)Dissociation /ionization of Sulphuric(IV)acid.

H2 SO3 (aq) -> 2H+(aq) + SO32- (aq)

(iii)Oxidation of SO32- (aq)and reduction of Fe3+ (aq)

SO32-(aq)+ 2Fe3+ (aq) +3H2O(l) -> SO42-(aq) + 2Fe2+(aq) + 2H+(aq)

(yellow) (green)

**(f)Reduces Nitric(V)acid to Nitrogen(IV)oxide gas**

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containingabout 3 cm3 of concentrated nitric(V)acid. or;

(ii)Place about 3cm3 of concentrated nitric(V)acid into a gas jar containing Sulphur(IV) oxide gas. Swirl.

Observation:

Brown fumes of a gas evolved/produced.

Explanation:

Sulphur(IV) oxide gas reduces concentrated nitric(V)acid to brown nitrogen(IV)oxide gas itself oxidized from SO32- ions in sulphuric(IV) acid to SO42- ions in sulphuric(VI) acid.

Chemical/ionic equation:

SO2(g) + 2HNO3 (l) -> H2 SO4 (l) + NO2 (g)

(brown fumes/gas)

**(g)Reduces Hydrogen peroxide to water**

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containingabout 3 cm3 of 20 volume hydrogen peroxide. Add four drops of Barium nitrate(V)or Barium chloride followed by five drops of 2M hydrochloric acid/ 2M nitric(V) acid.

Observation:

A white precipitate is formed that persist /remains on adding 2M hydrochloric acid/ 2M nitric(V) acid.

Explanation:

Sulphur(IV) oxide gas reduces 20 volume hydrogen peroxide and itself oxidized from SO32- ions in sulphuric(IV) acid to SO42- ions in sulphuric(VI) acid.

When Ba2+ ions in Barium Nitrate(V) or Barium chloride solution is added, a white precipitate of insoluble Barium salts is formed showing the presence of of either SO32- ,SO42- ,CO32- ions. i.e.

Chemical/ionic equation:

SO32-(aq)+ Ba2+ (aq) -> BaSO3(s)

white precipitate

SO42-(aq)+ Ba2+ (aq) -> BaSO4(s)

white precipitate

CO32-(aq)+ Ba2+ (aq) -> BaCO3(s)

white precipitate

If nitric(V)/hydrochloric acid is added to the three suspected insoluble white precipitates above, the white precipitate:

(i) persist/remains if SO42-(aq)ions (BaSO4(s)) is present.

(ii)dissolves if SO32-(aq)ions (BaSO3(s)) and CO32-(aq)ions (BaCO3(s))is present. This is because:

I. BaSO3(s) reacts with Nitric(V)/hydrochloric acid to produce acidic SO2 gas that turns Orange moist filter paper dipped in acidified Potassium dichromate to green.

Chemical equation

BaSO3(s) +2H+(aq)-> Ba2+ (aq) + SO2(g) + H2O(l)

I. BaCO3(s) reacts with Nitric(V)/hydrochloric acid to produce acidic CO2 gas that forms a white precipitate when bubbled in lime water.

Chemical equation

BaCO3(s) +2H+(aq)-> Ba2+ (aq) + CO2(g) + H2O(l)

**5.Sulphur(IV)oxide also act as an oxidizing agent as in the following examples.**

**(a)Reduction by burning Magnesium**

Experiment

Lower a burning Magnesium ribbon into agas jar containing Sulphur(IV)oxide gas

Observation

Magnesium ribbon continues to burn with difficulty.

White ash and yellow powder/speck

Explanation

Sulphur(IV)oxide does not support burning/combustion. Magnesium burns to produce enough heat energy to decompose Sulphur(IV)oxide to sulphur and oxygen.

The metal continues to burn on Oxygen forming white Magnesium oxide solid/ash.

Yellow specks of sulphur residue form on the sides of reaction flask/gas jar.

During the reaction, Sulphur(IV)oxide is reduced(oxidizing agent)while the metal is oxidized (reducing agent)

Chemical equation

SO2(g) + 2Mg(s) -> 2MgO(s) + S(s)

(white ash/solid) (yellow speck/powder)

**(b)Reduction by Hydrogen sulphide gas**

Experiment

Put two drops of water into a gas jar containing dry Sulphur(IV)oxide gas

Bubble hydrogen sulphide gas into the gas jar containing Sulphur(IV)oxide gas.

Or

Put two drops of water into a gas jar containing dry Sulphur(IV)oxide gas

Invert a gas jar full of hydrogen sulphide gas over the gas jar containing Sulphur(IV)oxide gas. Swirl

Observation

Yellow powder/speck

Explanation

Sulphur(IV)oxide oxidizes hydrogen sulphide to yellow specks of sulphur residue and itself reduced to also sulphur that form on the sides of reaction flask/gas jar.

A little moisture/water act as catalyst /speeds up the reaction.

Chemical equation

SO2(g) + 2H2S(g) -> 2H2O(l) + 3S(s)

(yellow speck/powder)

**6.Sulphur(IV)oxide has many industrial uses. State three.**

(i)In the contact process for the manufacture of Sulphuric(VI)acid

(ii)As a bleaching agent of pulp and paper.

(iii)As a fungicide to kill microbes’

(iv)As a preservative of jam, juices to prevent fermentation

**(ii) Sulphur(VI)oxide(SO3)**

**(a) Occurrence**

Sulphur (VI)oxide is does not occur free in nature/atmosphere

**(b) Preparation**

In a Chemistry school laboratory Sulphur (VI)oxide may prepared from:

Method 1;**Catalytic oxidation of sulphur(IV)oxide gas**.

Sulphur(IV)oxide gas and oxygen mixture are first dried by being passed through Concentrated Sulphuric(VI)acid .

The dry mixture is then passed through platinised asbestos to catalyse/speed up the combination to form Sulphur (VI)oxide gas.

Sulphur (VI)oxide gas readily solidify as silky white needles if passed through a freezing mixture /ice cold water.

The solid fumes out on heating to a highly acidic poisonous gas.

Chemical equation

2SO2(g) + O2(g) --platinised asbestos--> 2SO3 (g)

**Method 2; Heating Iron(II)sulphate(VI) heptahydrate**

When green hydrated Iron(II)sulphate(VI) heptahydrate crystals are heated in a boiling tube ,it loses the water of crystallization and colour changes from green to white.

Chemical equation

FeSO4.7H2O(s) -> FeSO4(s) + 7H2O(l)

(green solid) (white solid)

On further heating ,the white anhydrous Iron(II)sulphate(VI) solid decomposes to a mixture of Sulphur (VI)oxide and Sulphur (IV)oxide gas.

Sulphur (VI) oxide readily /easily solidify as white silky needles when the mixture is passed through a freezing mixture/ice cold water.

Iron(III)oxide is left as a brown residue/solid.

Chemical equation

2FeSO4 (s) -> Fe2O3(s) + SO2 (g) + SO3(g)

(green solid) (brown solid)

**Caution**

On exposure to air Sulphur (VI)oxide gas produces highly corrosive poisonous fumes of concentrated sulphuric(VI)acid and thus its preparation in a school laboratory is very risky.

**(c) Uses of sulphur(VI)oxide**

One of the main uses of sulphur(VI)oxide gas is as an intermediate product in the contact process for industrial/manufacture/large scale/production of sulphuric(VI)acid.

**(iii) Sulphuric(VI)acid(H2SO4)**

**(a) Occurrence**

Sulphuric(VI)acid(H2SO4) is one of the three mineral acids.There are three mineral acids;

Nitric(V)acid

Sulphuric(VI)acid

Hydrochloric acid.

Mineral acids do not occur naturally but are prepared in a school laboratory and manufactured at industrial level.

**(b)The Contact process for industrial manufacture of H2SO4 .**

I. Raw materials

The main raw materials for industrial preparation of Sulphuric(VI)acid include:

(i)**Sulphur** from Fraschs process or from heating metal sulphide ore like Galena(PbS),Zinc blende(ZnS)

(ii)**Oxygen** from fractional distillation of air

(iii)**Water** from rivers/lakes

II. Chemical processes

The contact process involve four main chemical processes:

**(i)Production of Sulphur(IV)oxide**

As one of the raw materials, Sulphur (IV)oxide gas is got from the following sources;

I. Burning/roasting sulphur in air.

**Sulphur** from Fraschs process is roasted/burnt in air to form Sulphur (IV)oxide gas in the **burners**

Chemical equation

S(s) + O2(g) --> SO2 (g)

II. Burning/roasting sulphide ores in air.

Sulphur (IV)oxide gas is produced as a by product in extraction of some metals like:

- Lead from Lead(II)sulphide/Galena,(PbS)

- Zinc from zinc(II)sulphide/Zinc blende, (ZnS)

- Copper from Copper iron sulphide/Copper pyrites, (CuFeS2)

On roasting/burning, large amount /quantity of sulphur(IV)oxide is generated/produced.

Chemical equation

(i)2PbS (s) + 3O2 (g) -> 2PbO(s) + 2SO2 (g)

(ii)2ZnS (s) + 3O2 (g) -> 2ZnO(s) + 2SO2 (g)

(ii)2CuFeS2 (s) + 4O2 (g) -> 2FeO(s) + 3SO2 (g) + Cu2O(s)

Sulphur(IV)oxide easily/readily liquefies and thus can be transported to a far distance safely.

**(ii)Purification of Sulphur(IV)oxide**

Sulphur(IV)oxide gas contain dust particles and Arsenic(IV)oxide as impurities. These impurities “poison”/impair the catalyst by adhering on/covering its surface.

The impurities are removed by electrostatic precipitation method .

In the contact process Platinum or Vanadium(V)oxide may be used. Vanadium(V)oxide is preferred because it is :

(i) cheaper/less expensive

(ii) less easily poisoned by impurities

**(iii)Catalytic conversion of Sulphur(IV)oxide to Sulphur(VI)oxide**

Pure and dry mixture of Sulphur (IV)oxide gas and Oxygen is heated to 450oC in a heat exchanger.

The heated mixture is passed through long pipes coated with pellets of Vanadium (V)oxide catalyst.

The close “contact” between the reacting gases and catalyst give the process its name.

Vanadium (V)oxide catalyse the conversion/oxidation of Sulphur(IV)oxide to Sulphur(VI)oxide gas.

Chemical equation

2SO2 (g) + O2(g) -- V2O5 --> 2SO2 (g)

This reaction is exothermic (-∆H) and the temperatures need to be maintained at around 450oC to ensure that:

(**i**)reaction rate/time taken for the formation of Sulphur(VI)oxide is not too **slow/long** at **lower** temperatures below 450oC

(**ii**) Sulphur(VI)oxide gas does not **decompose** back to Sulphur(IV)oxide gas and Oxygen gas at **higher** temperatures than 450oC.

**(iii)Conversion of Sulphur(VI)oxide of Sulphuric(VI)acid**

Sulphur(VI)oxide is the acid anhydride of concentrated Sulphuric(VI)acid. Sulphur(VI)oxide reacts with water to form thick mist of fine droplets of very/highly corrosive concentrated Sulphuric(VI)acid because the reaction is highly exothermic.

To prevent this, Sulphur (VI)oxide is a passed up to meet downward flow of 98% Sulphuric(VI)acid in the absorption chamber/tower.

The reaction forms a very **viscous oily** liquid called **Oleum/fuming Sulphuric (VI) acid/ pyrosulphuric (VI) acid.**

Chemical equation

H2SO4 (aq) + SO3 (g) -> H2S2O7 (l)

Oleum/fuming Sulphuric (VI) acid/ pyrosulphuric (VI) acid is diluted carefully with distilled water to give concentrated sulphuric (VI) acid .

Chemical equation

H2S2O7 (l) + H2O(l) -> 2H2SO4 (l)

The acid is stored ready for market/sale.

III. Environmental effects of contact process

Sulphur(VI)oxide and Sulphur(IV)oxide gases are atmospheric pollutants that form acid rain if they escape to the atmosphere.

In the Contact process, about 2% of these gases do not form sulphuric (VI) acid.

The following precautions prevent/minimize pollution from Contact process:

(i)recycling back any unreacted Sulphur(IV)oxide gas back to the heat exchangers.

(ii)dissolving Sulphur(VI)oxide gas in concentrated sulphuric (VI) acid instead of water.

This prevents the formation of fine droplets of the corrosive/ toxic/poisonous fumes of concentrated sulphuric (VI) acid.

(iii)scrubbing-This involves passing the exhaust gases through very tall chimneys lined with quicklime/calcium hydroxide solid.

This reacts with Sulphur (VI)oxide gas forming harmless calcium(II)sulphate (IV) /CaSO3

Chemical equation

Ca(OH)2 (aq) + SO2(g) --> CaSO3 (aq) + H2O (g)

III. Uses of Sulphuric(VI)acid

Sulphuric (VI) acid is used:

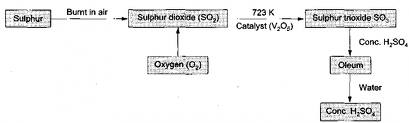
(i) in making dyes and paint

(ii)as acid in Lead-acid accumulator/battery

(iii) for making soapless detergents

(iv) for making sulphate agricultural fertilizers

III. Sketch chart diagram showing the Contact process



**(c) Properties of Concentrated sulphuric(VI)acid**  (i)Concentratedsulphuric(VI)acid is a colourless oily liquid with a density of 1.84gcm-3.It has a boiling point of 338oC.

(ii) Concentratedsulphuric(VI)acid is very soluble in water.

The solubility /dissolution of the acid very highly exothermic.

The concentrated acid should thus be diluted slowly in excess water.

Water should never be added to the acid because the hot acid scatters highly corrosive fumes out of the container.

(iii) Concentratedsulphuric (VI)acid is a covalent compound. It has no free H+ ions.

Free H+ ions are responsible for turning the blue litmus paper red. Concentratedsulphuric (VI) acid thus do not change the blue litmus paper red.

(iv) Concentratedsulphuric (VI)acid is hygroscopic. It absorbs water from the atmosphere and do not form a solution.

This makes concentratedsulphuric (VI) acid very suitable as drying agent during preparation of gases.

(v)The following are some chemical properties of concentratedsulphuric (VI) acid:

I. As a dehydrating agent

**Experiment I;**

Put about two spatula end fulls of brown sugar and glucose in separate boiling tubes .Carefully add about 10cm3 of concentratedsulphuric (VI) acid .Allow to stand for about 10 minutes.

Observation;

Colour in brown sugar change from brown to black.

Colour in glucose change from white to black.

Boiling tube becomes very hot.

Explanation

Concentratedsulphuric (VI) acid is strong dehydrating agent.

It removes chemically and physically combined elements of water(Hydrogen and Oxygen in ratio 2:1)from compounds.

When added to sugar /glucose a vigorous reaction that is highly exothermic take place.

The sugar/glucose is charred to black mass of carbon because the acid dehydrates the sugar/glucose leaving carbon.

Chemical equation

Glucose: C6H12O6(s) --conc.H2SO4--> 6C (s) + 6H2O(l)

(white) (black)

Sugar: C12H22O11(s) --conc.H2SO4--> 12C (s) +11H2O(l)

(brown) (black)

**Experiment II;**

Put about two spatula end full of hydrated copper(II)sulphate(VI)crystals in a boiling tube .Carefully add about 10cm3 of concentratedsulphuric (VI) acid .Warm .

Observation;

Colour change from blue to white.

Explanation

Concentratedsulphuric (VI) acid is strong dehydrating agent.It removes physically combined elements of water(Hydrogen and Oxygen in ratio 2:1)from hydrated compounds.

The acid dehydrates blue copper(II)sulphate to white anhydrous copper(II)sulphate .

Chemical equation

CuSO4.5H2O(s) --conc.H2SO4--> CuSO4 (s) + 5H2O(l)

(blue) (white)

**Experiment III;**

Put about 4cm3 of absolute ethanol in a boiling tube .Carefully add about 10cm3 of concentratedsulphuric (VI) acid. Place moist/damp/wet filter paper dipped in acidified potassium dichromate(VI)solution on the mouth of the boiling tube. Heat strongly.

Caution: Absolute ethanol is highly flammable.

Observation;

Colourless gas produced.

Orange acidified potassium dichromate (VI) paper turns to green.

Explanation

Concentratedsulphuric (VI) acid is strong dehydrating agent.

It removes chemically combined elements of water(Hydrogen and Oxygen in ratio 2:1)from compounds.

The acid dehydrates ethanol to ethene gas at about 170oC.

Ethene with =C=C= double bond turns orange acidified potassium dichromate (VI) paper turns to green.

Chemical equation

C2H5OH(l) --conc.H2SO4/170oC --> C2H4 (g) + H2O(l)

NB: This reaction is used for the school laboratory preparation of ethene gas

**Experiment IV;**

Put about 4cm3 of methanoic acid in a boiling tube .Carefully add about 6 cm3 of concentratedsulphuric (VI) acid. Heat gently

Caution: This should be done in a fume chamber/open

Observation;

Colourless gas produced.

Explanation

Concentratedsulphuric (VI) acid is strong dehydrating agent. It removes chemically combined elements of water (Hydrogen and Oxygen in ratio 2:1)from compounds.

The acid dehydrates methanoic acid to poisonous/toxic carbon(II)oxide gas.

Chemical equation

HCOOH(l) --conc.H2SO4 --> CO(g) + H2O(l)

NB: This reaction is used for the school laboratory preparation of small amount carbon (II)oxide gas

**Experiment V;**

Put about 4cm3 of ethan-1,2-dioic/oxalic acid in a boiling tube .Carefully add about 6 cm3 of concentratedsulphuric (VI) acid. Pass any gaseous product through lime water.Heat gently

Caution: This should be done in a fume chamber/open

Observation;

Colourless gas produced.

Gas produced forms a white precipitate with lime water.

Explanation

Concentratedsulphuric (VI) acid is strong dehydrating agent.

It removes chemically combined elements of water (Hydrogen and Oxygen in ratio 2:1)from compounds.

The acid dehydrates ethan-1,2-dioic/oxalic acid to a mixture of poisonous/toxic carbon(II)oxide and carbon(IV)oxide gases.

Chemical equation

HOOCCOOH(l) --conc.H2SO4 --> CO(g) + CO2(g) + H2O(l)

NB: This reaction is also used for the school laboratory preparation of small amount carbon (II) oxide gas.

Carbon (IV) oxide gas is removed by passing the mixture through concentrated sodium/potassium hydroxide solution.

II. As an Oxidizing agent

**Experiment I**

Put about 2cm3 of Concentratedsulphuric (VI) acid into three separate boiling tubes. Place a thin moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution on the mouth of the boiling tube. Put about 0.5g of Copper turnings, Zinc granule and Iron filings to each boiling tube separately.

Observation;

Effervescence/fizzing/bubbles

Blue solution formed with copper,

Green solution formed with Iron

Colourless solution formed with Zinc

Colourless gas produced that has a pungent irritating choking smell.

Gas produced turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

Explanation

Concentratedsulphuric (VI) acid is strong oxidizing agent.

It oxidizes metals to metallic sulphate(VI) salts and itself reduced to sulphur(IV)oxide gas.

Sulphur (IV) oxide gas turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

CuSO4(aq) is a blue solution. ZnSO4(aq) is a colourless solution. FeSO4(aq) is a green solution.

Chemical equation

Cu(s) + 2H2SO4(aq) --> CuSO4(aq) + SO2(g) + 2H2O(l)

Zn(s) + 2H2SO4(aq) --> ZnSO4(aq) + SO2(g) + 2H2O(l)

Fe(s) + 2H2SO4(aq) --> FeSO4(aq) + SO2(g) + 2H2O(l)

**Experiment II**

Put about 2cm3 of Concentratedsulphuric (VI) acid into two separate boiling tubes. Place a thin moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution on the mouth of the boiling tube.

Put about 0.5g of powdered charcoal and sulphur powder to each boiling tube separately.

Warm.

Observation;

Black solid charcoal dissolves/decrease

Yellow solid sulphur dissolves/decrease

Colourless gas produced that has a pungent irritating choking smell.

Gas produced turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

Explanation

Concentratedsulphuric (VI) acid is strong oxidizing agent. It oxidizes non-metals to non metallic oxides and itself reduced to sulphur(IV)oxide gas. Sulphur (IV) oxide gas turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

Charcoal is oxidized to carbon(IV)oxide. Sulphur is oxidized to Sulphur(IV)oxide .

Chemical equation

C(s) + 2H2SO4(aq) --> CO2(aq) + 2SO2(g) + 2H2O(l)

S(s) + 2H2SO4(aq) --> 3SO2(g) + 2H2O(l)

III. As the least volatile acid

Study the table below showing a comparison in boiling points of the three mineral acids

|  |  |  |
| --- | --- | --- |
| Mineral acid | Relative molecula mass | Boiling point(oC) |
| Hydrochloric acid(HCl) | 36.5 | 35.0 |
| Nitric(V)acid(HNO3) | 63.0 | 83.0 |
| Sulphuric(VI)acid(H2SO4) | 98.0 | 333 |

**1.Which is the least volatile acid? Explain**

Sulphuric(VI)acid(H2SO4) because it has the largest molecule and joined by Hydrogen bonds making it to have the highest boiling point/least volatile.

**2. Using chemical equations, explain how sulphuric(VI)acid displaces the less volatile mineral acids.**

(i)Chemical equation

KNO3(s) + H2SO4(aq) --> KHSO4(l) + HNO3(g)

NaNO3(s) + H2SO4(aq) --> NaHSO4(l) + HNO3(g)

This reaction is used in the school laboratory preparation of Nitric(V) acid (HNO3).

(ii)Chemical equation

KCl(s) + H2SO4(aq) --> KHSO4(s) + HCl(g)

NaCl(s) + H2SO4(aq) --> NaHSO4(s) + HCl(g)

This reaction is used in the school laboratory preparation of Hydrochloric acid (HCl).

**(d) Properties of dilute sulphuric(VI)acid.**

Dilute sulphuric(VI)acid is made when about 10cm3 of concentrated sulphuric

(VI) acid is carefully added to about 90cm3 of distilled water.

Diluting concentrated sulphuric (VI) acid should be done carefully because the reaction is highly exothermic.

Diluting concentrated sulphuric (VI) acid decreases the number of moles present in a given volume of solution which makes the acid less corrosive.

On diluting concentrated sulphuric(VI) acid, water ionizes /dissociates the acid fully/wholly into two(**dibasic**)free H+(aq) and SO42-(aq)ions:

H2SO4 (aq) -> 2H+(aq) + SO42-(aq)

The presence of free H+(aq)ions is responsible for ;

(**i**)turn litmus red because of the presence of free H+(aq)ions

(ii)have pH 1/2/3 because of the presence of many free H+(aq)ions hence a strongly acidic solution.

(iii)Reaction with metals

Experiment:

Place 5cm3 of 0.2M dilute sulphuric(VI)acid into four separate clean test tubes. Add about 0.1g of Magnesium ribbon to one test tube. Cover the mixture with a finger as stopper. Introduce a burning splint on top of the finger and release the finger “stopper”. Repeat by adding Zinc, Copper and Iron instead of the Magnesium ribbon.

Observation:

**No** effervescence/ bubbles/ fizzing with copper

Effervescence/ bubbles/ fizzing with Iron ,Zinc and Magnesium

Colourless gas produced that extinguishes burning splint with a “pop” sound.

Colourless solution formed with Zinc and Magnesium.

Green solution formed with Iron

Explanation:

When a metal higher than hydrogen in the reactivity/electrochemical series is put in a test tube containing dilute sulphuric(VI)acid, effervescence/ bubbling/ fizzing takes place with evolution of Hydrogen gas.

Impure hydrogen gas extinguishes burning splint with a “pop” sound.

A sulphate (VI) salts is formed. Iron, Zinc and Magnesium are higher than hydrogen in the reactivity/electrochemical series.

They form Iron (II)sulphate(VI), Magnesium sulphate(VI) and Zinc sulphate(VI).

. When a metal lower than hydrogen in the reactivity/electrochemical series is put in a test tube containing dilute sulphuric(VI)acid, there is no effervescence/ bubbling/ fizzing that take place.

Copper thus do not react with dilute sulphuric(VI)acid.

Chemical/ionic equation

Mg(s) + H2SO4(aq) --> MgSO4(aq) + H2(g)

Mg(s) + 2H+(aq) --> Mg2+ (aq) + H2(g)

Zn(s) + H2SO4(aq) --> ZnSO4(aq) + H2(g)

Zn(s) + 2H+(aq) --> Zn2+ (aq) + H2(g)

Fe(s) + H2SO4(aq) --> FeSO4(aq) + H2(g)

Fe(s) + H+(aq) --> Fe2+ (aq) + H2(g)

NB:(i) Calcium,Lead and Barium forms insoluble sulphate(VI)salts that cover/coat the unreacted metals.

(ii)Sodium and Potassium react explosively with dilute sulphuric(VI)acid

(**iv**)Reaction with metal carbonates and hydrogen carbonates

Experiment:

Place 5cm3 of 0.2M dilute sulphuric(VI)acid into four separate clean boiling tubes. Add about 0.1g of sodium carbonate to one boiling tube. Introduce a burning splint on top of the boiling tube. Repeat by adding Zinc carbonate, Copper (II)carbonate and Iron(II)Carbonate in place of the sodium hydrogen carbonate.

Observation:

Effervescence/ bubbles/ fizzing.

Colourless gas produced that extinguishes burning splint.

Colourless solution formed with Zinc carbonate, sodium hydrogen carbonate and sodium carbonate.

Green solution formed with Iron(II)Carbonate

Blue solution formed with Copper(II)Carbonate

Explanation:

When a metal carbonate or a hydrogen carbonates is put in a test tube containing dilute sulphuric(VI)acid, effervescence/ bubbling/ fizzing takes place with evolution of carbon(IV)oxide gas. carbon(IV)oxide gas extinguishes a burning splint and forms a white precipitate when bubbled in lime water.

A sulphate (VI) salts is formed.

Chemical/ionic equation

ZnCO3(s) + H2SO4(aq) --> ZnSO4(aq) + H2O(l) + CO2(g)

ZnCO3(s) + 2H+(aq) --> Zn2+ (aq) + H2O(l) + CO2(g)

CuCO3(s) + H2SO4(aq) --> CuSO4(aq) + H2O(l) + CO2(g)

CuCO3(s) + 2H+(aq) --> Cu2+ (aq) + H2O(l) + CO2(g)

FeCO3(s) + H2SO4(aq) --> FeSO4(aq) + H2O(l) + CO2(g)

FeCO3(s) + 2H+(aq) --> Fe2+ (aq) + H2O(l) + CO2(g)

2NaHCO3(s) + H2SO4(aq) --> Na2SO4(aq) + 2H2O(l) + 2CO2(g)

NaHCO3(s) + H+(aq) --> Na+ (aq) + H2O(l) + CO2(g)

Na2CO3(s) + H2SO4(aq) --> Na2SO4(aq) + H2O(l) + CO2(g)

NaHCO3(s) + H+(aq) --> Na+ (aq) + H2O(l) + CO2(g)

(NH4)2CO3(s) + H2SO4(aq) --> (NH4)2SO4 (aq) + H2O(l) + CO2(g)

(NH4)2CO3 (s) + H+(aq) --> NH4+ (aq) + H2O(l) + CO2(g)

2NH4HCO3(aq) + H2SO4(aq) --> (NH4)2SO4 (aq) + H2O(l) + CO2(g)

NH4HCO3(aq) + H+(aq) --> NH4+ (aq) + H2O(l) + CO2(g)

NB:

Calcium, Lead and Barium carbonates forms insoluble sulphate(VI)salts that cover/coat the unreacted metals.

(**v**)Neutralization-reaction of metal oxides and alkalis/bases

Experiment I:

Place 5cm3 of 0.2M dilute sulphuric(VI)acid into four separate clean boiling tubes. Add about 0.1g of copper(II)oxide to one boiling tube. Stir.

Repeat by adding Zinc oxide, calcium carbonate and Sodium (II)Oxide in place of the Copper(II)Oxide.

Observation:

Blue solution formed with Copper(II)Oxide

Colourless solution formed with other oxides

Explanation:

When a metal oxide is put in a test tube containing dilute sulphuric(VI)acid, the oxide dissolves forming a sulphate (VI) salt.

Chemical/ionic equation

ZnO(s) + H2SO4(aq) --> ZnSO4(aq) + H2O(l)

ZnO(s) + 2H+(aq) --> Zn2+ (aq) + H2O(l)

CuO(s) + H2SO4(aq) --> CuSO4(aq) + H2O(l)

CuO(s) + 2H+(aq) --> Cu2+ (aq) + H2O(l)

MgO(s) + H2SO4(aq) --> MgSO4(aq) + H2O(l)

MgO(s) + 2H+(aq) --> Mg2+ (aq) + H2O(l)

Na2O(s) + H2SO4(aq) --> Na2SO4(aq) + H2O(l)

Na2O(s) + 2H+(aq) --> 2Na+ (aq) + H2O(l)

K2CO3(s) + H2SO4(aq) --> K2SO4(aq) + H2O(l)

K2O(s) + H+(aq) --> 2K+ (aq) + H2O(l)

NB:

Calcium, Lead and Barium oxides forms insoluble sulphate(VI)salts that cover/coat the unreacted metals oxides.

Experiment II:

Fill a burette wuth 0.1M dilute sulphuric(VI)acid. Pipette 20.0cm3 of 0.1Msodium hydroxide solution into a 250cm3 conical flask. Add three drops of phenolphthalein indicator.Titrate the acid to get a permanent colour change.Repeat with0.1M potassium hydroxide solution inplace of 0.1Msodium hydroxide solution

Observation:

Colour of phenolphthalein changes from pink to colourless at the end point.

Explanation

Like other (mineral) acids dilute sulphuric(VI)acid neutralizes bases/alkalis to a sulphate salt and water only.

Colour of the indicator used changes when a slight excess of acid is added to the base at the end point

Chemical equation:

2NaOH(aq) + H2SO4(aq) --> Na2SO4(aq) + H2O(l)

OH-(s) + H+(aq) --> H2O(l)

2KOH(aq) + H2SO4(aq) --> K2SO4(aq) + H2O(l)

OH-(s) + H+(aq) --> H2O(l)

2NH4OH(aq) + H2SO4(aq) --> (NH4)2SO4(aq) + H2O(l)

OH-(s) + H+(aq) --> H2O(l)

**(iv) Hydrogen sulphide(H2S)**

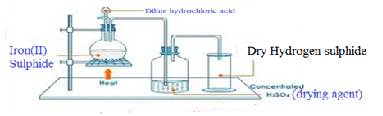
**(a) Occurrence**

Hydrogen sulphide is found in volcanic areas as a gas or dissolved in water from geysers and hot springs in active volcanic areas of the world e.g. Olkaria and Hells gate near Naivasha in Kenya.

It is present in rotten eggs and human excreta.

**(b) Preparation**

Hydrogen sulphide is prepared in a school laboratory by heating Iron (II) sulphide with dilute hydrochloric acid.



**(c) Properties of Hydrogen sulphide(Questions)**

**1. Write the equation for the reaction for the school laboratory preparation of Hydrogen sulphide.**

Chemical equation: FeS (s) + 2HCl (aq) -> H2S (g) FeCl2 (aq)

**2. State three physical properties unique to Hydrogen sulphide.**

Hydrogen sulphide is a colourless gas with characteristic pungent poisonous smell of rotten eggs. It is soluble in cold water but insoluble in warm water. It is denser than water and turns blue litmus paper red.

**3. Hydrogen sulphide exist as a dibasic acid when dissolved in water. Using a chemical equation show how it ionizes in aqueous state.**

H2S(aq) **->** H+(aq) + HS-(aq)

H2S(aq) **-> 2**H+(aq) + S2- (aq)

Hydrogen sulphide therefore can form both normal and acid salts e.g

Sodium hydrogen sulphide and sodium sulphide both exist

**4. State and explain one gaseous impurity likely to be present in the gas jar containing hydrogen sulphide above.**

Hydrogen/ H2

Iron(II)sulphide contains Iron as impurity .The iron will react with dilute hydrochloric acid to form iron(II)chloride and produce hydrogen gas that mixes with hydrogen sulphide gas.

**5. State and explain the observations made when a filter paper dipped in Lead(II) ethanoate /Lead (II) nitrate(V) solution is put in a gas jar containing hydrogen sulphide gas.**

Observations

Moist Lead(II) ethanoate /Lead (II) nitrate(V) paper turns black.

Explanation

When hydrogen sulphide is bubbled in a metallic salt solution, a metallic sulphide is formed.

All sulphides are insoluble black salts except sodium sulphide, potassium sulphide and ammonium sulphides.

Hydrogen sulphide gas blackens moist Lead (II) ethanoate /Lead (II) nitrate(V) paper .

The gas reacts with Pb2+ in the paper to form black Lead(II)sulphide.

This is the chemical test for the presence of H2S other than the physical smell of rotten eggs.

Chemical equations

Pb2+(aq) + H2S -> PbS + 2H+(aq)

(black)

Fe2+(aq) + H2S -> FeS + 2H+(aq)

(black)

Zn2+(aq) + H2S -> ZnS + 2H+(aq)

(black)

Cu2+(aq) + H2S -> CuS + 2H+(aq)

(black)

2Cu+(aq) + H2S -> Cu2S + 2H+(aq)

(black)

**6. Dry hydrogen sulphide was ignited as below.**

Dry Hydrogen sulphide gas

Flame A

**(i) State the observations made in flame A**

Hydrogen sulphide burns in excess air with a blue flame to form sulphur(IV)oxide gas and water.

Chemical equation: 2H2S(g) + 3O2(g) -> 2H2O(l) + 2SO2(g)

Hydrogen sulphide burns in limited air with a blue flame to form sulphur solid and water.

Chemical equation: 2H2S(g) + O2(g) -> 2H2O(l) + 2S(s)

**7. Hydrogen sulphide is a strong reducing agent that is oxidized to yellow solid sulphur as precipitate. The following experiments illustrate the reducing properties of Hydrogen sulphide.**

**(a)Turns Orange acidified potassium dichromate(VI) to green**

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containingacidified potassium dichromate (VI) solution. or;

(ii)Dip a filter paper soaked in acidified potassium dichromate (VI) into a gas jar containing Hydrogen sulphide gas.

Observation:

Orange acidified potassium dichromate (VI) turns to green.

Yellow solid residue.

Explanation:

Hydrogen sulphide gas reduces acidified potassium dichromate(VI) from orange Cr2O72- ions to green Cr3+ ions leaving a yellow solid residue as itself is oxidized to sulphur.

Chemical/ionic equation:

4H2S(aq)+ Cr2O72-(aq) +6H+(aq) -> 4S(aq) + 2Cr3+(aq) + 7H2O(l)

This test is used for differentiating Hydrogen sulphide and sulphur (IV)oxide gas.

Sulphur(IV)oxide also reduces acidified potassium dichromate(VI) from orange Cr2O72- ions to green Cr3+ ions without leaving a **yellow** residue.

**(b)Decolorizes acidified potassium manganate(VII)**

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containingacidified potassium manganate(VII) solution. or;

(ii)Dip a filter paper soaked in acidified potassium manganate(VII) into a gas jar containing Hydrogen Sulphide gas.

Observation:

Purple acidified potassium manganate(VII) turns to colourless/ acidified potassium manganate(VII) is decolorized.

Yellow solid residue.

Explanation:

Hydrogen sulphide gas reduces acidified potassium manganate(VII) from purple MnO4- ions to green Mn2+ ions leaving a residue as the gas itself is oxidized to sulphur.

Chemical/ionic equation:

5H2S(g)+ 2MnO4- (aq) +6H+(aq) -> 5S (s) + 2Mn2+(aq) + 8H2O(l)

(purple) (colourless)

This is another test for differentiating Hydrogen sulphide and Sulphur(IV) oxide gas.

Sulphur(IV) oxide also decolorizes acidified potassium manganate(VII) from purple MnO4- ions to colourless Mn2+ ions leaving no  **yellow** residue.

**(c)Decolorizes bromine water**

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containingbromine water . or;

(ii)Put three drops of bromine water into a gas jar containing Hydrogen sulphide gas. Swirl.

Observation:

Yellow bromine water turns to colourless/ bromine water is decolorized.

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces yellow bromine water to colourless hydrobromic acid (HBr) leaving a yellow residue as the gas itself is oxidized to sulphur.

Chemical/ionic equation:

H2 S(g)+ Br2 (aq) -> S (s) + 2HBr(aq)

(yellow solution) (yellow solid) (colourless)

This is another test for differentiating Hydrogen sulphide and Sulphur(IV) oxide gas.

Sulphur(IV) oxide also decolorizes acidified potassium manganate(VII) from purple MnO4- ions to colourless Mn2+ ions leaving no  **yellow** residue.

**(d)Reduces Iron(III)** Fe3+ **salts to Iron(II) salts** Fe2+

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containingabout 3 cm3 of Iron (III)chloride solution. or;

(ii)Place about 3cm3 of Iron (III)chloride solution into a gas jar containing Hydrogen sulphide gas. Swirl.

Observation:

Yellow/brown Iron (III)chloride solution turns to green.

Yellow solid solid

Explanation:

Hydrogen sulphide gas reduces Iron (III)chloride solution from yellow/brown Fe3+ ions to green Fe2+ ions leaving a yellow residue.The gas is itself oxidized to sulphur.

Chemical/ionic equation:

H2S(aq)+ 2Fe3+ (aq) -> S (s) + Fe2+(aq) + 2H+(aq)

(yellow solution) (yellow residue) (green)

**(e)Reduces Nitric(V)acid to Nitrogen(IV)oxide gas**

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containingabout 3 cm3 of concentrated nitric(V)acid. or;

(ii)Place about 3cm3 of concentrated nitric(V)acid into a gas jar containing Hydrogen sulphide gas. Swirl.

Observation:

Brown fumes of a gas evolved/produced.

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces concentrated nitric(V)acid to brown nitrogen(IV)oxide gas itself oxidized to yellow sulphur.

Chemical/ionic equation:

H2S(g) + 2HNO3 (l) -> 2H2O(l) + S (s) + 2NO2 (g)

(yellow residue) (brown fumes)

**(f)Reduces sulphuric(VI)acid to Sulphur**

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containingabout 3 cm3 of concentrated sulphuric(VI)acid. or;

(ii)Place about 3cm3 of concentrated sulphuric (VI) acid into a gas jar containing Hydrogen sulphide gas. Swirl.

Observation:

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces concentrated sulphuric(VI)acid to yellow sulphur.

Chemical/ionic equation:

3H2S(g) + H2SO4 (l) -> 4H2O(l) + 4S (s)

(yellow residue)

**(g)Reduces Hydrogen peroxide to water**

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containingabout 3 cm3 of 20 volume hydrogen peroxide.

Observation:

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces 20 volume hydrogen peroxide to water and itself oxidized to yellow sulphur

Chemical/ionic equation:

H2S(g) + H2O2 (l) -> 2H2O(l) + S (s)

(yellow residue)

**8.Name the salt formed when:**

**(i)equal volumes of equimolar hydrogen sulphide neutralizes sodium hydroxide solution:**

Sodium hydrogen sulphide

Chemical/ionic equation:

H2S(g) + NaOH (l) -> H2O(l) + NaHS (aq)

**(ii) hydrogen sulphide neutralizes excess concentrated sodium hydroxide solution:**

Sodium sulphide

Chemical/ionic equation:

H2S(g) + 2NaOH (l) -> 2H2O(l) + Na2S (aq)

**Practice**

**Hydrogen sulphide gas was bubbled into a solution of metallic nitrate(V)salts as in the flow chart below**

Hydrogen sulphide

Brown solution

Blue solution

Green solution

Black solid

**(a)Name the black solid**  Copper(II)sulphide

**(b)Identify the cation responsible for the formation of:**

**I. Blue solution** Cu2+(aq)

**II. Green solution** Fe2+(aq)

**III. Brown solution** Fe3+(aq)

**(c)Using acidified potassium dichromate(VI) describe how you would differentiate between sulphur(IV)Oxide and hydrogen sulphide**

-Bubble the gases in separate test tubes containing acidified Potassium dichromate(VI) solution.

-Both changes the Orange colour of acidified Potassium dichromate(VI) solution to green.

-Yellow solid residue/deposit is formed with Hydrogen sulphide

Chemical/ionic equation:

4H2S(aq)+ Cr2O72-(aq) +6H+(aq) -> 4S(aq) + 2Cr3+(aq) + 7H2O(l)

3SO32-(aq)+ Cr2O72-(aq) +8H+(aq) -> 3SO42-(aq) + 2Cr3+(aq) + 4H2O(l)

**(d)State and explain the observations made if a burning splint is introduced at the mouth of a hydrogen sulphide generator.**

ObservationGas continues burning with a blue flame

Explanation: Hydrogen sulphide burns in excess air with a blue flame to form sulphur(IV)oxide gas and water.

Chemical equation: 2H2S(g)+ 3O2(g) -> 2H2O(l) + 2SO2 (g)

**(v)Sulphate (VI) (SO42-)and Sulphate(IV) (SO32-) salts**

1. Sulphate (VI) (SO42-) salts are normal and acid salts derived from Sulphuric (VI)acid H2SO4.

2. Sulphate(IV) (SO32-) salts are normal and acid salts derived from Sulphuric (IV)acid H2SO3.

3. Sulphuric (VI)acid H2SO4 is formed when sulphur(VI)oxide gas is bubbled in water.

The acid exist as a dibasic acid with two ionisable hydrogen. It forms therefore the Sulphate (VI) (SO42-) and hydrogen sulphate (VI) (HSO4-) salts.

i.e.

H2SO4 (aq)-> 2H+(aq) + SO42-(aq)

H2SO4 (aq)-> H+(aq) + HSO4 -(aq)

All Sulphate (VI) (SO42-) salts **dissolve** in water/are soluble except Calcium (II) sulphate (VI) (CaSO4), Barium (II) sulphate (VI) (BaSO4) and Lead (II) sulphate (VI) (PbSO4)

All Hydrogen sulphate (VI) (HSO3-) salts **exist** in solution/dissolved in water. Sodium (I) hydrogen sulphate (VI) (NaHSO4), Potassium (I) hydrogen sulphate (VI) (KHSO4) and Ammonium hydrogen sulphate (VI) (NH4HSO4) exist also as solids.

Other Hydrogen sulphate (VI) (HSO4-) salts do not **exist** except those of Calcium (II) hydrogen sulphate (VI) (Ca (HSO4)2) and Magnesium (II) hydrogen sulphate (VI) (Mg (HSO4)2).

4. Sulphuric (IV)acid H2SO3 is formed when sulphur(IV)oxide gas is bubbled in water.

The acid exist as a dibasic acid with two ionisable hydrogen. It forms therefore the Sulphate (IV) (SO32-) and hydrogen sulphate (VI) (HSO4-) salts.

i.e.

H2SO3 (aq)-> 2H+(aq) + SO32-(aq)

H2SO3 (aq)-> H+(aq) + HSO3 -(aq)

All Sulphate (IV) (SO32-) salts **dissolve** in water/are soluble except Calcium (II) sulphate (IV) (CaSO3), Barium (II) sulphate (IV) (BaSO3) and Lead (II) sulphate (IV) (PbSO3)

All Hydrogen sulphate (IV) (HSO3-) salts **exist** in solution/dissolved in water. Sodium (I) hydrogen sulphate (IV) (NaHSO3), Potassium (I) hydrogen sulphate (IV) (KHSO3) and Ammonium hydrogen sulphate (IV) (NH4HSO3) exist also as solids.

Other Hydrogen sulphate (IV) (HSO3-) salts do not **exist** except those of Calcium (II) hydrogen sulphate (IV) (Ca (HSO3)2) and Magnesium (II) hydrogen sulphate (IV) (Mg (HSO3)2).

5.The following experiments show the effect of heat on sulphate(VI) (SO42-)and sulphate(IV) (SO32-) salts:

Experiment:

In a clean dry test tube place separately about 1.0g of :

Zinc(II)sulphate (VI), Iron(II)sulphate(VI), Copper(II)sulphate(VI),Sodium (I) sulphate (VI), Sodium (I) sulphate (IV).Heat gently then strongly. Test any gases produced using litmus papers.

Observations:

-Colourless droplets of liquid forms on the cooler parts of the test tube in all cases.

-White solid residue is left in case of Zinc (II)sulphate(VI),Sodium (I) sulphate (VI) and Sodium (I) sulphate (IV).

-Colour changes from green to brown /yellow in case of Iron (II)sulphate(VI)

-Colour changes from blue to white then black in case of Copper (II) sulphate (VI)

-Blue litmus paper remain and blue and red litmus paper remain red in case of Zinc(II)sulphate(VI), Sodium (I) sulphate (VI) and Sodium (I) sulphate (IV)

-Blue litmus paper turns red and red litmus paper remain red in case of Iron (II)sulphate(VI) and Copper (II) sulphate (VI).

Explanation

(i)All Sulphate (VI) (SO42-) salts exist as **hydrated** salts with water of crystallization that condenses and collects on cooler parts of test tube as a colourless liquid on gentle heating. e.g.

K2SO4.10H2O(s) -> K2SO4(s) + 10H2O(l)

Na2SO4.10H2O(s) -> Na2SO4(s) + 10H2O(l)

MgSO4.7H2O(s) -> MgSO4(s) + 7H2O(l)

CaSO4.7H2O(s) -> CaSO4(s) + 7H2O(l)

ZnSO4.7H2O(s) -> ZnSO4(s) + 7H2O(l)

FeSO4.7H2O(s) -> FeSO4(s) + 7H2O(l)

Al2(SO4)3.6H2O(s) -> Al2(SO4)3 (s) + 6H2O(l)

CuSO4.5H2O(s) -> CuSO4(s) + 5H2O(l)

All Sulphate (VI) (SO42-) salts do not decompose on heating **except** Iron (II) sulphate (VI) and Copper (II) sulphate (VI).

(i)Iron (II) sulphate (VI) decomposes on strong heating to produce acidic sulphur (IV)oxide and sulphur(VI)oxide gases. Iron(III)oxide is formed as a brown /yellow residue.

Chemical equation

2FeSO4 (s) -> Fe2O3(s) + SO2(g) + SO3(g)

This reaction is used for the school laboratory preparation of small amount of sulphur(VI)oxide gas.

Sulphur (VI) oxide readily /easily solidifies as white silky needles when the mixture is passed through freezing mixture/ice cold water.

Sulphur (IV) oxide does not.

(ii) Copper(II)sulphate(VI) decomposes on strong heating to black copper (II) oxide and Sulphur (IV) oxide gas.

Chemical equation

2CuSO4 (s) -> CuO(s) + SO2(g)

This reaction is used for the school laboratory preparation of small amount of sulphur(IV)oxide gas.

6. The following experiments show the test for the presence of sulphate (VI) (SO42-)and sulphate(IV) (SO32-) ions in a sample of a salt/compound:

Experiments/Observations:

**(a)Using Lead(II)nitrate(V)**

I. To about 5cm3 of a salt solution in a test tube add four drops of Lead(II)nitrate(V)solution. Preserve.

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt | SO42- , SO32- , CO32- , Cl- ions |

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt persists | SO42- , Cl- ions |

Observation 2

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt dissolves | SO32- , CO32- , ions |

III.(a)To the preserved sample observation 1 in (II) above, Heat to boil.

Observation 1

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt persists on boiling | SO42- ions |

Observation 2

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt dissolves on boiling | Cl - ions |

.(b)To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

|  |  |
| --- | --- |
| Observation | Inference |
| (i)acidified potassium manganate(VII)decolorized  (ii)Orange colour of acidified potassium  dichromate(VI) turns to green | SO32- ions |

Observation 2

|  |  |
| --- | --- |
| Observation | Inference |
| (i)acidified potassium manganate(VII) not decolorized  (ii)Orange colour of acidified potassium  dichromate(VI) does not turns to green | CO32- ions |

Experiments/Observations:

**(b)Using Barium(II)nitrate(V)/ Barium(II)chloride**

I. To about 5cm3 of a salt solution in a test tube add four drops of Barium(II) nitrate (V) / Barium(II)chloride. Preserve.

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt | SO42- , SO32- , CO32- ions |

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt persists | SO42- , ions |

Observation 2

|  |  |
| --- | --- |
| Observation | Inference |
| White precipitate/ppt dissolves | SO32- , CO32- , ions |

III.To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

|  |  |
| --- | --- |
| Observation | Inference |
| (i)acidified potassium manganate(VII)decolorized  (ii)Orange colour of acidified potassium  dichromate(VI) turns to green | SO32- ions |

Observation 2

|  |  |
| --- | --- |
| Observation | Inference |
| (i)acidified potassium manganate(VII) not decolorized  (ii)Orange colour of acidified potassium  dichromate(VI) does not turns to green | CO32- ions |

Explanations

**Using Lead(II)nitrate(V)**

(i)Lead(II)nitrate(V) solution reacts with chlorides(Cl-), Sulphate (VI) salts (SO42- ), Sulphate (IV)salts (SO32-) and carbonates(CO32-) to form the insoluble white precipitate of Lead(II)chloride, Lead(II)sulphate(VI), Lead(II) sulphate (IV) and Lead(II)carbonate(IV).

Chemical/ionic equation:

Pb2+(aq) + Cl- (aq) -> PbCl2(s)

Pb2+(aq) + SO42+ (aq) -> PbSO4 (s)

Pb2+(aq) + SO32+ (aq) -> PbSO3 (s)

Pb2+(aq) + CO32+ (aq) -> PbCO3 (s)

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Lead(II)chloride and Lead(II)sulphate(VI) do not react with the acid and thus their white precipitates remain/ persists.

- Lead(II) sulphate (IV) and Lead(II)carbonate(IV) reacts with the acid to form **soluble** Lead(II) nitrate (V) and produce/effervesces/fizzes/bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

PbSO3 (s) + 2H+(aq) -> H2 O (l) + Pb2+(aq) + SO2 (g)

PbCO3 (s) + 2H+(aq) -> H2 O (l) + Pb2+(aq) + CO2 (g)

(iii)When Lead(II)chloride and Lead(II)sulphate(VI) are heated/warmed;

- Lead(II)chloride dissolves in hot water/on boiling(recrystallizes on cooling)

- Lead(II)sulphate(VI) do not dissolve in hot water thus its white precipitate persists/remains on heating/boiling.

(iv)When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green. **Carbon(IV)oxide will not.**

Chemical equation:

5SO32-(aq)+ 2MnO4- (aq) +6H+(aq) -> 5SO42-(aq) + 2Mn2+(aq) + 3H2O(l)

(purple) (colourless)

3SO32-(aq)+ Cr2O72-(aq) +8H+(aq) -> 3SO42-(aq) + 2Cr3+(aq) + 4H2O(l)

(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

Ca(OH)2(aq)+ CO2 (g) -> CaCO3(s) + H2O(l)

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

**Using Barium(II)nitrate(V)/ Barium(II)Chloride**

(i)Barium(II)nitrate(V) and/ or Barium(II)chlo**ride** solution reacts with Sulphate (VI) salts (SO42- ), Sulphate (IV)salts (SO32-) and carbonates(CO32-) to form the insoluble white precipitate of Barium(II)sulphate(VI), Barium(II) sulphate (IV) and Barium(II)carbonate(IV).

Chemical/ionic equation:

Ba2+(aq) + SO42+ (aq) -> BaSO4 (s)

Ba2+(aq) + SO32+ (aq) -> BaSO3 (s)

Ba2+(aq) + CO32+ (aq) -> BaCO3 (s)

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Barium (II)sulphate(VI) do not react with the acid and thus its white precipitates remain/ persists.

- Barium(II) sulphate (IV) and Barium(II)carbonate(IV) reacts with the acid to form **soluble** Barium(II) nitrate (V) and produce /effervesces /fizzes/ bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

BaSO3 (s) + 2H+(aq) -> H2 O (l) + Ba2+(aq) + SO2 (g)

BaCO3 (s) + 2H+(aq) -> H2 O (l) + Ba2+(aq) + CO2 (g)

(iii) When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green. **Carbon(IV)oxide will not.**

Chemical equation:

5SO32-(aq)+ 2MnO4- (aq) +6H+(aq) -> 5SO42-(aq) + 2Mn2+(aq) + 3H2O(l)

(purple) (colourless)

3SO32-(aq)+ Cr2O72-(aq) +8H+(aq) -> 3SO42-(aq) + 2Cr3+(aq) + 4H2O(l)

(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

Ca(OH)2(aq)+ CO2 (g) -> CaCO3(s) + H2O(l)

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

Summary test for Sulphate (VI) (SO42-)and Sulphate(IV) (SO32-) salts

Heat to boil

White ppt persist

on heating in SO42-

in CO32-

White ppt dissolves on heating in Cl-

Acidified KMnO4 decolorized in SO32-

White ppt with lime water in CO32-

Acidified KMnO4

K2Cr2O7 / Lime water

white ppt persist /remains in SO32- and CO32-

white ppt dissolves in SO32- and CO32-

Dilute nitric(V) acid

Lead(II)nitrate(V)

White precipitates of Cl-, SO42- , SO32- and CO32-

Unknown salt

Practice

1. Study the flow chart below and use it to answer the questions that follow

Sodium salt solution

Acidified K2Cr2O7

Gas G and colour of solution

changes orange to green

Colourless solution B

Barium nitrate(VI) (VI)(aq)

White precipitate

Dilute HCl

(a)**Identify the:**

**I: Sodium salt solution** Sodium sulphate(IV)/Na2SO3

**II: White precipitate** Barium sulphate(IV)/BaSO3 **III: Gas G** Sulphur (IV)Oxide /SO2 **IV: Colourless solution H** Barium chloride /BaCl2

(b)**Write an ionic equation for the formation of:**  **I.White precipitate**

Ionic equation Ba2+(aq) + SO32-(aq) -> BaSO3(s)

**II.Gas G**

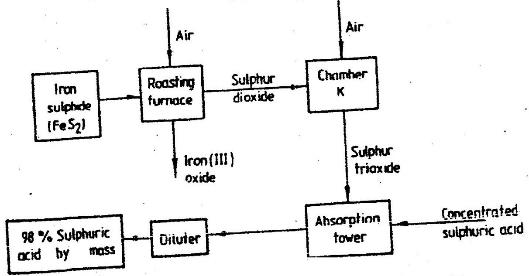
Ionic equation BaSO3(s)+ 2H+(aq) -> SO2 (g) + H2O (l) + Ba2+(aq)

**III. Green solution from the orange solution**

3SO32-(aq)+ Cr2O72-(aq) +8H+(aq) -> 3SO42-(aq) + 2Cr3+(aq) + 4H2O(l)

(Orange) (green)

2. Study the flow chart below and answer the questions that follow.



(i)Write equation for the reaction taking place at:

I.The roasting furnace (1mk)

II.The absorption tower (1mk)

III.The diluter (1mk)

1. The reaction taking place in chamber K is

SO2 (g) + 1/2O2 (g) SO3 (g)

I.Explain why it is necessary to use excess air in chamber K

II.Name another substance used in chamber K

**INTRODUCTION TO ELECTROLYSIS**

**(ELECTROLYTIC CELL)**

**1**.Electrolysis is defined simply as the **decomposition** of a **compound** by an electric current/**electricity**.

A compound that is decomposed by an electric current is called an electrolyte. Some electrolytes are **weak** while others are **strong**.

**2**.Strong electrolytes are those that are fully ionized/dissociated into (many) ions.

Common strong electrolytes include:

(i)all **mineral** acids

(ii)all strong **alkalis**/sodium hydroxide/potassium hydroxide.

(iii)all soluble **salts**

**3**.Weak electrolytes are those that are partially/partly ionized/dissociated into (few) ions.

Common weak electrolytes include:

(i)all **organic** acids

(ii)all **bases** except sodium hydroxide/potassium hydroxide.

(iii)**Water**

1. A compound that is **not** decomposed by an electric current is called nonelectrolyte.

Non-electrolytes are those compounds /substances that exist as molecules and thus cannot ionize/dissociate into(any) ions .

Common non-electrolytes include: (i) most organic solvents (e.g.

petrol/paraffin/benzene/methylbenzene/ethanol) (ii)all hydrocarbons(alkanes /alkenes/alkynes)

(iii)Chemicals of life(e.g. proteins, carbohydrates, lipids, starch, sugar)

1. An electrolytes in **solid** state have **fused** /joined ions and therefore do **not** conduct electricity but the **ions** (cations and anions) are **free** and **mobile** in **molten** and **aqueous** (solution, dissolved in water) state.

**6**.During electrolysis, the free ions are attracted to the **electrodes**.

An electrode is a rod through which current enter and leave the electrolyte during electrolysis.

An electrode that does not influence/alter the products of electrolysis is called an **inert electrode.**

Common inert electrodes include:

(i)**Platinum**

(ii)**Carbon graphite**

Platinum is not usually used in a school laboratory because it is very **expensive**. Carbon graphite is **easily**/readily and **cheaply** available (from used dry cells).

**7**.The **positive** electrode is called **Anode**.The anode is the electrode through which **current enter** the electrolyte/**electrons leave** the electrolyte

**8**.The **negative** electrode is called **Cathode**. The cathode is the electrode through which **current leave** the electrolyte / **electrons enter** the electrolyte

1. During the electrolysis, free **anions** are attracted to the **anode** where they **lose** /**donate** electrons to form **neutral** atoms/molecules. i.e.

M(l) -> M+(l) + e (for cations from molten electrolytes)

M(s) -> M+(aq) + e (for cations from electrolytes in aqueous state / solution / dissolved in water)

The neutral atoms /molecules form the **products** of electrolysis at the anode. This is called **discharge** at anode

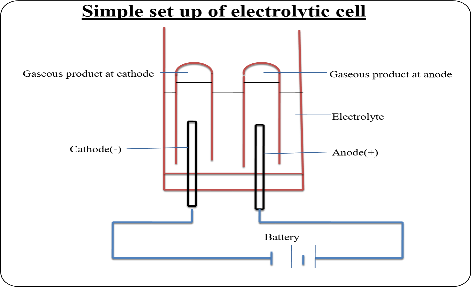
1. During electrolysis, free **cations** are attracted to the **cathode** where they **gain** /**accept/acquire** electrons to form **neutral** atoms/molecules.

X+ (aq) + 2e -> X(s) (for cations from electrolytes in aqueous state / solution / dissolved in water)

2X+ (l) + 2e -> X (l) (for cations from molten electrolytes)

The neutral atoms /molecules form the **products** of electrolysis at the cathode. This is called **discharge** at cathode.

1. The below set up shows an electrolytic cell.



1. For a compound /salt containing only two ion/binary salt the products of electrolysis in an electrolytic cell can be determined as in the below examples:

**a)To determine the products of electrolysis of molten Lead(II)chloride**

(i)Decomposition of electrolyte into free ions;

PbCl2 (l) -> Pb 2+(l) + 2Cl-(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

Pb 2+(l) + 2e -> Pb (l)

(Cation / Pb 2+ gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

2Cl-(l) -> Cl2 (g) + 2e

(Anion / Cl- donate/lose electrons to form free **atom** then agas **molecule)**

(iv)Products of electrolysis therefore are; I.At the cathode grey beads /solid lead metal.

II.At the anode pale green chlorine gas.

**b)To determine the products of electrolysis of molten Zinc bromide** (i)Decomposition of electrolyte into free ions;

ZnBr2 (l) -> Zn 2+(l) + 2Br-(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

Zn 2+(l) + 2e -> Zn(l)

(Cation / Zn2+ gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

2Br-(l) -> Br2 (g) + 2e

(Anion / Br- donate/lose electrons to form free **atom** then aliquid **molecule** whichchange to **gas** on heating)

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid Zinc metal.

II.At the anode **red** bromine **liquid** / **red/brown** bromine **gas.**

**c)To determine the products of electrolysis of molten sodium chloride**

(i)Decomposition of electrolyte into free ions;

NaCl (l) -> Na +(l) + Cl-(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

2Na+(l) + 2e -> Na (l)

(Cation / Na+ gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

2Cl-(l) -> Cl2 (g) + 2e

(Anion / Cl- donate/lose electrons to form free **atom** then agas **molecule)**

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid sodium metal.

II.At the anode pale green chlorine gas.

**d)To determine the products of electrolysis of molten Aluminium (III)oxide**

(i)Decomposition of electrolyte into free ions;

Al2O3 (l) -> 2Al 3+(l) + 3O2-(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

4Al 3+ (l) + 12e -> 4Al (l)

(Cation / Al 3+ gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

6O2-(l) -> 3O2 (g) + 12e

(Anion /6O2- donate/lose 12 electrons to form free **atom** then threegas **molecule)**

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid aluminium metal.

II.At the anode colourless gas that relights/rekindles glowing splint.

**13.**In industries electrolysis has the following uses/applications:

**(a)Extraction of reactive metals from their ores.**

Potassium, sodium ,magnesium, and aluminium are extracted from their ores using electrolytic methods.

**(b)Purifying copper after exraction from copper pyrites ores.**

Copper obtained from copper pyrites ores is not pure. After extraction, the copper is refined by electrolysing copper(II)sulphate(VI) solution using the **impure** copper as **anode** and a thin strip of **pure** copper as **cathode**. Electrode ionization take place there:

(i)At the cathode; **Cu2+ (aq) + 2e -> Cu(s)** (Pure copper deposits on the strip (ii)At the anode; **Cu(s) ->Cu2+ (aq) + 2e** (impure copper erodes/dissolves)

**(c)Electroplating**

The label EPNS(**E**lectro **P**lated **N**ickel **S**ilver) on some steel/metallic utensils mean they are plated/coated with silver and/or Nickel to **improve** their **appearance**(**add** their **aesthetic** value)and **prevent**/slow **corrosion**(**rusting** of iron). Electroplating is the process of coating a metal with another metal using an electric current.

During electroplating, the **cathode** is made of the metal to be **coated**/impure.

**Example:**

During the electroplating of a spoon with silver

(i)the spoon/impure is placed as the cathode(negative terminal of battery) (ii)the pure silver is placed as the anode(positive terminal of battery) (iii)the pure silver erodes/ionizes/dissociates to release electrons:

**Ag(s) ->Ag+ (aq) + e** (impure silver erodes/dissolves)

(iv) silver (**Ag+**)ions from electrolyte gain electrons to form pure silver deposits / coat /cover the spoon/impure

**Ag+ (aq) + e ->Ag(s)**  (pure silver deposits /coat/cover on spoon)



**GAS LAWS**

1. Matter is made up of small particle in accordance to Kinetic Theory of matter: Naturally, there are basically **three** states of matter: **Solid, Liquid** and **gas:** (i)A solid is made up of particles which are very closely packed with a definite/fixed shape and fixed/definite volume /occupies definite space. It has a very high density.

1. A liquid is made up of particles which have some degree of freedom. It thus has no definite/fixed shape. It takes the shape of the container it is put. A liquid has fixed/definite volume/occupies definite space.

1. A gas is made up of particles free from each other. It thus has no definite /fixed shape. It takes the shape of the container it is put. It has no fixed/definite volume/occupies every space in a container.

2.Gases are affected by **physical conditions**. There are **two** physical conditions:

(i)**Temperature**

(ii)**Pressure**

3. The SI unit of temperature is **Kelvin**(**K**).

**Degrees** Celsius/Centigrade(**oC**) are also used.

The two units can be interconverted from the relationship:

**oC + 273= K**

**K -273 = oC**

Practice examples

1. Convert the following into Kelvin. (i) O **oC**

**oC + 273 = K substituting :** O oC + 273 = **273 K**

1. -273 **oC**

**oC + 273 = K substituting :** -273oC + 273 = **0 K**

1. 25 **oC**

**oC + 273 = K substituting :** 25 oC + 273 = **298 K**

1. 100 **oC**

**oC + 273 = K substituting :** 100 oC + 273 = **373 K**

2. Convert the following into degrees Celsius/Centigrade(oC).

(i) 10 K

**K -273 = oC substituting:** 10 – 273 = **-263 oC** (ii) (i) 1 K

**K -273 = oC substituting:** 1 – 273 =  **-272 oC**

1. 110 K
   1. **-273 = oC substituting: 110 – 273 = -163 oC**

1. -24 K
   1. **-273 = oC substituting: -24 – 273 = -297 oC**

The **standard** temperature is **273K** = **0 oC**.

The **room** temperature is assumed to be **298K** = **25oC**

1. The SI unit of pressure is Pascal(**Pa**) / Newton per metre squared (**Nm-2**) . Millimeters’ of mercury(**mmHg**) ,centimeters of mercury(**cmHg**) and **atmospheres** are also commonly used.

The units are **not** interconvertible but Pascals(Pa) are equal to Newton per metre squared(Nm-2).

The **standard** pressure is the **atmospheric** pressure.

Atmospheric pressure is **equal** to about:

* 1. 101325 Pa
  2. 101325 Nm-2
  3. 760 mmHg (iv)76 cmHg

(v)one atmosphere.

1. Molecules of gases are always in continuous random motion at high speed. This motion is affected by the physical conditions of temperature and pressure. Physical conditions change the volume occupied by gases in a **closed** system. The effect of physical conditions of temperature and pressure was investigated and expressed in both Boyles and Charles laws.

1. Boyles law states that

“**the volume of a fixed mass of a gas is inversely proportional to the pressure at constant/fixed temperature** ” Mathematically:

Volume α 1 **(Fixed /constant Temperature)**

Pressure

V α 1 **(Fixed /constant T)** ie **PV = Constant**(k)P

From Boyles law , an **increase** in pressure of a gas cause a **decrease** in volume. i.e **doubling** the pressure cause the volume to be **halved**.

Graphically therefore a plot of volume(**V**) **against** pressure (**P**) produces a **curve**.

V

P

Graphically a plot of volume(**V**) **against** inverse/reciprocal of pressure (**1/p**) produces a **straight line**

**V**

**1**

**/**

**P**

For **two** gases then **P1 V1 = P2 V2**

P1 = Pressure of gas 1

V1 = Volume of gas 1

P2 = Pressure of gas 2

V2 = Volume of gas 2

Practice examples:

1. **A fixed mass of gas at 102300Pa pressure has a volume of 25cm3.Calculate its volume if the pressure is doubled.**

Working

P1 V1 = P2 V2 Substituting :102300 x 25 = (102300 x **2**) x V2

V2 = 102300 x 25 = **12.5cm3**

(102300 x 2)

1. **Calculate the pressure which must be applied to a fixed mass of 100cm3 of Oxygen for its volume to triple at 100000Nm-2.**

P1 V1 = P2 V2 Substituting :100000 x 100 = P2 x (100 x **3**)

V2 = 100000 x 100 = **33333.3333 Nm-2**

(100 x 3)

**3.A 60cm3 weather ballon full of Hydrogen at atmospheric pressure of 101325Pa was released into the atmosphere. Will the ballon reach stratosphere where the pressure is 90000Pa?**

P1 V1 = P2 V2 Substituting :101325 x 60 = 90000 x V2

V2 = 101325 x 60 = **67.55 cm3**

90000

The new volume at 67.55 cm3 **exceed** ballon capacity of 60.00 cm3.It will **burst** before reaching destination.

7.Charles law states that“**the volume of a fixed mass of a gas is directly proportional to the absolute temperature at constant/fixed pressure** ” Mathematically:

Volume α Pressure **(Fixed /constant pressure)**

V α T  **(Fixed /constant P)** ie **V = Constant**(k)

T

From Charles law , an **increase** in temperature of a gas cause an **increase** in volume. i.e **doubling** the temperature cause the volume to be **doubled**.

Gases expand/increase by **1/273** by volume on heating.Gases contact/decrease by **1/273** by volume on cooling at constant/fixed pressure.

The volume of a gas continue decreasing with decrease in temperature until at **-273oC /0 K** the volume is **zero**. i.e. there is no gas.

This temperature is called **absolute zero.** It is the **lowest** temperature at which a gas **can** exist.

Graphically therefore a plot of volume(**V**) **against** Temperature(**T**) in:

(**i**)oC produces a **straight line** that is **extrapolated** to the absolute zero of -273**oC** .

V

**-**

**273**

**o**

**C**

**0**

**o**

**C**

T

(

**o**

**C**

)

(**ii**)Kelvin/K produces a **straight line** from absolute zero of **O Kelvin**

**V**

**0**

**T**

**(**

**K**

**elvin**

**)**

For **two** gases then **V1  = V2**

**T1  T2**

T1 = Temperature **in Kelvin** of gas 1

V1 = Volume of gas 1

T2 = Temperature **in Kelvin** of gas 2

V2 = Volume of gas 2

Practice examples:

1. **500cm3 of carbon(IV)oxide at 0oC was transfered into a cylinder at -4oC. If the capacity of the cylinder is 450 cm3,explain what happened.**

V1  = V2 substituting 500 = V2

T1  T2 (0 +273) (-4 +273)

= 500 x (-4 x 273) = **492.674cm3** (0 + 273)

The capacity of cylinder (500cm3) is **less** than new volume(492.674cm3).

**7.326cm3**(500-492.674cm3)of carbon(IV)oxide gas did not fit into the cylinder.

1. **A mechanic was filling a deflated tyre with air in his closed garage using a hand pump. The capacity of the tyre was 40,000cm3 at room temperature. He rolled the tyre into the car outside. The temperature outside was 30oC.Explain what happens.**

V1  = V2 substituting 40000 = V2

T1  T2 (**25** +273) (**30** +273)

= 40000 x (30 x 273) = **40671.1409cm3** (25 + 273)

The capacity of a tyre (40000cm3) is **less** than new volume(40671.1409cm3). The tyre thus bursts.

1. **A hydrogen gas balloon with 80cm3 was released from a research station at room temperature. If the temperature of the highest point it rose is -30oC , explain what happened.**

V1  = V2 substituting 80 = V2

T1  T2 (**25** +273) (-**30** +273)

= 80 x (-30 x 273) = **65.2349cm3** (25 + 273)

The capacity of balloon (80cm3) is **more** than new volume (65.2349cm3). The balloon thus remained intact.

8. The continuous random motion of gases differ from gas to the other.The movement of molecules (of a gas) from region of high concentration to a region of low concentration is called **diffusion.**

The rate of diffusion of a gas depends on its density. i.e. **The higher the rate of diffusion, the less dense the gas**.

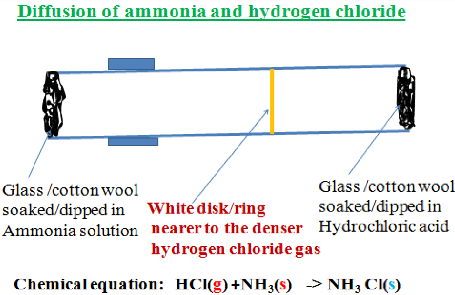
The density of a gas depends on its molar mass/relative molecular mass. i.e. **The** higher **the density the** higher **the molar mass/relative atomic mass and thus the** lower **the rate of diffusion.**

Examples

1.Carbon (IV)oxide(CO2) has a molar mass of 44g.Nitrogen(N2)has a molar mass of 28g. (N2)is thus lighter/less dense than Carbon (IV)oxide(CO2). N2 diffuses faster than CO2.

2.Ammonia(NH3) has a molar mass of 17g.Nitrogen(N2)has a molar mass of 28g. (N2)is thus about **twice** lighter/less dense than Ammonia(NH3). Ammonia(NH3)diffuses twice faster than N2.

3. Ammonia(NH3) has a molar mass of 17g.Hydrogen chloride gas has a molar mass of 36.5g.Both gases on contact react to form **white fumes** of ammonium chloride .When a glass/cotton wool dipped in ammonia and another glass/cotton wool dipped in hydrochloric acid are placed at opposite ends of a glass tube, both gases diffuse towards each other. A white disk appears near to glass/cotton wool dipped in hydrochloric acid. This is because hydrogen chloride is heavier/denser than Ammonia and thus its rate of diffusion is lower .



The rate of diffusion of a gas is in accordance to **Grahams law of diffusion**.

Grahams law states that:

**“the rate of diffusion of a gas is inversely proportional to the square root of its**

**density, at the same/constant/fixed temperature and pressure”** Mathematically

**R α 1**  and since density is proportional to mass then **R α 1**

**√ p √ m**

For two gases then:

**R1 = R2** where: R**1** and R**2** is the **rate** of diffusion of **1**st and **2**nd gas.

**√M2 √M1** M**1** and M**2** is the **molar mass** of **1**st and **2**nd gas.

Since rate is inverse of time. i.e. the higher the rate the less the time:

For two gases then:

T**1** = T2 where: T**1** and T**2** is the **time taken** for **1**st and **2**nd gas to diffuse.

**√**M1 **√**M2 M**1** and M**2** is the **molar mass** of **1**st and **2**nd gas.

**Practice examples**:

1. **It takes 30 seconds for 100cm3 of carbon(IV)oxide to diffuse across a porous plate. How long will it take 150cm3 of nitrogen(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure.**

**(C=12.0,N=14.0=16.0)**

Molar mass CO2=44.0 Molar mass NO2=46.0 Method 1

100cm3 CO2  takes 30seconds

150cm3 takes 150 x30 = 45seconds

100

* 1. CO2 = √ molar mass CO2 => 45seconds = √ 44.0

T NO2 √ molar mass NO2 T NO2 √ 46.0

T NO = **46.0114** seconds

Method 2

100cm3 CO2  takes 

1cm3 takes 100 x1 = **3.3333cm3sec**

30

R CO2 = √ molar mass NO2 => 3.3333cm3sec-1 = √ 46.0 R NO2 √ molar mass CO2 R NO2 √ 44.0

R NO2 = 3.3333cm3sec-1 x √ 44.0 = **3.2601**cm3sec-1

√ 46.0

3.2601cm3 takes 1seconds

150cm3 take 150cm3 = **46.0109seconds**

3.2601cm3

1. **How long would 200cm3 of Hydrogen chloride take to diffuse through a porous plug if carbon(IV)oxide takes 200seconds to diffuse through.**

Molar mass CO2 = 44g Molar mass HCl = 36.5g

* 1. CO2 = √ molar mass CO2 => 200 seconds = √ 44.0 T HCl √ molar mass HClT HCl√ 36.5

T HCl  = **182.1588** seconds

1. **Oxygen gas takes 250 seconds to diffuse through a porous diaphragm. Calculate the molar mass of gas Z which takes 227 second to diffuse.**

Molar mass O2 = 32g Molar mass Z = x g

|  |  |
| --- | --- |
| T O2 = √ molar mass O2 | => 250 seconds = √ 32.0 |
| T Z √ molar mass Z | 227seconds√ x |

 = **26.3828** grams

250

1. **25cm3 of carbon(II)oxide diffuses across a porous plate in 25seconds. How long will it take 75cm3 of Carbon(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,0=16.0)**

Molar mass CO2 = 44.0 Molar mass CO = 28.0 Method 1

25cm3 CO takes 25seconds

75cm3 takes 75 x25 = 75seconds

25

* 1. CO2 = √ molar mass CO2 => T CO2seconds = √ 44.0 T CO √ molar mass CO75√ 28.0

T CO = **94.0175** seconds

Method 2

25cm3 CO2  takes 25seconds

1cm3 takes 25 x1 = **1.0cm3sec-1**

25

R CO2 = √ molar mass CO => x cm3sec-1 = √ 28.0

R CO √ molar mass CO2 1.0cm3sec-1 √ 44.0

R CO = **0.7977**cm3sec-1

√ 44.0

0.7977cm3 takes 1 seconds

75cm3 takes 75cm3 = **94.0203seconds**

0.7977cm3

**MOLES**

**(b)Introduction to the mole, molar masses and Relative atomic masses**

1. The mole is the **SI** unit of the **amount** of substance.
2. The number of particles e.g. atoms, ions, molecules, electrons, cows, cars are all measured in terms of moles.
3. The number of particles in one mole is called the **Avogadros Constant**. It is denoted “**L**”.

The Avogadros Constant contain **6.023 x10 23** particles. i.e.

1mole = 6.023 x10 23 particles = 6.023 x10 **23**

2 moles = 2 x 6.023 x10 23 particles = 1.205 x10 **24**

0.2 moles = 0.2 x 6.023 x10 23 particles = 1.205 x10 **22**

0.0065 moles = 0.0065 x 6.023 x10 23 particles = 3.914 x10 **21**

3. The mass of one mole of a substance is called **molar mass**. The molar mass of: (i)an **element** has mass equal to relative **atomic** mass /RAM(in grams)of the element e.g.

Molar mass of carbon(C)= relative atomic mass = 12.0g

6.023 x10 23 particles of carbon = 1 mole =12.0 g

Molar mass of sodium(Na) = relative atomic mass = 23.0g

6.023 x10 23 particles of sodium = 1 mole =23.0 g

Molar mass of Iron (Fe) = relative atomic mass = 56.0g

6.023 x10 23 particles of iron = 1 mole =56.0 g

(ii)a **molecule** has mass equal to relative molecular mass /RMM (in grams)of the molecule. Relative molecular mass is the **sum** of the relative atomic masses of the elements making the molecule.

The number of atoms making a molecule is called **atomicity**. Most **gaseous** molecules are **diatomic (e.g. O2**, **H2**, **N2**, **F2**, **Cl2**, **Br2**, **I2**)noble gases are **monoatomic(e.g. He, Ar, Ne, Xe)**,Ozone gas(**O3**) is **triatomic** e.g.

Molar mass **Oxygen molecule(O2)** =relative molecular mass =(16.0x 2)g =32.0g

6.023 x10 23 particles of Oxygen molecule = 1 mole = 32.0 g

Molar mass **chlorine molecule(Cl2)** =relative molecular mass =(35.5x 2)g =71.0g

6.023 x10 23 particles of chlorine molecule = 1 mole = 71.0 g

Molar mass **Nitrogen molecule(N2)** =relative molecular mass =(14.0x 2)g =28.0g

6.023 x10 23 particles of Nitrogen molecule = 1 mole = 28.0 g

(ii)a **compound** has mass equal to relative formular mass /RFM (in grams)of the molecule. Relative formular mass is the **sum** of the relative atomic masses of the elements making the compound. e.g.

(i)Molar mass **Water(H2O)** = relative formular mass =[(1.0 x 2 ) + 16.0]g =18.0g

6.023 x10 23 particles of Water molecule = 1 mole = 18.0 g

6.023 x10 23 particles of Water molecule has:

* **2** x 6.023 x10 23 particles of Hydrogen atoms

**-1** x 6.023 x10 23 particles of Oxygen atoms

(ii)Molar mass **sulphuric(VI)acid(H2SO4)** = relative formular mass =[(1.0 x 2 ) + 32.0 + (16.0 x 4)]g =98.0g

6.023 x10 23 particles of sulphuric(VI)acid(H2SO4) = 1 mole = 98.0g

6.023 x10 23 particles of sulphuric(VI)acid(H2SO4) has:

* **2** x 6.023 x10 23 particles of **H**ydrogen atoms

-**1** x 6.023 x10 23 particles of **S**ulphur atoms

**-4** x 6.023 x10 23 particles of **O**xygen atoms

(iii)Molar mass **sodium carbonate(IV)(Na2CO3)** = relative formular mass =[(23.0 x 2 ) + 12.0 + (16.0 x 3)]g =106.0g

6.023 x10 23 particles of sodium carbonate(IV)(Na2CO3) = 1 mole = 106.0g

6.023 x10 23 particles of sodium carbonate(IV)(Na2CO3) has:

* **2** x 6.023 x10 23 particles of **S**odium atoms

-**1** x 6.023 x10 23 particles of **C**arbon atoms

**-3** x 6.023 x10 23 particles of **O**xygen atoms

(iv)Molar mass **Calcium carbonate(IV)(CaCO3)** = relative formular mass =[(40.0+ 12.0 + (16.0 x 3)]g =100.0g.

6.023 x10 23 particles of Calcium carbonate(IV)(CaCO3) = 1 mole = 100.0g 6.023 x10 23 particles of Calcium carbonate(IV)(CaCO3) has:

* **1** x 6.023 x10 23 particles of **C**alcium atoms

-**1** x 6.023 x10 23 particles of **C**arbon atoms

**-3** x 6.023 x10 23 particles of **O**xygen atoms

(v)Molar mass **Water(H2O)** = relative formular mass =[(2 x 1.0 )+ 16.0 ]g =18.0g

6.023 x10 23 particles of Water(H2O) = 1 mole = 18.0g

6.023 x10 23 particles of Water(H2O) has:

* **2** x 6.023 x10 23 particles of Hydrogen atoms

**-2** x 6.023 x10 23 particles of **O**xygen atoms

**Practice**

1. Calculate the number of moles present in:

1. 0.23 g of Sodium atoms

Molar mass of Sodium atoms = 23g

Moles = mass in grams = > 0.23g = **0.01moles**

Molar mass 23

1. 0.23 g of Chlorine atoms

Molar mass of Chlorine atoms = 35.5 g

Moles = mass in grams = > 0.23g = **0.0065moles /6.5 x 10-3 moles** Molar mass 35.5

1. 0.23 g of Chlorine molecules

Molar mass of Chlorine molecules =( 35.5 x 2) = **71.0** g

Moles = mass in grams = > 0.23g = **0.0032moles /3.2 x 10-3 moles** Molar mass 71

1. 0.23 g of dilute sulphuric(VI)acid

Molar mass of H2SO4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

Moles = mass in grams = > 0.23g = **0.0023moles /2.3 x 10-3 moles** Molar mass 98

1. Calculate the number of atoms present in:(Avogadros constant L = 6.0 x 10 23)

(i) 0.23 g of dilute sulphuric (VI)acid

**Method I**

Molar mass of H2SO4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

Moles = mass in grams = > 0.23g = **0.0023moles /2.3 x 10-3 moles** Molar mass 98

1 mole has 6.0 x 10 23 atoms

2.3 x 10-3 moles has (2.3 x 10-3 x 6.0 x 10 23) = **1.38 x 10 21**atoms

1

**Method II**

Molar mass of H2SO4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

**98.0**g = 1 mole has 6.0 x 10 23 atoms

0.23 g therefore has (0.23 g x 6.0 x 10 23 )= **1.38 x 10 21**atoms

98

1. 0.23 g of sodium carbonate(IV)decahydrate

Molar mass of Na2CO3.10H2 O=

[(2 x 23) + 12 + (3 x16) + (10 x 1.0) + (10 x 16)] = **276.0**g **Method I**

Moles = mass in grams = > 0.23g = **0.00083moles /**

Molar mass 276 **8.3 x 10-4 moles**

1 mole has 6.0 x 10 23 atoms

8.3 x 10-4 moles has (8.3 x 10-4 moles x 6.0 x 10 23) = **4.98 x 10 20**atoms

1

**Method II**

276.0g = 1 mole has 6.0 x 10 23 atoms

0.23 g therefore has (0.23 g x 6.0 x 10 23 )= **4.98 x 10 20**atoms

276.0

1. 0.23 g of Oxygen gas

Molar mass of O2 = (2 x16) = **32.0** g

**Method I**

Moles = mass in grams = > 0.23g = **0.00718moles /**

Molar mass 32 **7.18 x 10-3 moles**

1 mole has **2** x 6.0 x 10 23 **atoms** in O2

7.18 x 10-3moles has (7.18 x 10-3moles x **2** x 6.0 x 10 23) =**8.616 x 10 21**atoms

1

**Method II**

32.0g = 1 mole has **2** x 6.0 x 10 23 **atoms** in O2

0.23 g therefore has (0.23 g x **2** x 6.0 x 10 23 )= **8.616 x 10 21**atoms

32.0

1. 0.23 g of Carbon(IV)oxide gas

Molar mass of CO2 = [12 + (2 x16)] = **44.0** g

**Method I**

Moles = mass in grams = > 0.23g = **0.00522moles /**

Molar mass 44 **5.22 x 10-3 moles**

1 mole has **3** x 6.0 x 10 23 **atoms** in CO2

7.18 x 10-3moles has (5.22 x 10-3moles x **3** x 6.0 x 10 23) =**9.396 x 10 21**atoms

1

**Method II**

44.0g = 1 mole has **3** x 6.0 x 10 23 **atoms** in CO2

0.23 g therefore has (0.23 g x 3 x 6.0 x 10 23 )= **9.409 x 10 21**atoms

44.0

**(c)Empirical and molecular formula**

1.The empirical formula of a compound is its simplest formula. It is the simplest whole number ratios in which atoms of elements combine to form the compound. 2.It is mathematically the lowest common multiple (LCM) of the atoms of the elements in the compound

3.Practically the empirical formula of a compound can be determined as in the following examples.

**To determine the empirical formula of copper oxide**

**(a)**Method 1:From copper to copper(II)oxide **Procedure.**

Weigh a clean dry covered crucible(**M1**).Put two spatula full of copper **powder** into the crucible. Weigh again (**M2**).Heat the crucible on a strong Bunsen flame for five minutes. Lift the lid, and swirl the crucible carefully using a pair of tong. Cover the crucible and continue heating for another five minutes. Remove the lid and stop heating. Allow the crucible to cool. When cool replace the lid and weigh the contents again (**M3**).

**Sample results**

|  |  |
| --- | --- |
| Mass of crucible(**M1**) | 15.6g |
| Mass of crucible + copper before heating(**M2**) | 18.4 |
| Mass of crucible + copper after heating(**M3**) | 19.1 |

**Sample questions**

1. **Calculate the mass of copper powder used.**

Mass of crucible + copper before heating(**M2**) = 18.4

Less Mass of crucible(**M1**) = - 15.6g

Mass of copper **2.8 g**

1. **Calculate the mass of Oxygen used to react with copper.**

Method I

Mass of crucible + copper after heating(**M3**) = 19.1g

Mass of crucible + copper before heating(**M2**) = - 18.4g

Mass of Oxygen =  **0.7 g**

Method II

Mass of crucible + copper after heating(**M3**) = 19.1g

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  |  | Mass of crucible |  |  | = - 15.6g |
|  |  | Mass of copper(II)Oxide |  |  | = 3.5 g |
|  |  | Mass of copper(II)Oxide |  |  | = 3.5 g |
|  |  | Mass of copper |  |  | = - 2.8 g |
|  |  | Mass of Oxygen |  |  | =  **0.7 g** |

1. **Calculate the number of moles of: (i) copper used** (Cu = 63.5)

number of moles of copper = mass used => 2.8 = **0.0441moles**

Molar mass 63.5

**(ii) Oxygen used** (O = 16.0)

number of moles of oxygen = mass used => 0.7 = **0.0441moles**

Molar mass 16.0

1. **Determine the mole ratio of the reactants**

Moles of copper = 0.0441moles = 1 => Mole ratio Cu: O = **1:1** Moles of oxygen 0.0441moles 1

**5.What is the empirical, formula of copper oxide formed.**

CuO (copper(II)oxide

1. **State and explain the observations made during the experiment.**

Observation

Colour change from **brown** to **black**

Explanation

Copper powder is **brown**. On heating it reacts with oxygen from the air to form **black** copper(II)oxide

1. **Explain why magnesium ribbon/shavings would be unsuitable in a similar experiment as the one above.**

Hot magnesium generates enough heat energy to react with **both** Oxygen and Nitrogen in the air forming a white solid mixture of Magnesium **oxide** and magnesium **nitride**. This causes experimental mass errors.

**(b)**Method 2:From copper(II)oxide to copper **Procedure.**

Weigh a clean dry porcelain boat (**M1**). Put two spatula full of copper(II)oxide **powder** into the crucible. Reweigh the porcelain boat (**M2**).Put the porcelain boat in a glass tube and set up the apparatus as below;

**Determining empirical formula from** copper(II)oxide **to copper**

jgthungu@gmail.com

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HEAT

Copper(II)oxide

Hydrogen /Laboratory gas /

Ammonia gas/Carbon(II)Oxide

gas from a generator

Excess hydrogen

burning

Pass slowly(to prevent copper(II)oxide from being blown away)a stream of either dry Hydrogen /ammonia/laboratory gas/ carbon(II)oxide gas for about two minutes from a suitable generator.

When all the in the apparatus set up is driven out ,heat the copper(II)oxide strongly for about five minutes until there is no further change. Stop heating.

Continue passing the gases until the glass tube is cool.

Turn off the gas generator.

Carefully remove the porcelain boat form the combustion tube.

Reweigh (**M3**).

**Sample results**

|  |  |
| --- | --- |
| Mass of boat(**M1**) | 15.6g |
| Mass of boat before heating(**M2**) | 19.1 |
| Mass of boat after heating(**M3**) | 18.4 |

**Sample questions**

1. **Calculate the mass of copper(II)oxide used.**

Mass of boat before heating(**M2**) = 19.1

Mass of empty boat(**M1**) = - 15.6g

Mass of copper(II)Oxide **3.5 g**

1. **Calculate the mass of (i) Oxygen.**

Mass of boat before heating(**M2**) = 19.1

Mass of boat after heating (**M3**) = - 18.4g

Mass of oxygen = **0.7 g**

**(ii)Copper**

Mass of copper(II)Oxide = 3.5 g

Mass of oxygen = 0.7 g

Mass of oxygen = **2.8 g**

1. **Calculate the number of moles of: (i) Copper used (Cu = 63.5)**

number of moles of copper = mass used => 2.8 = **0.0441moles**

Molar mass 63.5

**(ii) Oxygen used (O = 16.0)**

number of moles of oxygen = mass used => 0.7 = **0.0441moles**

Molar mass 16.0

1. **Determine the mole ratio of the reactants**

Moles of copper = 0.0441moles = 1 => Mole ratio Cu: O = **1:1** Moles of oxygen 0.0441moles 1

**5.What is the empirical, formula of copper oxide formed.**

CuO (copper(II)oxide

1. **State and explain the observations made during the experiment.**

Observation

Colour change from **black** to **brown**

Explanation

Copper(II)oxide powder is **black**. On heating it is reduced by a suitable reducing agent to **brown** copper metal.

1. **Explain why magnesium oxide would be unsuitable in a similar experiment as the one above.**

Magnesium is high in the reactivity series. None of the above reducing agents is strong enough to reduce the oxide to the metal.

1. **Write the equation for the reaction that would take place when the reducing agent is:**

**(i) Hydrogen**

CuO(s) + H2(g) -> Cu(s) + H2O(l)

(Black) (brown) (colourless liquid form

on cooler parts )

**(ii)Carbon(II)oxide**

CuO(s) + CO (g) -> Cu(s) + CO2(g)

(Black) (brown) (colourless gas, form

white ppt with lime water )

**(iii)Ammonia**

3CuO(s) + 2NH3(g) -> 3Cu(s) + N2 (g) + 3H2O(l)

(Black) (brown) (colourless liquid form

on cooler parts )

1. **Explain why the following is necessary during the above experiment;**

* 1. **A stream of dry hydrogen gas should be passed before heating copper (II) Oxide.**

Air combine with hydrogen in presence of heat causing an explosion

* 1. **A stream of dry hydrogen gas should be passed after heating copper (II) Oxide has been stopped.**

Hot metallic copper can be re-oxidized back to copper(II)oxide

* 1. **A stream of excess carbon (II)oxide gas should be ignited to burn**

Carbon (II)oxide is highly poisonous/toxic. On ignition it burns to form less toxic carbon (IV)oxide gas.

1. **State two sources of error in this experiment.**

(i)All copper(II)oxide may not be reduced to copper.

(ii)Some copper(II)oxide may be blown out the boat by the reducing agent.

4.Theoreticaly the empirical formula of a compound can be determined as in the following examples.

1. **A oxide of copper contain 80% by mass of copper. Determine its empirical formula. (Cu = 63.5, 16.0)**

% of Oxygen = 100% - % of Copper => 100- 80 = **20%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Copper | Oxygen |
| Symbol | Cu | O |
| Moles present = % composition  Molar mass |  |  |
| Divide by the smallest value | 1.25 |  |
|  | 1.25 | 1.25 |
| Mole ratios | 1 | 1 |

Empirical formula is **CuO**

1. **1.60g of an oxide of Magnesium contain 0.84g by mass of Magnesium.**

**Determine its empirical formula(Mg = 24.0, 16.0)**

**Mass** of Oxygen = 1.60 – 0.84 => **0.56 g** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Magnesium | Oxygen |
| Symbol | Mg | O |
| Moles present = % composition  Molar mass |  | 0.56 |
| Divide by the smallest value | 0.35 |  |
|  | 0.35 | 0.35 |
| Mole ratios | 1 | 1 |

Empirical formula is **MgO**

**(c)An oxide of Silicon contain 47% by mass of Silicon. What is its empirical formula(Si = 28.0, 16.0)**

**Mass** of Oxygen = 100 – 47 => **53%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Silicon | Oxygen |
| Symbol | Si | O |
| Moles present = % composition  Molar mass |  |  |
| Divide by the smallest value | 1.68 |  |
|  | 1.68 | 1.68 |
| Mole ratios | 1 | 1.94 = 2 |

Empirical formula is **SiO2**

**(d)A compound contain 70% by mass of Iron and 30% Oxygen. What is its empirical formula(Fe = 56.0, 16.0)**

**Mass** of Oxygen = 100 – 47 => **53%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Silicon | Oxygen |
| Symbol | Si | O |
| Moles present = % composition  Molar mass |  |  |
| Divide by the smallest value | 1.68 |  |
|  | 1.68 | 1.68 |
| Mole ratios | 1 | 1.94 = 2 |

Empirical formula is **SiO2**

**2.During heating of a hydrated copper (II)sulphate(VI) crystals, the following readings were obtained: Mass of evaporating dish =300.0g**

**Mass of evaporating dish + hydrated salt = 305.0g**

**Mass of evaporating dish + anhydrous salt = 303.2g**

**Calculate the number of water of crystallization molecules in hydrated copper (II)sulphate(VI)**

**(Cu =64.5, S = 32.0,O=16.0, H = 1.0)**

**Working**

Mass of Hydrated salt = 305.0g -300.0g = **5.0g**

Mass of anhydrous salt = 303.2 g -300.0g = **3.2 g**

Mass of water in hydrated salt = 5.0g -3.2 g = **1.8g**

Molar mass of water(H2O) = **18.0g**

Molar mass of anhydrous copper (II)sulphate(VI) (CuSO4) = **160.5g**

|  |  |  |
| --- | --- | --- |
| Element/compound | anhydrous copper (II) sulphate(VI) | Oxygen |
| Symbol | CuSO4 | O |
| Moles present = composition by mass  Molar mass | 3,2 160.5 | 1.8 |
| Divide by the smallest value | 0.0199  0.0199 |  |
| Mole ratios | 1 | 5 |

The **empirical formula** of hydrated salt = CuSO4.**5**H2O

Hydrated salt has **five/5 molecules** of water of crystallizations

4. The molecular formula is the actual number of each kind of atoms present in a **molecule** of a compound.

The empirical formula of an ionic compound is the same as the chemical formula but for simple molecular structured compounds, the empirical formula may not be the same as the chemical formula.

The molecular formula is a multiple of empirical formula .It is determined from the relationship:

1. **n**  = Relative formular mass

Relative empirical formula where **n** is a whole number.

1. Relative empirical formula x **n** = Relative formular mass where **n** is a whole number.

**Practice sample examples**

1. **A hydrocarbon was found to contain 92.3% carbon and the remaining Hydrogen.**

**If the molecular mass of the compound is 78, determine the molecular formula(C=12.0, H =1.0)**

**Mass** of Hydrogen = 100 – 92.3 => **7.7%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Carbon | Hydrogen |
| Symbol | C | H |
| Moles present = % composition  Molar mass |  | 7.7 |
| Divide by the smallest value | 7.7 |  |
|  | 7.7 | 7.7 |
| Mole ratios | 1 | 1 |

Empirical formula is **CH**

The molecular formular is thus determined :

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | **n**  = Relative formular mass |  |  | = 78 = 6 |
|  | Relative empirical formula |  |  | 13 |

The molecular formula is (C H ) x 6 = **C6H6**

1. **A compound of carbon, hydrogen and oxygen contain 54.55% carbon, 9.09% and remaining 36.36% oxygen.**

**If its relative molecular mass is 88, determine its molecular formula(C=12.0, H =1.0, O= 16.0)**

|  |  |  |  |
| --- | --- | --- | --- |
| Element | Carbon | Hydrogen | Oxygen |
| Symbol | C | H | O |
| Moles present = % composition  Molar mass |  | 9.09 |  |
| Divide by the smallest value | 4.5458 |  |  |
|  | 2.2725 |  | 2.2725 |
| Mole ratios | 2 | 4 | 1 |

Empirical formula is **C2H4O**

The molecular formula is thus determined :

**n**  = Relative formular mass = 88 = 2

Relative empirical formula 44

The molecular formula is (C2H4O ) x 2 = **C4H8O2**.

**4.A hydrocarbon burns completely in excess air to form 5.28 g of carbon (IV) oxide and 2,16g of water.**

**If the molecular mass of the hydrocarbon is 84, draw and name its molecular structure.**

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then: Mass of carbon in CO2 = Mass of C in COx mass of CO2 =>



12 x 5.28

44

Mass of Hydrogen in H2O = Mass of C in H2Ox mass of H2O=>

Molar mass of H2O

2 x 2.16 = **0.24g√**

18

|  |  |  |
| --- | --- | --- |
| Element | Carbon | Hydrogen |
| Symbol | C | H |
| Moles present = mass  Molar mass | 1.44g  12 | 0.24g**√**  1 |
| Divide by the smallest value | 0.12  0.12 | 0.24  0.12 |
| Mole ratios | 1 | 2**√** |

Empirical formula is **CH2√**

The molecular formular is thus determined : **n**  = Relative formular mass = 84 = **6√**

Relative empirical formula 14

The molecular formula is (CH2 ) x 6 = **C6H12**. **√**

molecular name **Hexene√/Hex-1-ene (or any position isomer of Hexene)**

Molecular structure

**H H H H H H**

**C**

**C**

**C**

**C**

**C**

**C**

**H H√**

**H H H H**

**5. Compound A contain 5.2% by mass of Nitrogen .The other elements present are Carbon, hydrogen and Oxygen. On combustion of 0.085g of A in excess Oxygen,0.224g of carbon(IV)oxide and 0.0372g of water was formed. Determine the empirical formula of A (N=14.0, O=16.0 , C=12.0 , H=1.0)**

|  |  |
| --- | --- |
| Mass of **N** in A | = 5.2% x 0.085 **= 0.00442 g** |
| Mass of C in A | = 12 x 0.224 = **0.0611g** |
|  | 44 |
| Mass of H in A | = 2 x 0.0372 = **0.0041g** |
|  | 18 |
| Mass of O in A | = 0.085g – 0.004442g = **0.0806g** (Mass of C,H,O) |
|  | => 0.0611g + 0.0041g = **0.0652g (Mass of C,H)** |

0.0806g (Mass of C,H,O)- 0.0652g (Mass of C,H) = **0.0154 g**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Nitrogen | Carbon | Hydrogen | Oxygen |
| Symbol | N | C | H | O |
| Moles present = mass  Molar mass | 0.00442 g  14 | 0.0611g  12 | 0.0041g  1 | 0.0154 g  16 |
| Divide by the smallest value | 0.00032  0.00032 | 0.00509  0.00032 | 0.0041g  0.00032 | 0.00096  0.00032 |
| Mole ratios | 1 | 16 | 13 | 3 |

**Empirical formula = C16H13NO3**

**(d)Molar gas volume**

The volume occupied by one mole of all gases at the same temperature and pressure is a constant.It is:

1. 24dm3/24litres/24000cm3 at **r**oom **t**emperature(25oC/298K)and **p**ressure(**r.t.p**).

i.e. 1mole of all gases =24dm3/24litres/24000cm3 atr.t.p Examples

**1mole of O2** = 32g =6.0 x1023 particles**= 24dm3/24litres/24000cm3 at r.t.p**

**1mole of H2**= 2g =6.0 x1023 particles **=24dm3/24litres/24000cm3 at r.t.p**

**1mole of CO2**= 44g = 6.0 x1023 particles **=24dm3/24litres/24000cm3 at r.t.p 1mole of NH3**= 17g =6.0 x1023 particles **= 24dm3/24litres/24000cm3 at r.t.p**

**1mole of CH4** = 16g =6.0 x1023 particles **=24dm3/24litres/24000cm3 at r.t.p** (**ii**)22.4dm3/22.4litres/22400cm3 at **s**tandard **t**emperature(0oC/273K) and **p**ressure(**s.t.p**)

i.e. 1mole of all gases =22.4dm3/22.4litres/22400cm3 ats.t.p Examples

**1mole of O2** = 32g =6.0 x1023 particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of H2**= 2g =6.0 x1023 particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of CO2**= 44g = 6.0 x1023 particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of NH3**= 17g =6.0 x1023 particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of CH4** = 16g =6.0 x1023 particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p** The volume occupied by one mole of a gas at r.t.p or s.t.p is commonly called the **molar gas volume.** Whether the molar gas volume is at r.t.p or s.t.p must always be **specified**.

From the above therefore a less or more volume can be determined as in the examples below.

**Practice examples**

**1. Calculate the number of particles present in:**

(Avogadros constant =6.0 x1023mole-1 **)**

1. **2.24dm3 of Oxygen.**

|  |  |  |
| --- | --- | --- |
|  | 22.4dm3 -> | 6.0 x1023 |
|  | 2.24dm3  **->** | 2.24 x 6.0 x1023 |
|  |  | 22.4 |

=6.0 x1022 **molecules** = 2 x 6.0 x1022. = 1.2 x1023 **atoms**

1. **2.24dm3 of Carbon(IV)oxide.**

22.4dm3 ->6.0 x1023

2.24dm3  **->** 2.24 x 6.0 x1023

22.4

=6.0 x1022 **molecules** = (CO2) = 3 x 6.0 x1022. = 1.8 x1023 **atoms**

**2. 0.135 g of a gaseous hydrocarbon X on complete combustion produces 0.41g of carbon(IV)oxide and 0.209g of water.0.29g of X occupy 120cm3 at room temperature and 1 atmosphere pressure .Name X and draw its molecular structure.(C=12.0,O= 16.O,H=1.0,1 mole of gas occupies 24dm3 at r.t.p)**

Molar mass CO2= 44 gmole-1√ Molar mass H2O = 18 gmole-1√

Molar mass X = 0.29 x (24 x 1000)cm3 = 58 gmole-1√

120cm3

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then:

Mass of carbon in CO2 = Mass of C in CO2 x mass of CO2 =>

Molar mass of CO

12 x 0.41

44

Mass of Hydrogen in H2O = Mass of C in H2Ox mass of H2O=>

Molar mass of HO

2 x 0.209

18

|  |  |  |
| --- | --- | --- |
| Element | Carbon | Hydrogen |
| Symbol | C | H |
| Moles present = % composition  Molar mass | 0.g118  12 | 0.0232g√  1 |
| Divide by the smallest value | 0.0093  0.0093 | 0.0232  0.0093√ |
| Mole ratios | 1 x2 | 2.5x2 |
|  | 2 | 5√ |

Empirical formula is **C2H5**√

The molecular formular is thus determined : **n**  = Relative formular mass = 58 = **2**√

Relative empirical formula 29

The molecular formula is (**C2H5** ) x 2 = **C4H10**.√

|  |  |  |
| --- | --- | --- |
| Molecule name **Butane** Molecula structure |  |  |
| H H | H | H |

H C C C C H√

H H H H

(e)**Gravimetric analysis**

Gravimetric analysis is the relationship between reacting masses and the volumes and /or masses of products. All reactants are in mole **ratios** to their products in accordance to their stoichiometric equation. Using the mole ration of reactants and products any volume and/or mass can be determined as in the examples:

1. Calculate the volume of carbon(IV)oxide at r.t.p produced when 5.0 g of calcium carbonate is strongly heated.(Ca=40.0, C= 12.0,O = 16.0,1 mole of gas =22.4 at r.t.p)

Chemical equation

CaCO3(s) -> CaO(s) + CO2(g)

Mole ratios 1: 1: 1

Molar Mass CaCO3 =100g

**Method 1**

|  |  |
| --- | --- |
| 100g CaCO3(s) | -> 24dm3 CO2(g) at r.t.p |
| 5.0 g CaCO3(s) | -> 5.0 g x 24dm3 = **1.2dm3/1200cm3** |
|  | 100g |

**Method 2**

Moles of 5.0 g CaCO3(s) = 5.0 g = **0.05** moles

100 g

Mole ratio 1:1

Moles of CO2(g) = **0.05**moles

Volume of CO2(g) = 0.05 x 24000cm3 =**1200cm3 /1.2dm3**

1. **1.0g of an alloy of aluminium and copper were reacted with excess hydrochloric acid. If 840cm3 of hydrogen at s.t.p was produced, calculate the % of copper in the alloy.(Al =27.0,one mole of a gas at s.t.p =22.4dm3 )**

Chemical equation

Copper does not react with hydrochloric acid

2Al(s) + 6HCl(aq) -> 2AlCl3(aq) + 3H2(g)

Method 1

3H2(g) = 3 moles x (22.4 x 1000)cm3 => 2 x 27 g Al

840cm3 => 840cm3 x 2 x 27 = **0.675**g of Aluminium

3 x 22.4 x 1000

Total mass of alloy – mass of aluminium = mass of copper

=> 1.0g - 0.675g =**0.325**g of copper

% copper = mass of copper x100% = **32.5%**

Mass of alloy

Method 2

Mole ratio 2Al: 3H2 = 2:3

Moles of Hydrogen gas = volume of gas => 840cm3 = **0.0375moles**

Molar gas volume 22400cm3

Moles of Al = 2/3 moles of H2 => 2/3x 0.0375moles = **0.025moles**

Mass of Al = moles x molar mass =>0.025moles x 27 = **0.675g**

Total mass of alloy – mass of aluminium = mass of copper => 1.0g - 0.675g = **0.325** g of copper

% copper = mass of copper x100% = **32.5%**

Mass of alloy

**(f)Gay Lussac‟s law**

Gay Lussacs law states that **“when gases combine/react they do so in simple volume ratios to each other and to their gaseous products at constant/same**

**temperature and pressure”**

Gay Lussacs law thus only apply to gases

Given the volume of one gas reactant, the other gaseous reactants can be deduced thus:

**Examples**

**1. Calculate the volume of Oxygen required to completely react with 50cm3 of Hydrogen.**

Chemical equation: 2H2 (g) + O2 (g) -> 2H2O(l)

Volume ratios 2 : 1 : 0

Reacting volumes 50cm3 : **25cm3**

**50cm3 of Oxygen is used**

**2. Calculate the volume of air required to completely reacts with 50cm3 of**

**Hydrogen.(assume Oxygen is 21% by volume of air)**

Chemical equation: 2H2 (g) + O2 (g) -> 2H2O(l)

Volume ratios 2 : 1 : 0

Reacting volumes 50cm3 : **25cm3** **50cm3 of Oxygen is used**

21% = 25cm3

100% = 100 x 25 =

21

3.If 5cm3 of a hydrocarbon CxHy burn in 15cm3 of Oxygen to form 10cm3 of Carbon(IV)oxide and 10cm3 of water vapour/steam, obtain the equation for the reaction and hence find the value of x and y in CxHy.

|  |  |  |
| --- | --- | --- |
| Chemical equation: CxHy (g) | + O2 (g) | -> H2O(g) + CO2(g) |
| Volumes 5cm3 | : 15cm3 | : 10cm3 : 10cm3 |
| Volume ratios 5cm3 | : 15cm3 | : 10cm3 : 10cm3 |
| (divide by lowest volume) **5** | **5** | **5 5** |

**Reacting volume ratios** 1volume 3 volume 2 volume 2 volume

Balanced chemical equation: **CxHy (g) + 3O2 (g) -> 2H2O(g) + 2CO2(g)**

If “4H” are in 2H2O(g) the y=4 If “2C” are in 2CO2 (g) the x=2

Thus(i) chemical formula of hydrocarbon = **C2H4**

(ii) chemical name of hydrocarbon = **Ethene**

**4.100cm3 of nitrogen (II)oxide NO combine with 50cm3 of Oxygen to form 100cm3 of a single gaseous compound of nitrogen. All volumes measured at the same temperature and pressure. Obtain the equation for the reaction and name the gaseous product.**

|  |  |  |
| --- | --- | --- |
| Chemical equation: NO (g) | + O2 (g) | -> NOx |
| Volumes 100cm3 | : 50cm3 | : 100 |
| Volume ratios 100cm3 | : 50cm3 | : 100cm3 |
| (divide by lowest volume) **50** | **50** | **50** |
| **Reacting volume ratios** 2volume | 1 volume | 2 volume |

Balanced chemical equation: 2 **NO (g) + O2 (g) -> 2NO x(g)**

Thus(i) chemical formula of the nitrogen compound = **2 NO2**

(ii) chemical name of compound = **Nitrogen(IV)oxide**

5.When 15cm3 of a gaseous hydrocarbon was burnt in 100cm3 of Oxygen ,the resulting gaseous mixture occupied70cm3 at room temperature and pressure. When the gaseous mixture was passed through, potassium hydroxide its volume decreased to 25cm3.

(a)What volume of Oxygen was used during the reaction.(1mk)

**Volume of Oxygen used =100-25 =75cm3√**

**(P was completely burnt)**

(b)Determine the molecular formula of the hydrocarbon(2mk) **CxHy + O2 -> xCO2 + yH2O**

**15cm3 : 75cm3**

**15 15**

**1 : 3√**

**=> 1 atom of C react with 6 (3x2)atoms of Oxygen**

**Thus x = 1 and y = 2 => P has molecula formula CH4√**

**(g) Ionic equations**

An ionic equation is a chemical statement showing the movement of ions (cations and anions ) from reactants to products.

**Solids**, **gases** and **liquids** do not ionize/dissociate into free ions. **Only** ionic compounds in **aqueous/solution** or **molten** state ionize/dissociate into free cations and anions (**ions**)

An ionic equation is usually derived from a stoichiometric equation by using the following guidelines

Guidelines for writing ionic equations

1.Write the balanced stoichiometric equation

2.Indicate the state symbols of the reactants and products

3.**Split** into cations and anions all the reactants and products that exist in **aqueous** state.

4.**Cancel out** any cation and anion that appear on **both** the product and reactant side.

5. Rewrite the chemical equation. It is an ionic equation.

**Practice**

(**a**)Precipitation of an insoluble salt

All insoluble salts are prepared in the laboratory from double decomposition /precipitation. This involves mixing **two soluble** salts to form **one soluble** and **one insoluble** salt

1. When silver nitrate(V) solution is added to sodium chloride solution,sodium nitrate(V) solution and a white precipitate of silver chloride are formed.

Balanced stoichiometric equation

AgNO3(aq) + NaCl(aq) -> AgCl (**s**) + NaNO3 (aq)

Split reactants product existing in aqueous state as cation/anion

Ag+(aq) + NO3- (aq) + Na+(aq) + Cl-(aq) -> AgCl(**s**) + Na+(aq)+ NO3- (aq)

Cancel out ions appearing on reactant and product side

Ag+(aq) + NO3- (aq) + Na+(aq) + Cl-(aq) -> AgCl(**s**) + Na+(aq)+ NO3- (aq)

Rewrite the equation

**Ag+(aq) + Cl-(aq) -> AgCl(s) (**ionic equation**)**

1. When barium nitrate(V) solution is added to copper(II)sulphate(VI) solution, copper(II) nitrate(V) solution and a white precipitate of barium sulphate(VI) are formed.

Balanced stoichiometric equation

Ba(NO3)2(aq) + CuSO4(aq) -> BaSO4 (**s**) + Cu(NO3) 2 (aq)

Split reactants product existing in aqueous state as cation/anion

Ba2+(aq) + **2**NO3- (aq) + Cu2+(aq) + SO42-(aq) -> BaSO4 (**s**) + **2**NO3- (aq)+ Cu2+(aq)

Cancel out ions appearing on reactant and product side

Ba2+(aq) + 2NO3- (aq) +Cu2+ (aq) + SO42-(aq)-> BaSO4(**s**) + **2**NO3- (aq) + Cu2+(aq)

Rewrite the equation

**Ba2+(aq) + SO42-(aq) -> BaSO4(s) (**ionic equation**)**

**3.A yellow precipitate of Potassium Iodide is formed from the reaction of Lead(II)nitrate(v) and potassium iodide.**

Balanced stoichiometric equation

Pb(NO3)2(aq) + 2KI(aq) -> PbI2 (**s**) + 2KNO3  (aq)

Split reactants product existing in aqueous state as cation/anion

Pb2+(aq) + **2**NO3- (aq) + 2K +(aq) + 2I - (aq) -> PbI2 (**s**) + **2**NO3- (aq)+ 2K +(aq)

Cancel out ions appearing on reactant and product side

Pb2+(aq) + **2**NO3- (aq) + 2K +(aq) + 2I - (aq) -> PbI2 (**s**) + **2**NO3- (aq)+ 2K +(aq)

Rewrite the equation

**Pb2+(aq) + 2I- (aq) -> PbI2 (s) (**ionic equation**)**

(**b**)Neutralization

Neutralization is the reaction of an acid with a **soluble** base/alkali or **insoluble** base.

(i)Reaction of alkalis with acids

1.Reaction of nitric(V)acid with potassium hydroxide

Balanced stoichiometric equation

HNO3(aq) + KOH(aq) -> H2O (**l**) + KNO3  (aq)

Split reactants product existing in aqueous state as cation/anion

H+(aq) + NO3- (aq) + K +(aq) + OH - (aq) -> H2O (**l**) + NO3- (aq)+ K +(aq)

Cancel out ions appearing on reactant and product side

H+(aq) + NO3- (aq) + K +(aq) + OH - (aq) -> H2O (**l**) + NO3- (aq)+ K +(aq)

Rewrite the equation

**H+ (aq) + OH - (aq) -> H2O (l) (ionic equation)**

2.Reaction of sulphuric(VI)acid with ammonia solution

Balanced stoichiometric equation

H2SO4(aq) + 2NH4OH(aq) -> H2O (**l**) + (NH4) 2SO4  (aq)

Split reactants product existing in aqueous state as cation/anion

2H+(aq) + SO42- (aq) + 2NH4 +(aq)+ 2OH - (aq) ->2H2O (**l**) +SO42- (aq)+ 2NH4 + (aq)

Cancel out ions appearing on reactant and product side

2H+(aq) + SO42- (aq) + 2NH4 +(aq)+ 2OH - (aq) ->2H2O (**l**) +SO42- (aq)+ 2NH4 + (aq)

Rewrite the equation

|  |  |
| --- | --- |
| 2**H+ (aq) + 2OH - (aq) ->** | **2H2O (l)** |
| **H+ (aq) + OH - (aq) ->** | **H2O (l) (ionic equation)** |

3.Reaction of hydrochloric acid with Zinc hydroxide

Balanced stoichiometric equation

**2**HCl(aq) + Zn(OH)2 (**s**) -> 2H2O (**l**) + ZnCl 2  (aq)

Split reactants product existing in aqueous state as cation/anion

2H+(aq) + 2Cl- (aq) + Zn(OH)2 (**s**) ->2H2O (**l**) + 2Cl- (aq)+ Zn 2+ (aq)

Cancel out ions appearing on reactant and product side

2H+(aq) + 2Cl- (aq) + Zn(OH)2 (**s**) ->2H2O (**l**) + 2Cl- (aq)+ Zn 2+ (aq)

Rewrite the equation

2H+(aq) + Zn(OH)2 (**s**) ->2H2O (**l**) + Zn 2+ (aq) **(ionic equation)**

**(h)Molar solutions**

A molar solution is one whose concentration is known. The SI unit of concentration is **Molarity** denoted **M**.

Molarity may be defined as the number of moles of solute present in **one** cubic **decimeter** of solution.

One cubic decimeter is equal to **one litre** and also equal to **1000cm3**. The higher the molarity the higher the concentration and the higher/more solute has been dissolved in the solvent to make one cubic decimeter/ litre/1000cm3 **solution.**

**Examples**

2M sodium hydroxide means 2 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm3 uniform **solution** mixture of sodium hydroxide and water.

0.02M sodium hydroxide means 0.02 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm3 uniform **solution** mixture of sodium hydroxide and water.

“2M” is **more concentrated** than“0.02M”.

**Preparation of molar solution**

Procedure

Weigh accurately 4.0 g of sodium hydroxide pellets into a 250cm3 volumetric flask.

Using a wash bottle add about 200cm3 of distilled water.

Stopper the flask.

Shake vigorously for three minutes.

Remove the stopper for a second then continue to shake for about another two minutes until **all** the solid has dissolved.

Add more water slowly upto **exactly** the 250 cm3 **mark**.

Sample questions

**1.Calculate the number of moles of sodium hydroxide pellets present in:**

1. **4.0 g.**

Molar mass of NaOH = (23 + 16 + 1) = 40g

Moles = Mass => 4.0 = **0.1**  /  **1.0 x 10 -1** moles

Molar mass 40

1. **250 cm3 solution in the volumetric flask.**

Moles in 250 cm3 = **0.1 / 1.0 x 10 -1** moles

1. **one decimeter of solution**

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.0 x 10 -1 moles x 1000cm3 =

250cm3

= **0.4 M / 0.4 molesdm-3**

Method 2

250cm3 solution contain 1.0 x 10 -1 moles

1000cm3 solution = Molarity contain 1000 x 1.0 x 10 -1 moles

250 cm3

= **0.4 M / 0.4 molesdm-3**

**Theoretical sample practice**

**1. Calculate the molarity of a solution containing:**

**(i) 4.0 g sodium hydroxide dissolved in 500cm3 solution**

Molar mass of NaOH = (23 + 16 + 1) = 40g

Moles = Mass => 4.0 = **0.1**  /  **1.0 x 10 -1** moles

Molar mass 40

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.0 x 10 -1 moles x 1000cm3

500cm3

= **0.2 M / 0.2 molesdm-3**

Method 2

500 cm3 solution contain 1.0 x 10 -1 moles

1000cm3 solution = Molarity contain 1000 x 1.0 x 10 -1 moles

500 cm3

= **0.2 M / 0.2 molesdm-3**

1. **5.3 g anhydrous sodium carbonate dissolved in 50cm3 solution** Molar mass of Na2CO3 = (23 x 2 + 12 + 16 x 3) = 106 g

Moles = Mass => 5.3 = **0.05 / 5. 0 x 10-2**  moles

Molar mass 106

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.0 moles x 1000cm3 =

50cm3

=**1.0 M**

Method 2

* 1. cm3 solution contain 5.0 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 5.0 x 10 -2 moles

* + 1. cm3 = **1.0M / 1.0 molesdm-3**

1. **5.3 g hydrated sodium carbonate decahydrate dissolved in 50cm3 solution**

Molar mass of Na2CO3.10H2O = (23 x 2 + 12 + 16 x 3 + 20 x 1 + 10 x 16) =286g

Moles = Mass => 5.3 = **0.0185 / 1.85 x 10 -2**  moles

Molar mass 286

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.85 x 10 -2 moles x 1000cm3 =

50cm3

= **0.37 M/0.37 molesdm-3**

Method 2

* 1. cm3 solution contain 1.85 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 1.85 x 10 -2 moles

* + 1. cm3

= **3.7 x 10**-1 **M / 3.7 x 10**-1 **molesdm-3**

1. **7.1 g of anhydrous sodium sulphate(VI)was dissolved in 20.0 cm3 solution. Calculate the molarity of the solution.**

Method 1

20.0cm3 solution ->7.1 g

1000cm3 solution -> 1000 x 71 = 3550 g **dm-3**

20

Molar mass Na2SO4 = 142 g

Moles **dm-3** = Molarity = Mass 3550 = 2.5 M/ moles**dm-3**

Molar mass 142

Method 2

Molar mass Na2SO4 = 142 g

Moles = Mass => 7.1 = **0.05 / 5.0 x 10 -2**  moles

Molar mass 142

Method 2(a)

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 5.0 x 10 -2 moles x 1000cm3

20cm3

= **2.5 M/2.5 molesdm-3**

Method 2(b)

20 cm3 solution contain 5.0 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 5.0 x 10 -2 moles

20 cm3 = **2.5 M/2.5 molesdm-3**

**(iv) The density of sulphuric(VI) is 1.84gcm-3 Calculate the molarity of the acid.**

Method 1

1.0cm3 solution ->1.84 g

1000cm3 solution -> 1000 x 1.84 = 1840 g **dm-3**

1

Molar mass H2SO4 = 98 g

Moles **dm-3** = Molarity = Mass = 1840

Molar mass 98

= **18.7755 M**/ moles**dm-3**

Method 2

Molar mass H2SO4 = 98 g

Moles = Mass => 1.84 = **0.0188 / 1.88 x 10 -2**  moles

Molar mass 98

Method 2(**a**)

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.88 x 10 -2 moles x 1000cm3

1.0cm3

= **18.8M/18.8 molesdm-3**

Method 2(**b**)

20 cm3 solution contain 1.88 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 1.88 x 10 -2 moles

1.0 cm3 = **18.8M/18.8 molesdm-3**

**2. Calculate the mass of :**

1. **25 cm3 of 0.2M sodium hydroxide solution(Na =23.0.O =16.0, H=1.0)**

Molar mass NaOH = **40g**

Moles in 25 cm3 = Molarity x volume => 0.2 x 25 = **0.005/5.0 x 10-3**moles

1000 1000

Mass of NaOH =Moles x molar mass = 5.0 x 10-3 x 40 = **0.2 g**

1. **20 cm3 of 0.625 M sulphuric(VI)acid (S =32.0.O =16.0, H=1.0)**

Molar mass H2SO4 = **98g**

Moles in 20 cm3 = Molarity x volume=> 0.625 x 20 = **0.0125/1.25.0 x 10-3**moles

1000 1000

Mass of H2SO4 =Moles x molar mass => 5.0 x 10-3 x 40 = **0.2 g**

1. **1.0 cm3 of 2.5 M Nitric(V)acid (N =14.0.O =16.0, H=1.0)**

Molar mass HNO3 = **63 g**

Moles in 1 cm3 = Molarity x volume => 2.5 x 1 = **0.0025 / 2.5. x 10-3**moles

1000 1000

Mass of HNO3 =Moles x molar mass => 2.5 x 10-3 x 40 = **0.1 g**

**3. Calculate the volume required to dissolve :**

**(a)(i) 0.25moles of sodium hydroxide solution to form a 0.8M solution** Volume (in cm3) = moles x 1000 => 0.25 x 1000 = **312.5cm3**

Molarity 0.8

**(ii) 100cm3 was added to the sodium hydroxide solution above. Calculate the concentration of the solution.**

C1 x V1 = C2 x V2 where:

C1 = molarity/concentration before diluting/adding water

C2 = molarity/concentration after diluting/adding water

V1 = volume before diluting/adding water V2 = volume after diluting/adding water

=> 0.8M x 312.5cm3 = C2 x (312.5 + 100**)**

C2 = 0.8M x 312.5cm3 = **0.6061M**

412.5

**(b)(ii) 0.01M solution containing 0.01moles of sodium hydroxide solution .** Volume (in cm3) = moles x 1000 => 0.01 x 1000 = **1000 cm3**

Molarity 0.01

**(ii) Determine the quantity of water which must be added to the sodium hydroxide solution above to form a 0.008M solution.**

C1 x V1 = C2 x V2 where:

C1 = molarity/concentration before diluting/adding water

C2 = molarity/concentration after diluting/adding water V1 = volume before diluting/adding water

V2 = volume after diluting/adding water

=> 0.01M x 1000 cm3 = 0.008 x V2

V2 = 0.01M x 1000cm3 = **1250cm3**

0.008

Volume added = 1250 - 1000 = **250cm3**

**(c)Volumetric analysis/Titration**

Volumetric analysis/Titration is the process of determining unknown concentration of one reactant from a known concentration and volume of another. Reactions take place in simple mole ratio of reactants and products.

Knowing the concentration/ volume of one reactant, the other can be determined from the relationship:

**M1V1 = M2V2** where**:**

**n1 n2**

M1 = Molarity of 1st reactant

M2 = Molarity of 2nd reactant

V1 = Volume of 1st reactant V1 = Volume of 2nd reactant n1 = number of moles of 1st reactant from stoichiometric equation n2 = number of moles of 2nd reactant from stoichiometric equation

Examples

**1.Calculate the molarity of MCO3 if 5.0cm3 of MCO3 react with 25.0cm3 of 0.5M hydrochloric acid.(C=12.0 ,O =16.0)**

Stoichiometric equation:MCO3(s) + 2HCl(aq) -> MCl2(aq) + CO2(g) + H2O(l) Method 1

M1V1 = M2V2 -> M1 x 5.0cm3 = 0.5M x 25.0cm3 n1 n2 1 2

=> M1 = 0.5 x 25.0 x1 = **1.25M / 1.25 moledm-3**

5.0 x 2

Method 2

Moles of HCl used = molarity x volume

1000

=> 0.5 x 25.0 = **0.0125 /1.25 x 10-2moles**

1000

Mole ratio MCO3 : HCl = 1:2

Moles MCO3 = 0.0125 /1.25 x 10-2moles = **0.00625 / 6.25 x 10-3 moles**

2

Molarity MCO3 = moles x 1000 => 0.00625 / 6.25 x 10-3  x 1000

Volume 5

= **1.25M / 1.25 moledm-3**

**2. 2.0cm3 of 0.5M hydrochloric acid react with 0.1M of M2CO3. Calculate the volume of 0.1M M2CO3 used.**

Stoichiometric equation:M2CO3 (aq) + 2HCl(aq) -> 2MCl (aq) + CO2(g) + H2O(l) Method 1

M1V1 = M2V2 -> 0.5 x 2.0cm3 = 0.1M x V2 cm3

n1 n2 2 1

=> V2 = 0.5 x 2.0 x1 = **1.25M / 1.25 moledm-3**

0.1 x 2

Method 2

Moles of HCl used = molarity x volume

1000

=> 0.5 x 2.0 = **0.0125 /1.25 x 10-2moles**

1000

Mole ratio M2CO3 : HCl = 1:2

Moles M2CO3 = 0.0125 /1.25 x 10-2moles = **0.00625 / 6.25 x 10-3 moles**

2

Molarity M2CO3 = moles x 1000 => 0.00625 / 6.25 x 10-3  x 1000

Volume 5

= **1.25M / 1.25 moledm-3**

**3. 5.0cm3 of 0.1M sodium iodide react with 0.1M of Lead(II)nitrate(V).**

**Calculate(i) the volume of Lead(II)nitrate(V) used.**

**(ii)the mass of Lead(II)Iodide formed**

(Pb=207.0, I =127.0)

Stoichiometric equation: 2NaI(aq) + Pb(NO3)2(aq) -> 2NaNO3(aq) + PbI2(**s**)

(i)Volume of Lead(II)nitrate(V) used

Method 1

M1V1 = M2V2 -> 5 x 0.1cm3 = 0.1M x V2 cm3

n1 n2 2 1

=> V2 = 0.1 x 5.0 x 1 = **1.25M / 1.25 moledm-3**

0.1 x 2

Method 2

Moles of HCl used = molarity x volume

1000

=> 0.1 x 5.0 = **0.0125 /1.25 x 10-2moles**

1000

Mole ratio M2CO3 : HCl = 1:2

Moles M2CO3 = 0.0125 /1.25 x 10-2moles = **0.00625 / 6.25 x 10-3 moles**

2

Molarity M2CO3 = moles x 1000 => 0.00625 / 6.25 x 10-3  x 1000

Volume 5

= **1.25M / 1.25 moledm-3**

**4. 0.388g of a monobasic organic acid B required 46.5 cm3 of 0.095M sodium hydroxide for complete neutralization. Name and draw the structural formula of B**

Moles of NaOH used = molarity x volume

1000

=> 0.095 x 46.5 = **0.0044175 /4.4175 x 10-3moles**

1000

Mole ratio B: NaOH = 1:1

Moles B= **0.0044175 /4.4175 x 10-3moles**

Molar mass B = mass => 0.388

moles 0.0044175 /4.4175 x 10-3moles

= **87.8324 gmole-1**

X-COOH = 87.8324 where X is an alkyl group

X =87.8324- 42 = 42.8324=43

By elimination: CH3 = 15 CH3CH2 = 29 **CH3CH2 CH2  = 43**

**Molecula formula : CH3CH2 CH2COOH**

**Molecule name : Butan-1-oic acid**

**Molecular structure**

**H H H O**

**C**

**C**

**C**

**C**

**H OH**

H H H H

**5. 10.5 g of an impure sample containing ammonium sulphate (VI) fertilizer was warmed with 250cm3 of o.8M sodium hydroxide solution.The excess of the alkali was neutralized by 85cm3 of 0.5M hydrochloric acid. Calculate the % of impurities in the ammonium sulphate (VI)fertilizer.**

**(N=14.0,S=32.0,O=16.0, H=1.0)**

Equation for neutralization

NaOH(aq) + HCl(aq) -> NaOH(aq) + H2O(l)

Mole ratio NaOH(aq):HCl(aq)= 1:1

Moles of HCl = Molarity x volume => 0.5 x 85 = **0.0425 moles**

1000 1000

Excess moles of NaOH(aq)= **0.0425 moles**

Equation for reaction with ammonium salt

2NaOH(aq) + (NH4) 2SO4(aq) -> Na 2SO4(aq) + 2NH3 (g)+ 2H2O(l)

Mole ratio NaOH(aq): (NH4) 2SO4(aq)= 2:1

Total moles of NaOH = Molarity x volume => 0.8 x 250 = **0.2 moles**

1000 1000

Moles of NaOH that reacted with(NH4) 2SO4 = 0.2 - 0.0425 = **0.1575moles**

Moles (NH4) 2SO4 = ½ x 0.1575moles = **0. 07875moles**

Molar mass (NH4) 2SO4= **132 gmole-1**

Mass of in impure sample = moles x molar mass =>0. 07875 x 132 = **10.395 g**

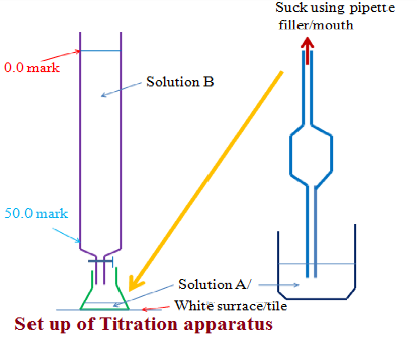
Mass of impurities = 10.5 -10.395 = **0.105 g** % impurities = 0.105 x 100 =  **1.0 %**

10.5

Practically volumetric analysis involves **titration**.

Titration generally involves filling a burette with known/unknown concentration of a solution then adding the solution to unknown/known concentration of another solution in a conical flask until there is complete reaction. If the solutions used are both colourless, an **indicator** is added to the conical flask. When the reaction is over, a **slight**/**little** excess of burette contents **change** the colour of the indicator. This is called the **end point**.

Set up of titration apparatus



The titration process involve involves determination of **titre.** The titre is the volume of burette contents/reading **before** and **after** the end point. Burette contents/reading **before** titration is usually called the **Initial** burette reading.

Burette contents/reading **after** titration is usually called the **Final** burette reading.

The titre value is thus a sum of the **Final** less **Initial** burette readings.

To reduce errors, titration process should be repeated at least once more. The results of titration are recorded in a **titration table** as below

Sample titration table

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | 20.0 | 20.0 | 20.0 |
| Initial burette reading (cm3) | 0.0 | 0.0 | 0.0 |
| Volume of solution used(cm3) | 20.0 | 20.0 | 20.0 |

As **evidence** of a titration **actually** done examining body requires the candidate to record their burette readings before and after the titration.

For KCSE candidates burette readings **must** be recorded **in** a titration table in the **format** **provided** by the Kenya National Examination Council.

As **evidence** of all titration **actually** done Kenya National Examination Council require the candidate to record their burette readings before and after the titration to complete the titration table **in the format** **provided**.

Calculate the average volume of solution used

24.0 + 24.0 + 24.0 = **24.0 cm3**

3

As evidence of understanding the degree of accuracy of burettes , all readings must be recorded to **a** decimal point.

As evidence of accuracy in carrying the out the titration , candidates value should be **within 0.2** of the **school** **value** .

The school value is the **teachers** readings presented to the examining body/council based on the concentrations of the solutions s/he presented to her/his candidates. Bonus mark is awarded for averaged reading **within 0.1** school value as Final answer.

Calculations involved after the titration require candidates **thorough** practical and theoretical **practice mastery** on the:

(**i**)relationship among the mole, molar mass, mole ratios, concentration, molarity.

(**ii**) mathematical application of 1st principles.

Very useful information which candidates forget appears usually in the beginning of the question paper as: “**You are provided with…**”

All calculation must be to the **4th decimal point** unless they divide fully to a lesser decimal point.

Candidates are expected to use a non programmable scientific calculator.

(**a**)Sample Titration Practice 1 **(Simple Titration)**

**You are provided with:**

0.1M sodium hydroxide solution A

Hydrochloric acid solution B

You are required to determine the concentration of solution B in moles per litre.

Procedure

Fill the burette with solution B. Pipette 25.0cm3 of solution A into a conical flask. Titrate solution A with solution B using phenolphthalein indicator to complete the titration table 1

Sample results Titration table 1

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | **20.0** | **20.0** | **20.0** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of solution B used(cm3) | **20.0** | **20.0** | **20.0** |

Sample worked questions

1. **Calculate the average volume of solution B used**

Average titre = Titre 1 + Titre 2 +Titre 3 => ( 20.0 +20.0 +20.0 ) = **20.0cm3**

3 3

1. **How many moles of:**

**(i)solution A were present in 25cm3 solution.**

Moles of solution A = Molarity x volume = 0.1 x 25 = **2.5 x 10-3** moles

1000 1000

**(ii)solution B were present in the average volume.**

Chemical equation: NaOH(aq) + HCl(aq) -> NaCl(aq) + H2O(l)

Mole ratio 1:1 => Moles of A = Moles of B = **2.5 x 10-3** moles

**(iii) solution B in moles per litre.**

Moles of B per litre = moles x 1000 = 2.5 x 10-3 x 1000 = **0.1M**

Volume 20

(**b**)Sample Titration Practice 2 **(Redox Titration)**

**You are provided with:**

Acidified Potassium manganate(VII) solution A

0.1M of an iron (II)salt solution B

8.5g of ammonium iron(II)sulphate(VI) crystals(NH4)2 SO4FeSO4.xH2O solid C

You are required to

(i)standardize acidified potassium manganate(VII)

(ii)determine the value of x in the formula (NH4)2 SO4FeSO4.xH2O.

**Procedure 1**

Fill the burette with solution A. Pipette 25.0cm3 of solution B into a conical flask.

Titrate solution A with solution B until a pink colour just appears. Record your results to complete table 1.

Table 1:**Sample results**

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | **20.0** | **20.0** | **20.0** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of solution A used(cm3) | **20.0** | **20.0** | **20.0** |

Sample worked questions

1. **Calculate the average volume of solution A used**

Average titre = Titre 1 + Titre 2 +Titre 3 => ( 20.0 +20.0 +20.0 ) = **20.0cm3**

3 3

1. **How many moles of:**

**(i)solution B were present in 25cm3 solution.**

Moles of solution A = Molarity x volume = 0.1 x 25 = **2.5 x 10-3** moles

1000 1000

**(ii)solution A were present in the average volume. Assume one mole of B react with five moles of B**

Mole ratio A : B = 1:5

=> Moles of A = Moles of B = 2.5 x 10-3 moles = **5.0 x 10 -4 moles**

5 5

**(iii) solution B in moles per litre.**

Moles of B per litre = moles x 1000 = 2.5 x 10-3 x 1000

Volume 20

= **0.025** M /moles per litre /moles l-1

**Procedure 2**

Place all the solid C into the 250cm3 volumetric flask carefully. Add about 200cm3 of distilled water. Shake to dissolve. Make up to the 250cm3 of solution by adding more distilled water. Label this solution C. Pipette 25cm3 of solution C into a conical flask, Titrate solution C with solution A until a permanent pink colour just appears. Complete table 2.

Table 2:**Sample results**

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | **20.0** | **20.0** | **20.0** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of solution A used(cm3) | **20.0** | **20.0** | **20.0** |

Sample worked questions

1. **Calculate the average volume of solution A used**

Average titre = Titre 1 + Titre 2 +Titre 3 => ( 20.0 +20.0 +20.0 ) = **20.0cm3**

3 3

1. **How many moles of:**

**(i)solution A were present inin the average titre.**

Moles of solution A = Molarity x volume = 0.025 x 20 = **5.0 x 10-4** moles

1000 1000

**(ii)solution C in 25cm3 solution given the equation for the reaction:**

**MnO4- (aq) + 8H+(aq) + 5Fe2+ (aq) -> Mn2+(aq) + 5Fe3+(aq) + 4H2O(l)**

Mole ratio MnO4- (aq): 5Fe2+ (aq) = 1:5 =>

Moles of 5Fe2+ (aq) = Moles of MnO4- (aq) = 5.0 x 10-4 moles = **1.0 x 10 -4 moles**

5 5

**(iii) solution B in 250cm3.**

Moles of B per litre = moles x 250 = 1.0 x 10 -4  x 250 = **1.0 x 10 -3 moles**

Volume 25

1. Calculate the molar mass of solid C and hence the value of x in the chemical formula (NH4)2SO4FeSO4.xH2O. (N=14.0, S=32.0, Fe=56.0, H=1.0 O=16.0)

Molar mass = mass perlitre = 8.5 =  **8500 g** Moles per litre 1.0 x 10 -3 moles NH4)2SO4FeSO4.xH2O = 8500 284 + 18x =8500

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  |  |  |  | 8500 - 284 = 8216 | = | 18x =  **454.4444** |
|  |  |  |  | 18 |  | 18 |

x = **454** (whole number)

**(**c**)**Sample Titration Practice 3 **(Back titration)**

You are provided with:

(i)an impure calcium carbonate labeled M

(ii)Hydrochloric acid labeled solution N

(iii)solution L containing 20g per litre sodium hydroxide.

You are required to determine the concentration of N in moles per litre and the % of calcium carbonate in mixture M.

**Procedure 1**

Pipette 25.0cm3 of solution L into a conical flask. Add 2-3 drops of phenolphthalein indicator. Titrate with dilute hydrochloric acid solution N and record your results in table 1(4mark)

Sample Table 1

1 2 3

Final burette reading (cm3)  **6.5 6.5 6.5**

Initial burette reading (cm3)  **0.0 0.0 0.0**

Volume of N used (cm3)  **6.5 6.5 6.5**

|  |  |
| --- | --- |
| Sample questions |  |
| **(a) Calculate the average volume of solution N used** |  |

6.5 + 6.5 + 6.5 = **6.5** cm3

3

**(b) How many moles of sodium hydroxide are contained in 25cm3of solution L**

Molar mass NaOH =**40**g

Molarity of L = mass per litre => 20 = **0.5**M

Molar mass NaOH 40

Moles NaOH in 25cm3 = molarity x volume => 0.5M x 25cm3 =  **0.0125** moles

1000 1000

**(c)Calculate:**

**(i)the number of moles of hydrochloric acidthat react with sodium hydroxide in (b)above.**

Mole ratio NaOH : HCl from stoichiometric equation= **1:1** Moles HCl =Moles NaOH => **0.0125** moles

**(ii)the molarity of hydrochloric acid solution N.**

Molarity = moles x 1000 => 0.0125 moles x 1000 =**1.9231**M/moledm-3

6.5 6.5

**Procedure 2**

Place the 4.0 g of M provided into a conical flask and add 25.0cm3 of the dilute hydrochloric acid to it using a clean pipette. Swirl the contents of the flask vigorously until effervescence stop.Using a 100ml measuring cylinder add 175cm3 distilled waterto make up the solution up to 200cm3.Label this solution K.Using a clean pipettetransfer 25.0cm3 of the solution into a clean conical flask and titrate with solution L from the burette using 2-3 drops of methyl orange indicator.Record your observations in table 2.

Sample Table 2

1 2 3

Final burette reading (cm3)  **24.5 24.5 24.5**

Initial burette reading (cm3)  **0.0 0.0 0.0**

Volume of N used (cm3)  **24.5 24.5 24.5**

Sample calculations

**(a)Calculate the average volume of solution L used(1mk)**

24.5 + 24.5 + 24.5 = **24.5**cm3

3

**(b)How many moles of sodium hydroxide are present in the average volume of solution L used?**

Moles = molarity x average burette volume => 0.5 x 24.5

1000 1000

= **0.01225 /1.225 x 10-2** moles

**(c) How many moles of hydrochloric acid are present in the original 200cm3 of solution K?**

Mole ratio NaOH: HCl = 1:1 => moles of HCl = **0.01225 /1.225 x 10-2** moles

Moles in 200cm3 = 200cm3 x 0.01225 /1.225 x 10-2moles

25cm3(volume pipetted)

=**0.49 /4.9 x 10-1**moles

**(d)How many moles of hydrochloric acid were contained in original 25 cm3 solution N used**

Original moles = Original molarity x pipetted volume =>

1000cm3

1.9231M/moledm-3 x 25 =  **0.04807/4.807 x 10-2** moles

1000

**(e)How many moles of hydrochloric acid were used to react with calcium carbonate present?**

Moles that reacted = original moles –moles in average titre =>

0.04807/4.807 x 10-2moles - 0.01225 /1.225 x 10-2moles

= **0.03582/3.582 x 10 -2**moles

**(f)Write the equation for the reaction between calcium carbonate and hydrochloric acid.**

CaCO3(s) + 2HCl(aq) -> CaCl2(aq) + CO2(g) + H2O(l)

**(g)Calculate the number of moles of calcium carbonate that reacted with hydrochloric acid.**

From the equation CaCO3(s):2HCl(aq) = 1:2

=> Moles CaCO3(s) = 1/2moles HCl

= 1/2 x 0.03582/3.582 x 10 -2 moles

= **0.01791 /1.791 x 10-2**moles

**(h)Calculate the mass of calcium carbonate in 4.0g of mixture M (Ca=40.0,O =**

**16.0,C=12.0)**

Molar mass CaCO3 = 100g

Mass CaCO3 = moles x molar mass => 0.01791 /1.791 x 10-2moles x 100g

= **1.791**g

**(i)Determine the % of calcium carbonate present in the mixture**

% CaCO3 = mass of pure x 100% => 1.791g x 100% = **44.775**%

Mass of impure 4.0

(d)Sample titration practice 4 (**Multiple titration**)

You are provided with:

(i)sodium L containing 5.0g per litre of a dibasic organic acid H2X.2H2O.

(ii)solution M which is acidified potassium manganate(VII)

(iii)solution N a mixture of sodium ethanedioate and ethanedioic acid

1. 0.1M sodium hydroxide solution P
2. 1.0M sulphuric(VI)

You are required to:

(i)standardize solution M using solution L

(ii)use standardized solution M and solution P to determine the % of sodium ethanedioate in the mixture.

**Procedure 1**

Fill the burette with solution M. Pipette 25.0cm3 of solution L into a conical flask. Heat this solution to about 70oC(**but not to boil**).Titrate the hot solution L with solution M until a permanent pink colour just appears .Shake thoroughly during the titration. Repeat this procedure to complete table 1.

Sample Table 1

1 2 3

Final burette reading (cm3)  **24.0 24.0 24.0**

Initial burette reading (cm3)  **0.0 0.0 0.0**

Volume of N used (cm3)  **24.0 24.0 24.0**

Sample calculations

**(a)Calculate the average volume of solution L used (1mk)**

24.0 + 24.0 + 24.0 = **24.0**cm3

3

**(b)Given that the concentration of the dibasic acid is 0.05molesdm-3.determine the value of x in the formula H2X.2H2O (H=1.0,O=16.0)**

Molar mass H2X.2H2O= mass per litre => 5.0g/litre = **100**g

Moles/litre 0.05molesdm-3

H2X.2H2O =100

X = 100 – ((2 x1) + 2 x (2 x1) + (2 x 16) => 100 – 34 = **66**

**(c) Calculate the number of moles of the dibasic acid H2X.2H2O.**

Moles = molarity x pipette volume => 0.5 x 25 = **0.0125/1.25 x10 -2** moles

1000 1000

**(d)Given the mole ratio manganate(VII)(MnO4-): acid H2X is 2:5, calculate the number of moles of manganate(VII) (MnO4-) in the average titre.**

Moles H2X = 2/5 moles of MnO4-

=> 2/5 x 0.0125/1.25 x10 -2 moles

= **0.005/5.0 x 10 -3**moles

**(e)Calculate the concentration of the manganate(VII)(MnO4-) in moles per litre.**

Moles per litre/molarity = moles x 1000

average burette volume

=>0.005/5.0 x 10 -3moles x 1000 =  **0.2083** molesl-1/M

24.0

**Procedure 2**

With solution M still in the burette ,pipette 25.0cm3 of solution N into a conical flask. Heat the conical flask containing solution N to about 70oC.Titrate while hot with solution M.Repeat the experiment to complete table 2.

Sample Table 2

1 2 3

Final burette reading (cm3)  **12.5 12.5 12.5**

Initial burette reading (cm3)  **0.0 0.0 0.0**

Volume of N used (cm3)  **12.5 12.5 12.5**

Sample calculations

**(a)Calculate the average volume of solution L used (1mk)**

12.5 + 12.5 + 12.5 =**12.5**cm3

3

**(b)Calculations:**

**(i)How many moles of manganate(VII)ions are contained in the average volume of solution M used?**

Moles = molarity of solution M x average burette volume

1000

=> 0.2083 molesl-1/ M x 12.5 = **0.0026 / 2.5 x 10-3** moles

1000

**(ii)The reaction between manganate(VII)ions and ethanedioate ions that reacted with is as in the equation:**

**2MnO4- (aq) + 5C2O42- (aq) + 16H+ (aq) -> 2Mn2+(aq) + 10CO2(g) + 8H2O(l)**

**Calculate the number of moles of ethanedioate ions that reacted with manganate (VII) ions in the average volume of solution M.**

From the stoichiometric equation,mole ratio MnO4- (aq): C2O42- (aq) = 2:5

=> moles C2O42- = 5/2 moles MnO4- => 5/2 x 0.0026 / 2.5 x 10-3 moles

= **0.0065 /6.5 x10-3** moles

**(iii)Calculate the number of moles of ethanedioate ions contained in 250cm3 solution N.**

25cm3 pipette volume -> 0.0065 /6.5 x10-3 moles

250cm3 ->

0.0065 /6.5 x10-3 moles x 250 = **0.065 / 6.5 x10-2** moles

25

**Procedure 3**

Remove solution M from the burette and rinse it with distilled water. Fill the burette with sodium hydroxide solution P. Pipette 25cm3 of solution N into a conical flask and add 2-3 drops of phenolphthalein indicator. Titrate this solution N with solution P from the burette. Repeat the procedure to complete table 3.

Sample Table 2

1 2 3

Final burette reading (cm3)  **24.9 24.9 24.9**

Initial burette reading (cm3)  **0.0 0.0 0.0**

Volume of N used (cm3)  **24.9 24.9 24.9**

Sample calculations

**(a)Calculate the average volume of solution L used (1mk)**

24.9 + 24.9 + 24.9 =  **24.9 cm3**

3

**(b)Calculations:**

**(i)How many moles of sodium hydroxide solution P were contained in the average volume?**

Moles = molarity of solution P x average burette volume

1000

=> 0.1 molesl-1 x 24.9 = **0.00249 / 2.49 x 10-3** moles

1000

**(ii)Given that NaOH solution P reacted with the ethanedioate ions from the acid only and the equation for the reaction is:**

**2NaOH(aq) + H2C2O4 (aq) -> Na2C2O4(g) + 2H2O(l)**

**Calculate the number of moles of ethanedioic acid that were used in the reaction**

From the stoichiometric equation,mole ratio NaOH(aq): H2C2O4 (aq) = 2:1

=> moles H2C2O4 = 1/2 moles NaOH => 1/2 x 0.00249 / 2.49 x 10-3 moles

= **0.001245/1.245 x10-3** moles.

**(iii)How many moles of ethanedioic acid were contained in 250cm3 of solution N?**

25cm3 pipette volume -> 0.001245/1.245 x10-3 moles

250cm3 ->

0.001245/1.245 x10-3 moles x 250 = **0.01245/1.245 x10-2** moles

25

**(iii)Determine the % by mass of sodium ethanedioate in the micture**

**(H= 1.0,O=16.0,C=12.0 and total mass of mixture =2.0 g in 250cm3 solution)**

Molar mass H2C2O4 = **90.0**g

Mass of H2C2O4 in 250cm3 = moles in 250cm3 x molar mass H2C2O4

=>0.01245/1.245 x10-2 moles x 90.0

= **1.1205**g

% by mass of sodium ethanedioate

=(Mass of mixture - mass of H2C2O4) x 100%

Mass of mixture

=> 2.0 - 1.1205 g =  **43.975**%

2.0

**Note**

1. **L is 0.05M Oxalic acid**
2. **M is 0.01M KMnO4**
3. **N is 0.03M oxalic acid(without sodium oxalate)**

**Practice example 5.(Determining equation for a reaction)**  You are provided with

-0.1M hydrochloric acid solution A

-0.5M sodium hydroxide solution B

You are to determine the equation for thereaction between solution A and B **Procedure**

Fill the burette with solution A.Using a pipette and pipette filler transfer 25.0cm3 of solution B into a conical flask.Add 2-3 drops of phenolphthalein indicator.Run solution A into solution B until a permanent pink colour just appears.Record your results in Table 1.Repeat the experiment to obtain three concordant results to complete Table 1 Table 1(Sample results)

|  |  |  |  |
| --- | --- | --- | --- |
| Titration | 1 | 2 | 3 |
| Final volume(cm3) | **12.5** | **25.0** | **37.5** |
| Initial volume(cm3) | **0.0** | **12.5** | **25.0** |
| Volume of solution A used(cm3) | **12.5** | **12.5** | **12.5** |

**Sample questions**

**Calculate the average volume of solution A used.**

12.5+12.5+12.5 = **12.5cm3**

**3**

**Theoretical Practice examples**

**1. 1.0g of dibasic acid HOOC(CH2)xCOOH was dissolved in 250cm3 solution. 25.0 cm3 of this solution reacted with 30.0cm3 of 0.06M sodium hydroxide solution. Calculate the value of x in HOOC(CH2)xCOOH.**

**(C=12.0,H=1.0,O=16.)**

Chemical equation

2NaOH(aq) + H**2**X(aq) -> Na**2**X (aq) + 2H**2**O(aq)

Mole ratio NaOH(aq) :H**2**X(aq) = 2:1

Method 1

Ma Va = na => Ma x 25.0 = 1 => Ma =0.06 x 30.0 x1 Mb Vb = nb 0.06 x 30.0 2 25.0 x 2

Molarity of acid = **0.036M/Mole l-1**

Mass of acid per lite = 1.0 x1000 = **4.0 g/l**

250

0.036M/ Mole l-1  -> 4.0 g /l

1 mole= molar mass of HOOC(CH2)xCOOH = 4.0 x 1 = **111.1111** g

0.036

Molar mass (CH2)x = 111.1111 – (HOOCCOOH = 90.0) **= 21.1111**

(CH2)x = 14x = 21.1111 =  **1.5 = 1 (whole number)**

14

Method 2

Moles of sodium hydroxide = Molarity x volume = 0.06 x 30 = **1.8 x 10 -3moles**

1000

Moles of Hydrochloric acid = 1/2 x 1.8 x 10 -3moles = 9.0 x10 -4moles

Molarity of Hydrochloric acid = moles x 1000 = 9.0 x10 -4moles x1000 Volume 25

Molarity of acid = **0.036M/Mole l-1**

Mass of acid per lite = 1.0 x1000 = **4.0 g/l**

250

0.036M/ Mole l-1  -> 4.0 g /l

1 mole= molar mass of HOOC(CH2)xCOOH = 4.0 x 1 = **111.1111** g

0.036

Molar mass (CH2)x = 111.1111 – (HOOCCOOH = 90.0) **= 21.1111**

(CH2)x = 14x = 21.1111 = **1.5 = 1 (whole number)**

14

**2. 20.0cm3 of 0.05 M acidified potassium manganate(VII)solution oxidized 25.0cm3 of Fe2+(aq) ions in 40.0g/l of impure Iron (II)sulphate(VI) to Fe3+(aq) ions. Calculate the percentage impurities in the Iron (II)sulphate(VI).**

**MnO4- (aq) + 8H+(aq)+ 5Fe2+(aq)-> 5Fe3+(aq) + Mn2+(aq) + 4H2O(aq) Fe=56.0,S= 32.0, O=16.0).**

Moles of MnO4- (aq) = Molarity x volume = 0.05 x 20.0 = **0.001 Moles**

1000 1000

Mole ratio MnO4- (aq): 5Fe2+(aq)= 1:5

Moles 5Fe2+(aq) **=** 5 x0.001  **= 0.005 Moles**

Moles of 5Fe2+(aq) per litre/molarity = Moles x 1000 = 0005 x 1000

Volume 25.0

**= 0.2 M/ Moles/litre**

Molar mass =FeSO4=**152 g**

Mass of in the mixture = Moles x molar mass => 0.2 x 152 = **30.4** g

Mass of impurity = 40.0 – 30.4 =**9.6 g**

% impurity **=** 9.6 g x100 = **24.0 % impurity**

40.0

**3.9.7 g of a mixture of Potassium hydroxide and Potassium chloride was dissolved to make one litre solution.20.0cm3 of this solution required 25.0cm3 of 0.12M hydrochloric acid for completed neutralization. Calculate the percentage by mass of Potassium chloride.(K=39.0,Cl= 35.5)**

Chemical equation

KOH(aq) + HCl(aq) -> KCl(aq) + H2O(l)

Moles of HCl = Molarity x volume => 0.12 x 25.0 = **0.003/3.0 x 10 -3** moles

1000 1000

Mole ratio KOH(aq) : HCl(aq) -= 1:1

Moles KOH =**0.003/3.0 x 10 -3** moles

Method 1

Molar mass KOH =**56.0**g

Mass KOH in 25cm3 =0.003/3.0 x 10 -3 moles x56.0 = 0.168g

Mass KOH in 1000cm3/1 litre = 0.168 x1000= **8.4 g/l**

20

Mass of KCl = 9.7g - 8.4g = **1.3** g

% of KCl = 1.3 x 100 = **13.4021%**

9.7

Method 2

Moles KOH in 1000cm3 /1 litre = Moles in 20cm3 x 1000 =>0.003 x 1000

20 20

=**0.15M/Moles /litre**

Molar mass KOH =**56.0**g

Mass KOH in 1000/1 litre = 0.15M/Moles /litre x 56.0 = **8.4g/l**

Mass of KCl = 9.7g - 8.4g = **1.3** g

% of KCl = 1.3 x 100 = **13.4021%**

9.7

4.**A certain carbonate, GCO3, reacts with dilute hydrochloric acid according to the equation given below:**

**GCO3(s) + 2HCl(aq) -> GCl2 (aq) + CO2 (g) + H2O(l)**

**If 1 g of the carbonate reacts completely with 20 cm3 of 1 M hydrochloric acid**

**,calculate the relative atomic mass of G (C = 12.0 = 16.0)**

Moles of HCl = Molarity x volume=> 1 x20 = **0.02 moles**

1000 1000

Mole ratio HCl; GCO3  = 2:1

Moles of GCO3= 0.02 moles = **0.01moles**

2

Molar mass of GCO3  = mass => 1 = **100** g

moles 0.01moles

G= GCO3 - CO3 =>100g – (12+ 16 x3 = 60) = **40**(**no units**)

**5. 46.0g of a metal carbonate MCO3 was dissolved 160cm3 of 0.1M excess hydrochloric acid and the resultant solution diluted to one litre.25.0cm3 of this solution required 20.0cm3 of 0.1M sodium hydroxide solution for**

**complete neutralization. Calculate the atomic mass of „M‟** Equation

Chemical equation

NaOH(aq) + HCl(aq) -> KCl(aq) + H2O(l)

Moles of NaOH = Molarity x volume=> 0.1 x20 = **0.002 moles**

1000 1000

Mole ratio HCl; NaOH = 1:1

Excess moles of HCl =  **0.002 moles**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | 25cm3 -> | 0.002 moles |  |  |
|  | 1000cm3 -> | 1000 x 0.002 | = **0.08moles** |  |
|  |  | 25cm3 |  |  |

Original moles of HCl = Molarity x volume => 1M x 1litre = **1.0 moles**

Moles of HCl reacted with MCO3 = **1.0 - 0.08 moles** =  **0.92moles**

**Chemical equation**

MCO3(s) + 2HCl(aq) -> MCl2 (aq) + CO2 (g) + H2O(l)

Mole ratio MCO3(s) : HCl(aq) =1:2

Moles of MCO3  = 0.92moles => **0.46moles**

2

Molar mass of MCO3= mass => 46g = **100 g**  moles 0.46moles

M= MCO3 - CO3 =>100g – (12+ 16 x3 = 60) = **40**

**6. 25.0cm3 of a mixture of Fe2+ and Fe3+ ions in an aqueous salt was acidified with sulphuric(VI)acid then titrated against potassium manganate(VI).The salt required 15cm3 ofe0.02M potassium manganate(VI) for complete reaction.**

**A second 25cm3 portion of the Fe2+ and Fe3+ ion salt was reduced by Zinc then titrated against the same concentration of potassium manganate(VI).19.0cm3 of potassium manganate(VI)solution was used for complete reaction. Calculate the concentration of Fe2+ and Fe3+ ion in the solution on moles per litre.**

Mole ratio Fe2+ :Mn04- = 5:1

Moles Mn04- used = 0.02 x 15 = **3.0 x 10-4** moles

1000

Moles **Fe2+ =**  3.0 x 10-4 moles **= 6.0 x 10-5** moles

5

Molarity of **Fe2+ = 6.0 x 10-4** moles x 1000 =  **2.4 x 10-3 moles l-1**

25

Since Zinc reduces **Fe3+** to  **Fe2+**  in the mixture:

Moles Mn04- that reacted with all **Fe2+**= 0.02 x 19 = **3.8 x 10-4** moles

1000

Moles of all **Fe2+** = 3.8 x 10-4 moles = 7.6 x 10-5 moles

5

Moles of Fe3+ = 3.8 x 10-4 - 6.0 x 10-5 = **1.6 x 10-5** moles

Molarity of **Fe3+ =** 1.6 x 10-5 moles x 1000 = **4.0 x 10-4 moles l-1**

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ORGNIC CHEMISTRY

f

Julius G. Thungu

Comprehensive Tutorial notes

**Introduction to Organic chemistry**

Organic chemistry is the branch of chemistry that studies carbon compounds present in living things, once living things or synthetic/man-made.

Compounds that makes up living things whether alive or dead mainly contain carbon. Carbon is tetravalent.

It is able to form stable covalent bonds with itself and many non-metals like hydrogen, nitrogen ,Oxygen and halogens to form a variety of compounds. This is because:

(i) carbon uses all the four valence electrons to form four strong covalent bond.

(ii)carbon can covalently bond to form a single, double or triple covalent bond with itself.

(iii)carbon atoms can covalently bond to form a very long chain or ring.

When carbon covalently bond with Hydrogen, it forms a group of organic compounds called **Hydrocarbons**

**A. HYDROCARBONS (HCs)**

Hydrocarbons are a group of organic compounds containing /made up of hydrogen and carbon atoms only.

Depending on the type of bond that exist between the individual carbon atoms, hydrocarbon are classified as:

(i) Alk**a**nes

(ii) Alk**e**nes

(iii) Alk**y**nes

**(i) Alkanes**

**(a)Nomenclature/Naming**

These are hydrocarbons with a general formula **CnH2n+2** where **n** is the number of Carbon atoms in a molecule.

The carbon atoms are linked by single bond to each other and to hydrogen atoms.

They include:

|  |  |  |  |
| --- | --- | --- | --- |
| n | General/  Molecular  formula | Structural formula | Name |
| 1 | CH4 | H  H C H  H | Meth**ane** |
| 2 | C2H6 | H H  H C C H  H H | Eth**ane** |
| 3 | C3H8 | H H H  H C C C H  H H H | Prop**ane** |
| 4 | C4H10 | H H H H  H C C C C H  H H H H | But**ane** |
| 5 | C5H12 | H H H H H  H C C C C C H CH3 (CH2) 6CH3  H H H H H | Pent**ane** |
| 6 | C6H14 | H H H H H H  H C C C C C C H CH3 (CH2) 6CH3  H H H H H H | Hex**ane** |
| 7 | C7H16 | H H H H H H H  H C C C C C C C H  H H H H H H H | Hept**ane** |
| 8 | C8H18 | H H H H H H H H  H C C C C C C C C H  H H H H H H H H | Oct**ane** |
| 9 | C9H20 | H H H H H H H H H  H C C C C C C C C C H  H H H H H H H H H | Non**ane** |
| 10 | C10H22 | H H H H H H H H H H  H C C C C C C C C C C H  H H H H H H H H H H | dec**ane** |

**Note**

1.The **general formula**/**molecular formular** of a compound shows the number of each atoms of elements making the compound e.g.

Decane has a general/molecular formula **C10H22** ;this means there are 10 carbon atoms and 22 hydrogen atoms in a molecule of decane.

2.The **structural formula** shows the arrangement/bonding of atoms of each element making the compound e.g

Decane has the structural formula as in the table above ;this means the 1st carbon from left to right is bonded to three hydrogen atoms and one carbon atom.

The 2nd carbon atom is joined/bonded to two other carbon atoms and two Hydrogen atoms.

3.Since carbon is **tetravalent** ,each atom of carbon in the alkane **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons.

4.Since Hydrogen is **monovalent** ,each atom of hydrogen in the alkane **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

5.One member of the alkane differ from the next/previous by a CH2  group.

e.g

Propane differ from ethane by one carbon and two Hydrogen atoms form ethane. Ethane differ from methane also by one carbon and two Hydrogen atoms

6.A group of compounds that differ by a CH2  group from the next /previous **consecutively** is called a **homologous series**.

7.A homologous series:

(i) differ by a CH2  group from the next /previous consecutively

(ii)have similar chemical properties

(iii)have similar chemical formula that can be represented by a general formula e.g alkanes have the general formulaCnH2n+2.

(iv)the physical properties (e.g.melting/boiling points)show steady gradual change)

8.The 1st four alkanes have the prefix **meth\_,eth\_,prop\_** and **but\_** to represent 1,2,3 and 4 carbons in the compound. All other use the numeral prefix **pent\_,Hex\_,hept\_** , etc to show also the number of carbon atoms.

9.If one hydrogen atom in an alkane is removed, an alkyl group is formed.e.g

|  |  |  |  |
| --- | --- | --- | --- |
| Alkane name | molecular structure  CnH2n+**2** | **Alkyl name** | **Molecula structure**  **CnH2n+1** |
| methane | CH4 | **methyl** | **CH3** |
| ethane | CH3CH3 | **ethyl** | **CH3 CH2** |
| propane | CH3 CH2 CH3 | **propyl** | **CH3 CH2 CH2** |
| butane | CH3 CH2 CH2 CH3 | **butyl** | **CH3 CH2 CH2 CH2** |

**(b)Isomers of alkanes**

Isomers are compounds with the same molecular **general formula** but different molecular **structural formula**.

Isomerism is the existence of a compounds having the same general/molecular formula but different structural formula.

The 1st three alkanes do not form isomers.Isomers are named by using the IUPAC(**I**nternational **U**nion of **P**ure and **A**pplied **C**hemistry) system of **nomenclature/naming.**

The IUPAC system of nomenclatureuses the following basic rules/guidelines:

1.Identify the longest continuous carbon chain to get/determine the parent alkane.

2.Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane. Name them fluoro-,chloro-,bromo-,iodo- if they are halogens

4.Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of branches attached to the parent alkane.

Practice on IUPAC nomenclature of alkanes

**(a)Draw the structure of:**

**(i)2-methylpentane**

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

Butane is the parent name **CH3 CH2 CH2 CH3**

2. Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon **“2”**

3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

Position of the branch at carbon “2”

Number of branches at carbon “1”

Type of the branch “methyl” hence

Molecular formula

**CH3**

**CH3 CH CH2 CH3 //CH3 CH (**CH3 **) CH2CH3**

Structural formula

H H H H

H C C C C H

H H H

**H C H**

**H**

**(ii)2,2-dimethylpentane**

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

Butane is the parent name **CH3 CH2 CH2 CH3**

2. Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon **“2”**

3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

Position of the branch at carbon “2”

Number of branches at carbon “2”

Type of the branch two“methyl” hence

Molecular formular

**CH3**

**CH3 C CH2 CH3 //CH3 C (**CH3 **)2 CH2CH3**

**CH3**

Structural formula

H

H C H

H H H

H C C C C H

H H H

**H C H**

**H**

**(iii) 2,2,3-trimethylbutane**

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

Butane is the parent name **CH3 CH2 CH2 CH3**

2. Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon **“2 and 3”**

3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

Position of the branch at carbon “2 and 3”

Number of branches at carbon “3”

Type of the branch three “methyl” hence

Molecular formular

**CH3**

**CH3 C CH CH3 //CH3 C (**CH3 **)3 CH2CH3**

**CH3** **CH3**

Structural formula

H

H C H

H H

H C C C H

H H

H

HC C H

H

H C H

H

**(iv) 1,1,1,2,2,2-hexabromoethane**

Molecular formula

**CBr3 CBr3**

Structural formula

Br Br

Br C C Br

Br Br

**(v) 1,1,1-tetrachloro-2,2-dimethylbutane**

**CH3**

**CCl 3 C CH3 //C Cl 3 C (**CH3 **)2 CH3**

**CH3**

Structural formula

Cl

Cl C Cl

H H

H C C C H

H H

HC H

H

**(c)Occurrence and extraction**

**Crude oil** ,**natural gas** and **biogas** are the main sources of alkanes:

(i)Natural gas is found on top of crude oil deposits and consists mainly of methane.

(ii)Biogas is formed from the decay of waste organic products like animal dung and cellulose. When the decay takes place in absence of oxygen , 60-75% by volume of the gaseous mixture of methane gas is produced.

(iii)Crude oil is a mixture of many flammable hydrocarbons/substances. Using fractional distillation, each hydrocarbon fraction can be separated from the other. The hydrocarbon with lower /smaller number of carbon atoms in the chain have lower boiling point and thus collected first.

As the carbon **chain increase**, the **boiling** point, **viscosity** (ease of flow) and colour **intensity** increase as **flammability** decrease. Hydrocarbons in crude oil are not pure. They thus have no sharp fixed boiling point.

Uses of different crude oil fractions

|  |  |  |
| --- | --- | --- |
| Carbon atoms in a molecule | Common name of fraction | Uses of fraction |
| 1-4 | Gas | L.P.G gas for domestic use |
| 5-12 | Petrol | Fuel for petrol engines |
| 9-16 | Kerosene/Paraffin | Jet fuel and domestic lighting/cooking |
| 15-18 | Light diesel | Heavy diesel engine fuel |
| 18-25 | Diesel oil | Light diesel engine fuel |
| 20-70 | Lubricating oil | Lubricating oil to reduce friction. |
| Over 70 | Bitumen/Asphalt | Tarmacking roads |

**(d)School laboratory preparation of alkanes**

In a school laboratory, alkanes may be prepared from the reaction of a sodium alkanoate with solid sodium hydroxide/soda lime.

Chemical equation:

Sodium alkanoate + soda lime -> alkane + Sodium carbonate

CnH2n+1COONa(s) + NaO**H**(s) -> C n H2n+2 + Na2CO3(s)

The “H” in NaO**H** is transferred/moves to the CnH2n+1in CnH2n+1COONa(s) to form C n H2n+**2**.

Examples

1. **Meth**ane is prepared from the heating of a mixture of sodium **ethan**oate and soda lime and collecting over water

Sodium **ethan**oate + soda lime -> **meth**ane + Sodium carbonate

CH3COONa(s) + NaO**H**(s) -> CH4 + Na2CO3(s)

The “H” in NaO**H** is transferred/moves to the CH3in CH**3**COONa(s) to form CH**4**.

2. **Eth**ane is prepared from the heating of a mixture of sodium **prop**anoate and soda lime and collecting over water

Sodium **prop**anoate + soda lime -> **eth**ane + Sodium carbonate

CH3 CH2COONa(s) + NaO**H**(s) -> CH3 CH3 + Na2CO3(s)

The “H” in NaO**H** is transferred/moves to the CH3 CH2in CH3 CH2COONa (s) to form CH3 CH**3**

3. **Prop**ane is prepared from the heating of a mixture of sodium **but**anoate and soda lime and collecting over water

Sodium **but**anoate + soda lime -> **prop**ane + Sodium carbonate

CH3 CH2CH2COONa(s) + NaO**H**(s) -> CH3 CH2CH3 + Na2CO3(s)

The “H” in NaO**H** is transferred/moves to the CH3 CH2 CH2in CH3 CH2CH2COONa (s) to form CH3 CH2CH**3**

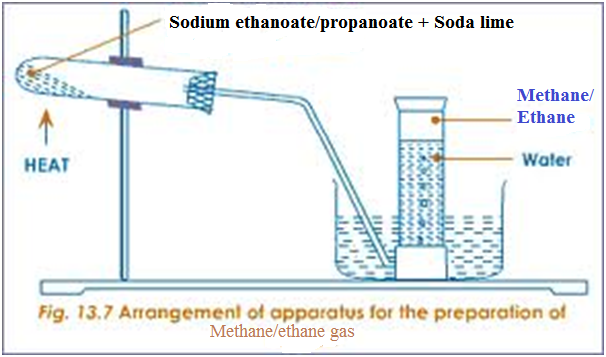
4. **But**ane is prepared from the heating of a mixture of sodium **pent**anoate and soda lime and collecting over water

Sodium **pent**anoate + soda lime -> **but**ane + Sodium carbonate

CH3 CH2 CH2CH2COONa(s)+NaO**H**(s) -> CH3 CH2CH2CH3 + Na2CO3(s)

The “H” in NaO**H** is transferred/moves to the CH3CH2 CH2 CH2in CH3 CH2CH2 CH2COONa (s) to form CH3 CH2 CH2CH**3**

Laboratory set up for the preparation of alkanes



**(d)Properties of alkanes**

**I. Physical properties**

Alkanes are colourless gases, solids and liquids that are not poisonous.

They are slightly soluble in water.

The solubility decrease as the carbon chain and thus the molar mass increase

The melting and boiling point increase as the carbon chain increase.

This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase.

The 1st four straight chain alkanes (methane,ethane,propane and butane)are therefore gases ,the nect six(pentane ,hexane, heptane,octane,nonane, and decane) are liquids while the rest from unidecane(11 carbon atoms) are solids .

The density of straight chain alkanes increase with increasing carbon chain as the intermolecular forces increases.

This reduces the volume occupied by a given mass of the compound.

Summary of physical properties of alkanes

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Alkane | General formula | Melting point(K) | Boiling point(K) | Density gcm-3 | State at room(298K) temperature and pressure atmosphere (101300Pa) |
| Methane | CH4 | 90 | 112 | 0.424 | gas |
| Ethane | CH3CH3 | 91 | 184 | 0.546 | gas |
| Propane | CH3CH2CH3 | 105 | 231 | 0.501 | gas |
| Butane | CH3(CH2)2CH3 | 138 | 275 | 0.579 | gas |
| Pentane | CH3(CH2)3CH3 | 143 | 309 | 0.626 | liquid |
| Hexane | CH3(CH2)4CH3 | 178 | 342 | 0.657 | liquid |
| Heptane | CH3(CH2)5CH3 | 182 | 372 | 0.684 | liquid |
| Octane | CH3(CH2)6CH3 | 216 | 399 | 0.703 | liquid |
| Nonane | CH3(CH2)7CH3 | 219 | 424 | 0.708 | liquid |
| Octane | CH3(CH2)8CH3 | 243 | 447 | 0.730 | liquid |

**II.Chemical properties**

**(i)Burning.**

Alkanes burn with a **blue**/non-luminous **non-sooty**/non-smoky flame in **excess** air to form carbon(IV) oxide and water.

Alkane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

Alkanes burn with a **blue**/non-luminous **no-sooty**/non-smoky flame in **limited** air to form carbon(II) oxide and water.

Alkane + Air -> carbon(**II**) oxide + water (limited air)

Examples

1.(a) Methane when ignited burns with a **blue** **non** **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Methane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

CH4(g) + 2O2(g) -> CO2(g) + 2H2O(l/g)

(b) Methane when ignited burns with a **blue** **non** **sooty** flame in **limited** air to form carbon(II) oxide and water.

Methane + Air -> carbon(**II**) oxide + water (excess air/oxygen)

2CH4(g) + 3O2(g) -> 2CO(g) + 4H2O(l/g)

2.(a) Ethane when ignited burns with a **blue** **non** **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Ethane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

2C2H6(g) + 7O2(g) -> 4CO2(g) + 6H2O(l/g)

(b) Ethane when ignited burns with a **blue** **non** **sooty** flame in **limited** air to form carbon(II) oxide and water.

Ethane + Air -> carbon(**II**) oxide + water (excess air/oxygen)

2C2H6(g) + 5O2(g) -> 4CO(g) + 6H2O(l/g)

3.(a) Propane when ignited burns with a **blue** **non** **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Propane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

C3H8(g) + 5O2(g) -> 3CO2(g) + 4H2O(l/g)

(b) Ethane when ignited burns with a **blue** **non** **sooty** flame in **limited** air to form carbon(II) oxide and water.

Ethane + Air -> carbon(**II**) oxide + water (excess air/oxygen)

2C3H8(g) + 7O2(g) -> 6CO(g) + 8H2O(l/g)

**ii)Substitution**

Substitution reaction is one in which a hydrogen atom is replaced by a halogen in presence of ultraviolet light.

Alkanes react with halogens in presence of ultraviolet light to form halogenoalkanes.

During substitution:

(i)the halogen molecule is split into free atom/radicals.

(ii)one free halogen radical/atoms knock /remove one hydrogen from the alkane leaving an alkyl radical.

(iii) the alkyl radical combine with the other free halogen atom/radical to form halogenoalkane.

(iv)the chlorine atoms substitute repeatedly in the alkane. Each substitution removes a hydrogen atom from the alkane and form hydrogen halide.

(v)substitution stops when all the hydrogen in alkanes are replaced with halogens.

Substitution reaction is a highly explosive reaction in presence of **sunlight / ultraviolet** light that act as **catalyst**.

Examples of substitution reactions

Methane has no effect on bromine or chlorine in diffused light/dark. In sunlight , a mixture of chlorine and methane explode to form colourless mixture of chloromethane and hydrogen chloride gas. The pale green colour of chlorine gas fades.

Chemical equation

1.(a)Methane + chlorine -> Chloromethane + Hydrogen chloride

CH4(g) + Cl2(g) -> CH3Cl (g) + HCl (g)

H H

H C **H** + **Cl** Cl -> H C Cl + **H** **Cl**

H H

(b) Chloromethane + chlorine -> dichloromethane + Hydrogen chloride

CH3Cl (g) + Cl2(g) -> CH2Cl2 (g) + HCl (g)

H H

H C **Cl** + **Cl** Cl -> H C Cl + **H** **Cl**

**H** **Cl**

(c) dichloromethane + chlorine -> trichloromethane + Hydrogen chloride

CH2Cl2 (g) + Cl2(g) -> CHCl3 (g) + HCl (g)

Cl H

H C **Cl** + **Cl** Cl -> Cl C Cl + **H** **Cl**

**H** **Cl**

(c) trichloromethane + chlorine -> tetrachloromethane + Hydrogen chloride

CHCl3 (g) + Cl2(g) -> CCl4 (g) + HCl (g)

H Cl

**Cl** C **Cl** + **Cl** Cl -> Cl C Cl + **H** **Cl**

**Cl** **Cl**

Ethane has no effect on bromine or chlorine in diffused light/dark. In sunlight , a mixture of bromine and ethane explode to form colourless mixture of bromoethane and hydrogen chloride gas. The red/brown colour of bromine gas fades.

Chemical equation

(a)Ethane + chlorine -> Chloroethane + Hydrogen chloride

CH3CH3(g) + Br2(g) -> CH3CH2Br (g) + HBr (g)

H H H H

H C C H + **Br** **Br** -> H C C H + **H** **Br**

H **H** H **Br**

Bromoethane

H **H** H Br

H C C H + **Br** **Br** -> H C C H + **H** **Br**

H **Br** H **Br**

1,1-dibromoethane

H **Br** H Br

H C C H + **Br** **Br** -> H C C Br + **H** **Br**

H **Br** H **Br**

1,1,1-tribromoethane

H **Br** H Br

H C C **Br** + **Br** **Br** -> H C C Br + **H** **Br**

H **Br** **Br** **Br**

1,1,1,2-tetrabromoethane

H **Br** H Br

H C C **Br** + **Br** **Br** -> **Br** C C Br + **H** **Br**

**Br** **Br** **Br** **Br**

1,1,1,2,2-pentabromoethane

H **Br** **Br** Br

**Br** C C **Br** + **Br** **Br** -> **Br** C C Br + **H** **Br**

**Br** **Br** **Br** **Br**

1,1,1,2,2,2-hexabromoethane

**Uses of alkanes**

1.Most alkanes are used as fuel e.g. Methane is used as biogas in homes.Butane is used as the Laboratory gas.

2.On cracking ,alkanes are a major source of Hydrogen for the manufacture of ammonia/Haber process.

3.In manufacture of Carbon black which is a component in printers ink.

4.In manufacture of useful industrial chemicals like methanol, methanol, and chloromethane.

**(ii) Alkenes**

**(a)Nomenclature/Naming**

These are hydrocarbons with a general formula **CnH2n** and **C C** double bond as the functional group . **n** is the number of Carbon atoms in the molecule.

The carbon atoms are linked by at least one **double** bond to each other and single bonds to hydrogen atoms.

They include:

|  |  |  |  |
| --- | --- | --- | --- |
| n | General/  Molecular  formula | Structural formula | Name |
| 1 |  | Does not exist |  |
| 2 | C2H6 | H H  H C C H    CH2 CH2 | Eth**ene** |
| 3 | C3H8 | H H H  H C C C H  H  CH2 CH CH3 | Prop**ene** |
| 4 | C4H10 | H H H H  H C C C C H  H H  CH2 CH CH2CH3 | But**ene** |
| 5 | C5H12 | H H H H H  H C C C C C H  H H H  CH2 CH (CH2)2CH3 | Pent**ene** |
| 6 | C6H14 | H H H H H H  H C C C C C C H  H H H H  CH2 CH (CH2)3CH3 | Hex**ene** |
| 7 | C7H16 | H H H H H H H  H C C C C C C C H  H H H H H H H  CH2 CH (CH2)4CH3 | Hept**ene** |
| 8 | C8H18 | H H H H H H H H  H C C C C C C C C H  H H H H H H  CH2 CH (CH2)5CH3 | Oct**ene** |
| 9 | C9H20 | H H H H H H H H H  H C C C C C C C C C H  H H H H H H H  CH2 CH (CH2)6CH3 | Non**ene** |
| 10 | C10H22 | H H H H H H H H H H  H C C C C C C C C C C H  H H H H H H H H  CH2 CH (CH2)7CH3 | dec**ene** |

**Note**

1.Since carbon is **tetravalent** ,each atom of carbon in the alkene **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons including at the double bond.

2.Since Hydrogen is **monovalent** ,each atom of hydrogen in the alkene **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

3.One member of the alkene ,like alkanes,differ from the next/previous by a CH2  group.They also form a homologous series.

e.g

Propene differ from ethene by one carbon and two Hydrogen atoms from ethene. 4.A homologous series of alkenes like that of alkanes:

(i) differ by a CH2  group from the next /previous consecutively

(ii)have similar chemical properties

(iii)have similar chemical formula represented by the general formula CnH2n

(iv)the physical properties also show steady gradual change

5.The = C= C = double bond in alkene is the functional group. A functional group is the **reacting site** of a molecule/compound.

6. The = C= C = double bond in alkene can easily be broken to accommodate more two more monovalent atoms. The = C= C = double bond in alkenes make it thus **unsaturated.**

**7.** An unsaturated hydrocarbon is one with a double =C=C= or triple – C C – carbon bonds in their molecular structure. Unsaturated hydrocarbon easily reacts to be **saturated**.

8.A saturated hydrocarbon is one without a double =C=C= or triple – C C – carbon bonds in their molecular structure.

Most of the reactions of alkenes take place at the = C **=** C =bond.

**(b)Isomers of alkenes**

Isomers are alkenes lie alkanes have the same molecular **general formula** but different molecular **structural formula**.

Ethene and propene do not form isomers. Isomers of alkenes are also named by using the IUPAC(**I**nternational **U**nion of **P**ure and **A**pplied **C**hemistry) system of **nomenclature/naming.**

The IUPAC system of nomenclatureof naming alkenesuses the following basic rules/guidelines:

1.Identify the longest continuous/straight carbon chain which contains the **=C = C= double** bond get/determine the **parent** alkene.

2.Number the longest chain form the end of the chain which contains the **=C = C= double** bond so he **=C = C= double** bond lowest number possible.

3 Indicate the positions by splitting “**alk**-positions-**ene**” e.g. but-2-ene, pent-1,3-diene.

4.The position **indicated** must be for the carbon atom at the **lower** position in the **=C = C= double bond.**i.e

But-2-ene means the double =C = C= is between Carbon “2”and “3”

Pent-1,3-diene means there are two double bond one between carbon “1” and “2”and another between carbon “3” and “4”

5. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of alkyl carbon chains attached to the alkene. Name them fluoro-,chloro-,bromo-,iodo- if they are halogens

6.Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of **double** C = C bonds and **branches** attached to the alkene.

7.Position isomers can be formed when the=C = C= double bond is shifted between carbon atoms e.g.

But-2-ene means the double =C = C= is between Carbon “2”and “3”

But-1-ene means the double =C = C= is between Carbon “1”and “2”

Both But-1-ene and But-2-ene are position isomers of Butene

8.Position isomers are molecules/compounds having the same general formular but different position of the functional group.i.e.

Butene has the molecular/general formular C4H8 position but can form both But-1-ene and But-2-ene as position isomers.

9. Like alkanes ,an alkyl group can be attached to the alkene. Chain/branch isomers are thus formed.

10.Chain/branch isomers are molecules/compounds having the same general formula but different structural formula e.g

Butene and 2-methyl propene both have the same general formualr but different branching chain.

Practice on IUPAC nomenclature of alkenes

Name the following isomers of alkene

H H H H

H C C C C H But-1-ene

H H

H H H H

H C C C C H But-2-ene

H H

H H H H H H

H C C C C C C H 4-methylhex-1-ene

H H H

H C H

H

H

H C H

H H H H H

H C C C C C C H 4,4-dimethylhex-1-ene

H H H

H C H

H

3. H

H C H

H H H H

H C C C C C H 4,4-dimethylpent -1- ene

H H

H C H

H

4. H

H C H

H H H H

H C C C C C H 5,5-dimethylhex-2- ene

H C H H H

H C H

H

H

5. H

H C H

H H H

H C C C C H 2,2-dimethylbut -2- ene

H H

H C H

H

8.H2C CHCH2 CH2 CH3 pent -1- ene

9.H2C C(CH3)CH2 CH2 CH3 2-methylpent -1- ene

10.H2C C(CH3)C(CH3)2 CH2 CH3 2,3,3-trimethylpent -1- ene

11.H2C C(CH3)C(CH3)2 C(CH3)2 CH3 2,3,3,4,4-pentamethylpent -1- ene

12.H3C C(CH3)C(CH3) C(CH3)2 CH3 2,3,4,4-tetramethylpent -2- ene

13. H2C C(CH3)C(CH3) C(CH3) CH3 2,3,4-trimethylpent -1,3- diene

14. H2C CBrCBr CBr CH3 2,3,4-tribromopent -1,3- diene

15. H2C CHCH CH2 But -1,3- diene

16. Br2C CBrCBr CBr2 1,1,2,3,4,4-hexabromobut -1,3- diene

17. I2C CICI CI2 1,1,2,3,4,4-hexaiodobut -1,3- diene

18. H2C C(CH3)C(CH3) CH2 2,3-dimethylbut -1,3- diene

**(c)Occurrence and extraction**

At indusrial level,alkenes are obtained from the cracking of alk**a**nes.Cracking is the process of breaking long chain alkanes to smaller/shorter alk**a**nes, an alk**e**ne and hydrogen gas at high temperatures.

Cracking is a major source of useful hydrogen gas for manufacture of ammonia/nitric(V)acid/HCl i.e.

Long chain alkane -> smaller/shorter alk**a**ne + Alk**e**ne + Hydrogen gas

Examples

1.When irradiated with high energy radiation,Propane undergo cracking to form methane gas, ethene and hydrogen gas.

Chemical equation

CH3CH2CH3 (g) -> CH4(g) + CH2=CH2(g) + H2(g)

2.Octane undergo cracking to form hydrogen gas, butene and butane gases

Chemical equation

CH3(CH2) 6 CH3 (g) -> CH3CH2CH2CH3(g) + CH3 CH2CH=CH2(g) + H2(g)

**(d)School laboratory preparation of alk**e**nes**

In a school laboratory, alk**e**nes may be prepared from dehydration of alkanols using:

(i) concentrated sulphuric(VI)acid(H2SO4).

(a) aluminium(III)oxide(Al2O3) i.e

Alkanol --Conc. H2SO4 --> Alkene + Water

Alkanol --Al2O3 --> Alkene + Water e.g.

1.(a)At about 180oC,concentrated sulphuric(VI)acid dehydrates/removes water from ethanol to form eth**e**ne.

The gas produced contain traces of carbon(IV)oxide and sulphur(IV)oxide gas as impurities.

It is thus passed through concentrated sodium/potassium hydroxide solution to remove the impurities.

Chemical equation

CH3CH2**OH** (l) --conc H2SO4/180oC--> CH2=CH2(g) + **H2O**(l)

(b)On heating strongly aluminium(III)oxide(Al2O3),it dehydrates/removes water from ethanol to form eth**e**ne.

Ethanol vapour passes through the hot aluminium (III) oxide which catalyses the dehydration.

Activated aluminium(III)oxide has a very high affinity for water molecules/elements of water and thus dehydrates/ removes water from ethanol to form eth**e**ne.

Chemical equation

CH3CH2**OH** (l) --(Al2O3/strong heat--> CH2=CH2(g) + **H2O**(l)

2(a) Propan-1-ol and Propan-2-ol(position isomers of propanol) are dehydrated by conc H2SO4 at about 180oC to prop**e**ne(propene has no position isomers).

Chemical equation

CH3CH2 CH2**OH** (l) -- conc H2SO4/180oC --> CH3CH2=CH2(g) + **H2O**(l)

Propan-**1**-ol Prop-**1**-ene

CH3CH**OH** CH3 (l) -- conc H2SO4/180oC --> CH3CH2=CH2(g) + **H2O**(l)

Propan-**2**-ol Prop-**1**-ene

(b) Propan-1-ol and Propan-2-ol(position isomers of propanol) are dehydrated by heating strongly aluminium(III)oxide(Al2O3) form prop**e**ne

Chemical equation

CH3CH2 CH2**OH** (l) -- Heat/Al2O3 --> CH3CH2=CH2(g) + **H2O**(l)

Propan-**1**-ol Prop-**1**-ene

CH3CH**OH** CH3 (l) -- Heat/Al2O3 --> CH3CH2=CH2(g) + **H2O**(l)

Propan-**2**-ol Prop-**1**-ene

3(a) Butan-1-ol and Butan-2-ol(position isomers of butanol) are dehydrated by conc H2SO4 at about 180oC to But-1-**e**ne and But-2-**e**ne respectively

Chemical equation

CH3CH2 CH2 CH2**OH** (l) -- conc H2SO4/180oC -->CH3 CH2CH2=CH2(g) + **H2O**(l)

Butan-**1**-ol But-**1**-ene

CH3CH**OH** CH2CH3 (l)-- conc H2SO4/180oC -->CH3CH=CH CH2(g) + **H2O**(l)

Butan-**2**-ol But-**2**-ene

(b) Butan-1-ol and Butan-2-ol are dehydrated by heating strongly aluminium (III) oxide (Al2O3) form But-1-**e**ne and But-2-**e**ne respectively.

Chemical equation

CH3CH2 CH2 CH2**OH** (l) -- Heat/Al2O3 --> CH3 CH2CH2=CH2(g) + **H2O**(l)

Butan-**1**-ol But-**1**-ene

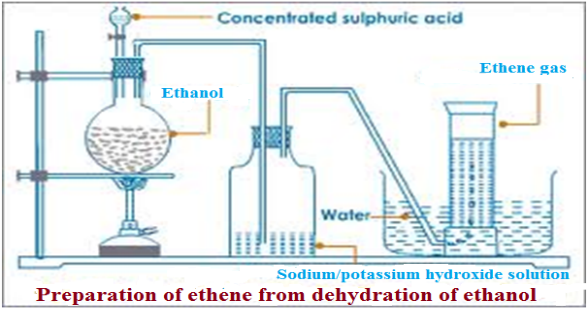
CH3CH**OH** CH2CH3 (l) -- Heat/Al2O3 --> CH3CH=CH CH2(g) + **H2O**(l)

Butan-**2**-ol But-**2**-ene

Laboratory set up for the preparation of alk**e**nes/**ethene**

**Caution:**Ethanol is highly inflammable**, and** Conc H2SO4 is highly corrosive on skin contact.

**(i)using conentrated sulphuric(VI)acid**



Some broken porcelain or sand should be put in the flask when heating to:

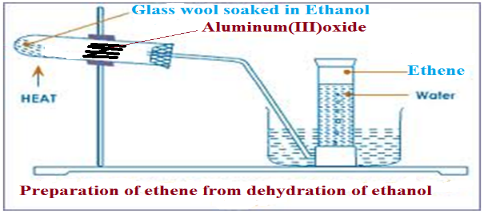
(i)prevent bumping which may break the flask.

(ii)ensure uniform and smooth boiling of the mixture

The temperatures should be maintained at above160oC.

At lower temperatures another compound -**ether** is predominantly formed instead of ethene gas**.**

**(ii)Using aluminium(III)oxide**



**(e)Properties of alk**e**nes**

**I. Physical properties**

Like alk**a**nes, alk**e**nes are colourles gases, solids and liquids that are not poisonous.

They are slightly soluble in water.

The solubility in water decrease as the carbon chain and as the molar mass increase but very soluble in organic solvents like tetrachloromethane and methylbenzene.

The melting and boiling point increase as the carbon chain increase.

This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase.

The 1st four straight chain alk**e**nes (eth**e**ne,propane,but-1-**e**ne and pent-1-**e**ne)are gases at room temperature and pressure.

The density of straight chain alk**e**nes,like alkanes, increase with increasing carbon chain as the intermolecular forces increases reducing the volume occupied by a given mass of the alk**e**ne.

Summary of physical properties of the 1st five alk**e**nes

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Alk**e**ne | General formula | Melting point(oC) | Boiling point(K) | State at room(298K) temperature and pressure atmosphere (101300Pa) |
| Eth**e**ne | CH2CH2 | -169 | -104 | gas |
| Prop**e**ne | CH3 CHCH2 | -145 | -47 | gas |
| But**e**ne | CH3CH2 CHCH2 | -141 | -26 | gas |
| Pent-1-**e**ne | CH3(CH2 CHCH2 | -138 | 30 | liquid |
| Hex-1-**e**ne | CH3(CH2) CHCH2 | -98 | 64 | liquid |

**II. Chemical properties**

**(a)Burning/combustion**

Alk**e**nes burn with a **yellow**/ luminous **sooty**/ smoky flame in **excess** air to form carbon(IV) oxide and water.

Alk**e**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

Alk**e**nes burn with a **yellow**/ luminous **sooty**/ smoky flame in **limited** air to form carbon(II) oxide and water.

Alk**e**ne + Air -> carbon(**II**) oxide + water (limited air)

Burning of alkenes with a **yellow**/ luminous **sooty**/ smoky flame is a confirmatory test for the **presence** of the  **=C=C=** double bond because they have **higher C:H ratio.**

A homologous series with C = C double or C C triple bond is said to be **unsaturated.**

A homologous series with C C single bond is said to be **saturated**.Most of the reactions of the unsaturated compound involve trying to be saturated to form a

C C single bond .

Examples of burning alkenes

1.(a) Eth**e**ne when ignited burns with a **yellow** **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Eth**e**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

C2H4(g) + 3O2(g) -> 2CO2(g) + 2H2O(l/g)

(b) Eth**e**ne when ignited burns with a **yellow** **sooty** flame in **limited** air to form carbon(II) oxide and water.

Eth**e**ne + Air -> carbon(**II**) oxide + water (limited air )

C2H4(g) + 3O2(g) -> 2CO2(g) + 2H2O(l/g)

2.(a) Prop**e**ne when ignited burns with a **yellow** **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Prop**e**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

2C3H6(g) + 9O2(g) -> 6CO2(g) + 6H2O(l/g)

(a) Prop**e**ne when ignited burns with a **yellow** **sooty** flame in **limited** air to form carbon(II) oxide and water.

Prop**e**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

C3H6(g) + 3O2(g) -> 3CO(g) + 3H2O(l/g)

**(b)Addition reactions**

An addition reaction is one which an unsaturated compound reacts to form a saturated compound.Addition reactions of alkenes are named from the reagent used to cause the addtion/convert the double =C=C= to single C-C bond.

**(i)Hydrogenation**

Hydrogenation is an addition reaction in which **hydrogen** in presence of **Palladium/Nickel** catalyst at high temperatures react with alk**e**nes to form alk**a**nes.

Examples

1.When Hydrogen gas is passed through liquid vegetable and animal **oil** at about 180oC in presence of Nickel catalyst,solid **fat** is formed.

Hydrogenation is thus used to **harden** oils to solid fat especially margarine.

During hydrogenation, one hydrogen atom in the hydrogen molecule attach itself to one carbon and the other hydrogen to the second carbon breaking the double bond to single bond.

Chemical equation

H2C=CH2 + H2  -Ni/Pa-> H3C - CH3

H H H H

C = C + H – H - Ni/Pa -> H - C – C - H

H H H H

2.Prop**e**ne undergo hydrogenation to form Propane

Chemical equation

H3C CH=CH2 + H2  -Ni/Pa-> H3C CH - CH3

H H H H H H

H C C = C + H – H - Ni/Pa-> H - C – C - C- H

H H H H H

3.Both But-1-**e**ne and But-2-**e**ne undergo hydrogenation to form Butane

Chemical equation

But-1-**e**ne + Hydrogen –Ni/Pa-> Butane

H3C CH2 CH=CH2 + H2  -Ni/Pa-> H3C CH2CH - CH3

H H H H H H H H

H C C - C = C + H – H - Ni/Pa-> H - C- C – C - C- H

H H H H H H H

But-2-**e**ne + Hydrogen –Ni/Pa-> Butane

H3C CH2 =CH CH2 + H2  -Ni/Pa-> H3C CH2CH - CH3

H H H H H H H H

H C C = C - C -H + H – H - Ni/Pa-> H - C- C – C - C- H

H H H H H H

4. But-1,3-di**e**ne should undergo hydrogenation to form Butane. The reaction uses **two** moles of hydrogen molecules/**four** hydrogen atoms to break the two double bonds.

But-1,3-di**e**ne + Hydrogen –Ni/Pa-> Butane

H2C CH CH=CH2 + **2**H2  -Ni/Pa-> H3C CH2CH - CH3

H H H H H H H H

H C C - C = C -H + **2**(H – H) - Ni/Pa-> H - C- C – C - C- H

H H H H

**(ii) Halogenation.**

Halogenation is an addition reaction in which a halogen (Fluorine, chlorine, bromine, iodine) reacts with an alk**e**ne to form an alkane.

The double bond in the alk**e**ne break and form a single bond.

The colour of the halogen **fades** as the number of moles of the halogens remaining unreacted decreases/reduces.

One bromine atom bond at the 1st carbon in the double bond while the other goes to the 2nd carbon.

Examples

1Ethene reacts with bromine to form 1,2-dibromoethane.

Chemical equation

H2C=CH2 + Br2  H2 Br C - CH2 Br

H H H H

C = C + Br – Br Br - C – C - Br

H H H H

Ethene + Bromine 1,2-dibromoethane

2.Propene reacts with chlorine to form 1,2-dichloropropane.

Chemical equation

H3C CH=CH2 + Cl2  H3C CHCl - CH2Cl

Propene + Chlorine 1,2-dichloropropane

H H H H H H

H C C = C + Cl – Cl H - C – C - C- Cl

H H H Cl H

H H H H H H H H

H C C - C = C + I – I H - C- C – C - C- I

H H H H H H H H

3.Both But-1-**e**ne and But-2-**e**ne undergo halogenation with iodine to form 1,2-diiodobutane and 2,3-diiodobutane

Chemical equation

But-1-**e**ne + iodine 1,2 diiodobutane

H3C CH2 CH=CH2 + I2  H3C CH2CH I - CH2I

But-2-**e**ne + Iodine 2,3-diiodobutane

H3C CH= CH-CH2 + F2  H3C CHICHI - CH3

H H H H H H H H

H C C = C - C -H + I – I H - C- C – C - C- H

H H H I I H

4. But-1,3-di**e**ne should undergo halogenation to form Butane. The reaction uses **two** moles of iodine molecules/**four** iodine atoms to break the two double bonds.

But-1,3-di**e**ne + iodine 1,2,3,4-tetraiodobutane

H2C= CH CH=CH2 + **2**I2  H2CI CHICHI - CHI

H H H H H H H H

H C C - C = C -H + **2**(I – I) H - C- C – C - C- H

I I I I

**(iii) Reaction with hydrogen halides.**

**Hydrogen halides** reacts with alk**e**ne to form a halogenoalkane. The double bond in the alk**e**ne break and form a single bond.

The main compound is one which the **hydrogen** atom bond at the carbon with **more** **hydrogen** .

Examples

1. Ethene reacts with hydrogen bromide to form bromoethane.

Chemical equation

H2C=CH2 + HBr H3 C - CH2 Br

H H H H

C = C + H – Br H - C – C - Br

H H H H

Ethene + Bromine bromoethane

2. Propene reacts with hydrogen iodide to form 2-iodopropane.

Chemical equation

H3C CH=CH2 + HI H3C CHI - CH3

Propene + Chlorine 2-chloropropane

Carbon atom with more Hydrogen atoms gets **extra** hydrogen

H H H H H H

H C C = **C** + H – Cl H - C – C - **C**- H

H H H Cl H

3. Both But-1-**e**ne and But-2-**e**ne reacts with hydrogen bromide to form 2- bromobutane

Chemical equation

But-1-**e**ne + hydrogen bromide 2-bromobutane

H3C CH2 CH=CH2 + HBr H3C CH2CHBr -CH3

H H H H H H H H

H C C - C = C + H – Br H - C- C – C - C- H

H H H H H Br H

But-2-**e**ne + Hydrogen bromide 2-bromobutane

H3C CH= CH-CH2 + HBr H3C CHBrCH2 - CH3

H H H H H H H H

H C C = C - C -H + Br – H H - C- C – C - C- H

H H H Br H H

4. But-1,3-di**e**ne react with hydrogen iodide to form 2,3- diiodobutane. The reaction uses **two** moles of hydrogen iodide molecules/**two** iodine atoms and two hydrogen atoms to break the two double bonds.

But-1,3-di**e**ne + iodine 2,3-diiodobutane

H2C= CH CH=CH2 + **2H**I2  H3CCHICHI - CH3

H H H H H H H H

H C C - C = C -H + **2**(H – I) H - C- C – C - C- H

H I I H

**(iv) Reaction with bromine/chlorine water.**

Chlorine and bromine water is formed when the halogen is dissolved in distilled water.Chlorine water has the formular HOCl(hypochlorous/chloric(I)acid) .Bromine water has the formular HOBr(hydrobromic(I)acid).

During the addition reaction .the halogen move to one carbon and the OH to the other carbon in the alkene at the =C=C= double bond to form a **halogenoalkanol.**

Bromine water + Alkene -> bromoalkan**o**l

Chlorine water + Alkene -> bromoalkan**o**l

Examples

1Ethene reacts with bromine water to form bromoethan**ol**.

Chemical equation

H2C=CH2 + HOBr H2 Br C - CH2 OH

H H H H

C = C + Br – OH Br - C – C - OH

H H H H

Ethene + Bromine water bromoethan**ol**

2.Propene reacts with chlorine water to form chloropropan-2-ol / 2-chloropropan-1-ol.

Chemical equation

**I.**H3C CH=CH2 + HOCl H3C CHCl - CH2OH

Propene + Chlorine water 2-chloropropane

H H H H H H

H C C = C + HO – Cl H - C – C - C- OH

H H H Cl H

**II.**H3C CH=CH2 + HOCl H3C CHOH - CH2Cl

Propene + Chlorine chloropropan-2-ol

H H H H H H

H C C = C + HO – Cl H - C – C - C- Cl

H H H OH H

3.Both But-1-**e**ne and But-2-**e**ne react with bromine water to form 2-bromobutan-1-**ol** /3-bromobutan-2-**ol** respectively

Chemical equation

**I.**But-1-**e**ne + bromine water 2-bromobutan-1-ol

H3C CH2 CH=CH2 + HOBr H3C CH2CH Br - CH2OH

H H H H H H H H

H C C - C = C + HO– Br H - C- C – C - C- OH

H H H H H Br H

**II.**But-2-**e**ne + bromine water 3-bromobutan-2-ol

H3C CH= CHCH3 + HOBr H3C CH2OHCH Br CH3

H H H H H H H H

H C C - C = C + HO– Br H - C- C – C - C- OH

H H H H H Br H

4. But-1,3-di**e**ne reacts with bromine water to form Butan-1,3-diol.

The reaction uses **two** moles of bromine water molecules to break the two double bonds.

But-1,3-di**e**ne + bromine water 2,4-dibromobutan-1,3-diol

H2C= CH CH=CH2 + **2**HOBr H2COH CHBrCHOH CHBr

H H H H H H H H

H C C - C = C -H + **2**(HO – Br) H - C- C – C - C- H

HO Br HO Br

**(v) Oxidation.**

Alkenes are oxidized to alkanols with **duo/double** functional groups by oxidizing agents.

When an alk**e**ne is bubbled into orange acidified potassium/sodium dichromate (VI) solution,the colour of the oxidizing agent changes to green.

When an alk**e**ne is bubbled into purple acidified potassium/sodium manganate(VII) solution, the oxidizing agent is decolorized.

Examples

1Ethene is oxidized to ethan-1,2-di**ol** by acidified potassium/sodium manganate(VII) solution/ acidified potassium/sodium dichromate(VI) solution.

The purple acidified potassium/sodium manganate(VII) solution is decolorized.

The orange acidified potassium/sodium dichromate(VI) solution turns to green.

Chemical equation

H2C=CH2  **[O]** in H+/K2Cr2O7  HO CH2 - CH2 OH

H H H H

C = C+ [O] in H+/KMnO4 H - C – C - H

H H OH OH

Ethene + [O] in H+/KMnO4 ethan-1,2-di**ol**

2. Propene is oxidized to propan-1,2-di**ol** by acidified potassium/sodium manganate(VII) solution/ acidified potassium/sodium dichromate(VI) solution.

The purple acidified potassium/sodium manganate(VII) solution is decolorized.

The orange acidified potassium/sodium dichromate(VI) solution turns to green.

Chemical equation

H3C CH=CH2  **[O]** in H+/KMnO4   H3C CHOH - CH2OH

Propene  **[O]** in H+/KMnO4   propan-1,2-diol

H H H H H H

H C C = C **[O]** in H+/KMnO4  H - C – C - C- OH

H H H OH H

3.Both But-1-**e**ne and But-2-**e**ne react with bromine water to form butan-1,2-diol and butan-2,3-diol

Chemical equation

**I.**But-1-**e**ne + **[O]** in H+/KMnO4 butan-1,2-diol

H3C CH2 CH=CH2 + [O] H3C CH2CHOH - CH2OH

H H H H H H H H

H C C - C = C + [O] H - C- C – C - C- OH

H H H H H OH H

**(v) Hydrolysis.**

Hydrolysis is the reaction of a compound with water/addition of H-OH to a compound.

Alkenes undergo hydrolysis to form alkanols .

This takes place in two steps:

(i)Alkenes react with **concentrated sulphuric(VI)acid** at room temperature and pressure to form **alkylhydrogen sulphate(VI)**.

Alkenes + concentrated sulphuric(VI)acid -> alkylhydrogen sulphate(VI)

(ii)On adding **water** to alkylhydrogen sulphate(VI) then warming, an alkanol is formed.

alkylhydrogen sulphate(VI) + water -warm-> Alkanol.

Examples

(i)Ethene reacts with cold concentrated sulphuric(VI)acid to form ethyl hydrogen sulphate(VII)

Chemical equation

H2C=CH2  **+** H2SO4  CH3 - CH2OSO3H

H H H O-SO3H

C = C  **+** H2SO4  H - C – C - H

H H H H

Ethene  **+** H2SO4  ethylhydrogen sulphate(VI)

(ii) Ethylhydrogen sulphate(VI) is hydrolysed by water to ethanol

Chemical equation

CH3 - CH2OSO3H  **+** H2O CH3 - CH2OH + H2SO4

H OSO3H H OH

H - C - C - H  **+** H2O H - C – C - H + H2SO4

H H H H

ethylhydrogen sulphate(VI) **+** H2O Ethanol

2. Propene reacts with cold concentrated sulphuric(VI)acid to form propyl hydrogen sulphate(VII)

Chemical equation

CH3H2C=CH2  **+** H2SO4  CH3CH2 - CH2OSO3H

H H H H H O-SO3H

C = C - C - H **+** H2SO4  H - C - C – C - H

H H H H H H

Propene  **+** H2SO4  propylhydrogen sulphate(VI)

(ii) Propylhydrogen sulphate(VI) is hydrolysed by water to propanol

Chemical equation

CH3 - CH2OSO3H  **+** H2O CH3 - CH2OH + H2SO4

H H OSO3H H H OH

H - C - C - C - H  **+** H2O H - C - C – C - H + H2SO4

H H H H H H

propylhydrogen sulphate(VI) **+** H2O propanol

**(vi) Polymerization/self addition**

Addition polymerization is the process where a small unsaturated monomer (alkene ) molecule join together to form a large saturated molecule.

**Only alkenes undergo addition polymerization.**

Addition polymers are named from the alkene/monomer making the polymer and adding the prefix “**poly**” before the name of monomer to form a **polyalkene**

During addition polymerization

(i)the double bond in alkenes break

(ii)free radicals are formed

(iii)the free radicals collide with each other and join to form a larger molecule. The more collisions the larger the molecule.

**Examples of addition polymerization**

1.Formation of Polyethene

Polyethene is an addition polymer formed when ethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting paticles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H H H H H H H H

Ethene + Ethene + Ethene + Ethene + **…**

(ii)the double bond joining the ethane molecule break to free readicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H H H H H H H H

Ethene radical + Ethene radical + Ethene radical + Ethene radical + **…**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

•C – C - C – C - C – C - C - C• + **…**

H H H H H H H H

Lone pair of electrons can be used to join more monomers to form longer polyethene.

Polyethene molecule can be represented as:

H H H H H H H H extension of

molecule/polymer

- C – C - C – C - C – C - C – C-

H H H H H H H H

Since the molecule is a **repetition** of one monomer, then the polymer is:

H H

( C – C )**n**

H H

Where **n** is the number of monomers in the polymer. The number of monomers in the polymer can be determined from the molar mass of the polymer and monomer from the relationship:

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

Examples

**Polythene has a molar mass of 4760.Calculate the number of ethene molecules in the polymer(C=12.0, H=1.0 )**

Number of monomers/repeating units in polyomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C2H4 )= 28 Molar mass polyethene = 4760

Substituting 4760 = 170 ethene molecules

28

The **commercial** name of polyethene is **polythene**.

It is an elastic, tough, transparent and durable plastic.

Polythene is used:

(i)in making plastic bag

(ii)bowls and plastic bags

(iii)packaging materials

2.Formation of Polychlorethene

Polychloroethene is an addition polymer formed when chloroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

chloroethene + chloroethene + chloroethene + chloroethene + **…**

(ii)the double bond joining the chloroethene molecule break to free radicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

•C – C - C – C - C – C - C - C• + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

Lone pair of electrons can be used to join more monomers to form longer polychloroethene.

Polychloroethene molecule can be represented as:

H H H H H H H H extension of

molecule/polymer

- C – C - C – C - C – C - C – C- + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

Since the molecule is a repetition of one monomer, then the polymer is:

H H

( C – C )**n**

H **Cl**

Examples

**Polychlorothene has a molar mass of 4760.Calculate the number of chlorethene molecules in the polymer(C=12.0, H=1.0,Cl=35.5 )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C2H3Cl )= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**

62.5

The **commercial** name of polychloroethene is **polyvinylchloride(PVC)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope

(ii)water pipes

(iii)crates and boxes

3.Formation of Polyphenylethene

Polyphenylethene is an addition polymer formed when phenylethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H **C6H5** H **C6H5** H **C6H5** H **C6H5**

phenylethene + phenylethene + phenylethene + phenylethene + **…**

(ii)the double bond joining the phenylethene molecule break to free radicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H **C6H5** H **C6H5** H **C6H5** H **C6H5**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

• C – C - C – C - C – C - C - C • + **…**

H **C6H5**  H **C6H5** H **C6H5** H **C6H5**

Lone pair of electrons can be used to join more monomers to form longer polyphenylethene.

Polyphenylethene molecule can be represented as:

H H H H H H H H

- C – C - C – C - C – C - C - C -

H **C6H5**  H **C6H5** H **C6H5** H **C6H5**

Since the molecule is a repetition of one monomer, then the polymer is:

H H

( C – C )**n**

H **C6H5**

Examples

**Polyphenylthene has a molar mass of 4760.Calculate the number of phenylethene molecules in the polymer(C=12.0, H=1.0, )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C8H8 )= 104 Molar mass polyethene = 4760

Substituting 4760 = 45.7692 =>45 polyphenylethene molecules**(whole number)**

104

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

(iii)clothe linings

4.Formation of Polypropene

Polypropene is an addition polymer formed when propene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H **CH3** H **CH3** H **CH3** H **CH3**

propene + propene + propene + propene + **…**

(ii)the double bond joining the phenylethene molecule break to free radicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H **CH3** H **CH3** H **CH3** H **CH3**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

• C – C - C – C - C – C - C - C • + **…**

H **CH3**  H **CH3** H **CH3** H **CH3**

Lone pair of electrons can be used to join more monomers to form longer propene.

propene molecule can be represented as:

H H H H H H H H

- C – C - C – C - C – C - C - C -

H **CH3**  H **CH3** H **CH3** H **CH3**

Since the molecule is a repetition of one monomer, then the polymer is:

H H

( C – C )**n**

H **CH3**

Examples

**Polypropene has a molar mass of 4760.Calculate the number of propene molecules in the polymer(C=12.0, H=1.0, )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass propene (C3H8 )= 44 Molar mass polyethene = 4760

Substituting 4760 = 108.1818 =>108 propene molecules**(whole number)**

44

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

(iii)clothe linings

5.Formation of Polytetrafluorothene

Polytetrafluorothene is an addition polymer formed when tetrafluoroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

F F F F F F F F

C = C + C = C + C = C + C = C + **…**

F F F F F F F F

tetrafluoroethene + tetrafluoroethene+ tetrafluoroethene+ tetrafluoroethene + **…**

(ii)the double bond joining the tetrafluoroethene molecule break to free radicals

F F F F F F F F

•C – C• + •C - C• + •C - C• + •C - C• + **…**

F F F F F F F F

(iii)the free radicals collide with each other and join to form a larger molecule

F F F F F F F F lone pair of electrons

•C – C - C – C - C – C - C - C• + **…**

F F F FF F F F

Lone pair of electrons can be used to join more monomers to form longer polytetrafluoroethene.

polytetrafluoroethene molecule can be represented as:

F F F F F F F F extension of

molecule/polymer

- C – C - C – C - C – C - C – C- + **…**

F F F F F F F F

Since the molecule is a repetition of one monomer, then the polymer is:

F F

( C – C )**n**

F F

Examples

**Polytetrafluorothene has a molar mass of 4760.Calculate the number of tetrafluoroethene molecules in the polymer(C=12.0, ,F=19 )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C2F4 )= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**

62.5

The **commercial** name of polytetrafluorethene(**P.T.F.E**) is **Teflon(P.T.F.E)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope

(ii)water pipes

(iii)crates and boxes

6.Formation of rubber from Latex

Natural rubber is obtained from rubber trees.

During harvesting an incision is made on the rubber tree to produce a milky white substance called **latex.**

Latex is a mixture of rubber and lots of water.

The latex is then added an acid to coagulate the rubber.

Natural rubber is a polymer of 2-methylbut-1,3-diene ;

H CH3 H H

CH2=C (CH3) CH = CH2 H - C  **=** C – C **=** C - H

During natural polymerization to rubber, one double C=C bond break to self add to another molecule.The double bond remaining move to carbon “2” thus;

H CH3 H H H CH3 H H

- C - C  **=** C  **-** C - C - C **=** C **-** C -

H H H H

Generally the structure of rubber is thus;

H CH3 H H

-(- C - C **=** C - C -)**n**-

H H

Pure rubber is soft and sticky.It is used to make erasers, car tyres. Most of it is vulcanized.Vulcanization is the process of heating rubber with sulphur to make it harder/tougher.

During vulcanization the sulphur atoms form a cross link between chains of rubber molecules/polymers. This decreases the number of C=C double bonds in the polymer.

H CH3 H H H CH3 H H

Sulphur atoms make cross link between polymers

- C - C  **-** C  **-** C - C - C **-** C **-** C -

H S H H S H

H CH3 S H H CH3 S H

- C - C  **-**  C  **-** C - C - C **-** C **-** C -

H H H H H H

Vulcanized rubber is used to make **tyres**, **shoes** and **valves**.

7.Formation of synthetic rubber

Synthetic rubber is able to resist action of oil,abrasion and organic solvents which rubber cannot.

Common synthetic rubber is a polymer of 2-chlorobut-1,3-diene ;

H Cl H H

CH2=C (Cl CH = CH2 H - C  **=** C – C **=** C - H

During polymerization to synthetic rubber, one double C=C bond is broken to self add to another molecule. The double bond remaining move to carbon “2” thus;

H Cl H H H Cl H H

- C - C  **=** C  **-** C - C - C **=** C **-** C -

H H H H

Generally the structure of rubber is thus;

H Cl H H

-(- C - C **=** C - C -)**n**-

H H

Rubber is thus strengthened through vulcanization and manufacture of synthetic rubber.

**(c)Test for the presence of – C = C – double bond.**

**(i)Burning/combustion**

All unsaturated hydrocarbons with a  **– C = C –** or  **– C = C –** bond burn with a yellow sooty flame.

Experiment

Scoop a sample of the substance provided in a clean metallic spatula. Introduce it on a Bunsen burner.

|  |  |
| --- | --- |
| Observation | Inference |
| Solid melt then burns with a yellow sooty flame | **– C = C –,**  **– C = C –** bond |

**(ii)Oxidation by acidified KMnO4/K2Cr2O7**

Bromine water ,Chlorine water and Oxidizing agentsacidified KMnO4/K2Cr2O7change to **unique** colour in presence of  **– C = C –**

or  **– C = C –** bond.

Experiment

Scoop a sample of the substance provided into a clean test tube. Add 10cm3 of distilled water. Shake. Take a portion of the solution mixture. Add three drops of acidified KMnO4/K2Cr2O7 .

|  |  |
| --- | --- |
| Observation | Inference |
| Acidified KMnO4 decolorized  Orange colour of acidified K2Cr2O7turns green  Bromine water is decolorized  Chlorine water is decolorized | **– C = C –**  **– C = C –** bond |

**(d)Some uses of Alkenes**

1. In the manufacture of plastic

2. Hydrolysis of eth**e**ne is used in industrial manufacture of ethanol.

3. In ripening of fruits.

4. In the manufacture of detergents.

**(iii) Alkynes**

**(a)Nomenclature/Naming**

These are hydrocarbons with a general formula **CnH2n**-**2** and **C C** double bond as the functional group . **n** is the number of Carbon atoms in the molecule.

The carbon atoms are linked by at least one **triple** bond to each other and single bonds to hydrogen atoms.

They include:

|  |  |  |  |
| --- | --- | --- | --- |
| n | General/  Molecular  formula | Structural formula | Name |
| 1 |  | Does not exist | - |
| 2 | C2H2 | H C C H  CH CH | Eth**y**ne |
| 3 | C3H4 | H  H C C C H  H  CH C CH3 | Prop**y**ne |
| 4 | C4H6 | H H  H C C C C H  H H  CH C CH2CH3 | But**y**ne |
| 5 | C5H8 | H H H  H C C C C C H  H H H  CH C (CH2)2CH3 | Pent**y**ne |
| 6 | C6H10 | H H H H  H C C C C C C H  H H H H  CH C (CH2)3CH3 | Hex**y**ne |

|  |  |  |  |
| --- | --- | --- | --- |
| 7 | C7H12 | H H H H H  H C C C C C C C H  H H H H H H H  CH C (CH2)4CH3 | Hept**y**ne |
| 8 | C8H14 | H H H H H H  H C C C C C C C C H  H H H H H H  CH C (CH2)5CH3 | Oct**y**ne |
| 9 | C9H16 | H H H H H H H  H C C C C C C C C C H  H H H H H H H  CH C (CH2)6CH3 | Non**y**ne |
| 10 | C10H18 | H H H H H H H H  H C C C C C C C C C C H  H H H H H H H H  CH C (CH2)7CH3 | Dec**y**ne |

**Note**

1. Since carbon is **tetravalent** ,each atom of carbon in the alk**y**ne **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons including at the triple bond.

2. Since Hydrogen is **monovalent** ,each atom of hydrogen in the alk**y**ne **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

3. One member of the alk**y**ne ,like alk**e**nes and alk**a**nes, differ from the next/previous by a CH2  group(molar mass of 14 **a**tomic **m**ass **u**nits).They thus form a homologous series.

e.g

Prop**y**ne differ from eth**y**ne by (14 a.m.u) one carbon and two Hydrogen atoms from eth**y**ne.

4.A homologous series of alkenes like that of alkanes:

(i) differ by a CH2  group from the next /previous consecutively

(ii) have similar chemical properties

(iii)have similar chemical formula with general formula CnH2n-2

(iv)the physical properties also show steady gradual change

5.The - C = C - triple bond in alk**y**ne is the functional group. The functional group is the **reacting site** of the alk**y**nes.

6. The - C = C - triple bond in alk**y**ne can easily be broken to accommodate more /four more monovalent atoms. The - C = C - triple bond in alk**y**nes make it thus **unsaturated** like alkenes**.**

7. Most of the reactions of alk**y**nes like alkenes take place at the - C **=** C- triple bond.

**(b)Isomers of alkynes**

Isomers of alk**y**nes have the same molecular **general formula** but different molecular **structural formula**.

Isomers of alk**y**nes are also named by using the IUPAC(**I**nternational **U**nion of **P**ure and **A**pplied **C**hemistry) system of **nomenclature/naming.**

The IUPAC system of nomenclatureof naming alk**y**nesuses the following basic rules/guidelines:

1.Identify the longest continuous/straight carbon chain which contains the  **- C = C- triple** bond to get/determine the **parent** alkene.

2. Number the longest chain form the end of the chain which contains the **-C = C- triple** bond so as **- C = C- triple** bond get lowest number possible.

3 Indicate the positions by splitting “**alk**-positions-**yne**” e.g. but-2-**y**ne, pent-1,3-di**y**ne.

4.The position **indicated** must be for the carbon atom at the **lower** position in the

**-C = C- triple bond.** i.e

But-2-**y**ne means the triple -C = C- is between Carbon “2”and “3”

Pent-1,3-di**y**ne means there are two triple bonds; one between carbon “1” and “2”and another between carbon “3” and “4”

5. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of alkyl carbon chains attached to the alk**y**ne. Name them fluoro-,chloro-,bromo-,iodo- if they are halogens

6.Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of **triple -** C = C- bonds and **branches** attached to the alk**y**ne.

7.Position isomers can be formed when the - C = C- triple bond is shifted between carbon atoms e.g.

But-2-**y**ne means the double - C = C- is between Carbon “2”and “3”

But-1-**y**ne means the double - C = C- is between Carbon “1”and “2”

Both But-1-**y**ne and But-2-yne are position isomers of But**y**ne.

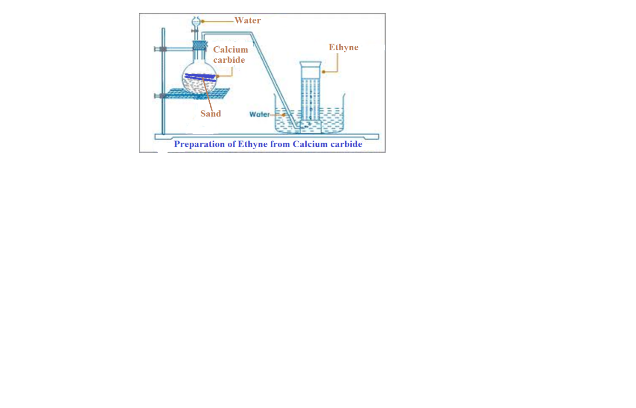
9. Like alkanes and alkynes , an alkyl group can be attached to the alk**y**ne. Chain/branch isomers are thus formed.

But**y**ne and 2-methyl prop**y**ne both have the same general formular but different branching chain.

(**More on powerpoint**)

**(c)Preparation of Alkynes.**

Ethyne is prepared from the reaction of water on calcium carbide. The reaction is highly exothermic and thus a layer of sand should be put above the calcium carbide to absorb excess heat to prevent the reaction flask from breaking. Copper(II)sulphate(VI) is used to catalyze the reaction



**(d)Properties of alkynes**

**I. Physical properties**

Like alk**a**nes and alk**e**nes,alkynes are colourles gases, solids and liquids that are not poisonous.

They are slightly soluble in water. The solubility in water decrease as the carbon chain and as the molar mass increase but very soluble in organic solvents like tetrachloromethane and methylbenzene. Ethyne has a pleasant taste when pure.

The melting and boiling point increase as the carbon chain increase.

This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase. The 1st three straight chain alk**y**nes (eth**y**ne,prop**y**ne and but-1-**y**ne)are gases at room temperature and pressure.

The density of straight chain alk**y**nes increase with increasing carbon chain as the intermolecular forces increases reducing the volume occupied by a given mass of the alk**y**ne.

Summary of physical properties of the 1st five alk**e**nes

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Alk**y**ne | General formula | Melting point(oC) | Boiling point(oC) | State at room(298K) temperature and pressure atmosphere (101300Pa) |
| Eth**y**ne | CH CH | -82 | -84 | gas |
| Prop**y**ne | CH3 C CH | -103 | -23 | gas |
| But**y**ne | CH3CH2 CCH | -122 | 8 | gas |
| Pent-1-**y**ne | CH3(CH2) 2 CCH | -119 | 39 | liquid |
| Hex-1-**y**ne | CH3(CH2) 3C CH | -132 | 71 | liquid |

**II. Chemical properties**

**(a)Burning/combustion**

Alk**y**nes burn with a **yellow**/ luminous very **sooty**/ smoky flame in **excess** air to form carbon(IV) oxide and water.

Alk**y**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

Alk**e**nes burn with a **yellow**/ luminous very**sooty**/ smoky flame in **limited** air to form carbon(II) oxide/carbon and water.

Alk**y**ne + Air -> carbon(**II**) oxide /carbon + water (limited air)

Burning of alk**y**nes with a **yellow**/ luminous **sooty**/ smoky flame is a confirmatory test for the **presence** of the  **- C = C –** triple bond because they have very **high C:H ratio.**

Examples of burning alk**y**nes

1.(a) Eth**y**ne when ignited burns with a **yellow** very **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Eth**y**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

2C2H2(g) + 5O2(g) -> 4CO2(g) + 2H2O(l/g)

(b) Eth**y**ne when ignited burns with a **yellow** **sooty** flame in **limited** air to form a mixture of unburnt carbon and carbon(II) oxide and water.

Eth**y**ne + Air -> carbon(**II**) oxide + water (limited air )

C2H2(g) + O2(g) -> 2CO2(g) + C + 2H2O(l/g)

2.(a) Prop**y**ne when ignited burns with a **yellow** **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Prop**y**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

C3H4(g) + 4O2(g) -> 3CO2(g) + 2H2O(l/g)

(a) Prop**y**ne when ignited burns with a **yellow** **sooty** flame in **limited** air to form carbon(II) oxide and water.

Prop**e**ne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

2C3H4(g) + 5O2(g) -> 6CO(g) + 4H2O(l/g)

**(b)Addition reactions**

An addition reaction is one which an unsaturated compound reacts to form a saturated compound. Addition reactions of alk**y**nes are also named from the reagent used to cause the addition/convert the triple - C = C- to single C- C bond.

**(i)Hydrogenation**

Hydrogenation is an addition reaction in which **hydrogen** in presence of **Palladium/Nickel** catalyst at 150oC temperatures react with alk**y**nes to form alk**e**nes then alk**a**nes.

Examples

1.During hydrogenation, **two** hydrogen atom in the hydrogen molecule attach itself to one carbon and the other **two** hydrogen to the second carbon breaking the **triple** bond to **double** the **single**.

Chemical equation

HC = CH + H2  -Ni/Pa -> H2C = CH2 + H2  -Ni/Pa -> H2C - CH2

H H H H H H

C = C + H – H - Ni/Pa -> H - C = C – H + H – H - Ni/Pa -> H - C - C – H

H H H H H H

2.Prop**y**ne undergo hydrogenation to form Propane

Chemical equation

H3C CH = CH2 + **2**H2  -Ni/Pa-> H3C CH - CH3

H H H H H H

H C C = C + **2**H – H - Ni/Pa-> H - C – C - C- H

H H H H H

3(a) But-1-**y**ne undergo hydrogenation to form Butane

Chemical equation

But-1-**y**ne + Hydrogen –Ni/Pa-> Butane

H3C CH2 C = CH + **2**H2  -Ni/Pa-> H3C CH2CH - CH3

H H H H H H H

H C C - C = C + **2**H – H - Ni/Pa-> H - C- C – C - C- H

H H H H H H

(b) But-2-**y**ne undergo hydrogenation to form Butane

Chemical equation

But-2-**y**ne + Hydrogen –Ni/Pa-> Butane

H3C C = C CH2 + **2**H2  -Ni/Pa-> H3C CH2CH - CH3

H H H H H H

H C C = C - C H + **2**H – H- Ni/Pa-> H - C- C – C - C- H

H H H H H H

**(ii) Halogenation.**

Halogenation is an addition reaction in which a halogen (Fluorine, chlorine, bromine, iodine) reacts with an alk**y**ne to form an alkene then alkane.

The reaction of alkynes with halogens with alkynes is **faster** than with alkenes. The triple bond in the alk**y**ne break and form a double then single bond.

The colour of the halogen **fades** as the number of moles of the halogens remaining unreacted decreases.

Two bromine atoms bond at the 1st carbon in the triple bond while the other two goes to the 2nd carbon.

Examples

1Ethyne reacts with brown bromine vapour to form 1,1,2,2-tetrabromoethane.

Chemical equation

HC = CH + 2Br2  H Br2 C - CH Br2

H H H H

C = C + **2**Br – Br Br - C – C - Br

Br Br

Eth**y**ne + Bromine 1,1,2,1-tetrabromoethane

2.Prop**y**ne reacts with chlorine to form 1,1,2,2-tetrachloropropane.

Chemical equation

H3C C = CH + **2**Cl2  H3C CHCl2 - CHCl2

Prop**y**ne + Chlorine 1,1,2,2-tetrachloropropane

H H Cl H

H C C = C + 2Cl – Cl H - C – C - C- Cl

H H H Cl Cl

Prop**y**ne + Iodine 1,1,2,2-tetraiodopropane

H3C C = CH + **2**I2  H3C CHI2 - CHI2

H H H H H I H

H C C - C = C + 2I – I H - C- C – C - C- I

H H H H I I

3(a)But-1-**y**ne undergo halogenation to form 1,1,2,2-tetraiodobutane with iodine

Chemical equation

But-1-**y**ne + iodine 1,1,2,2-tetrabromobutane

H3C CH2 C = CH + **2**I2  H3C CH2C I2 - CHI2

H H H H I I

H C C - C = C -H + **2**I – I H - C- C – C - C- H

H H H H H I I

(b) But-2-**y**ne undergo halogenation to form 2,2,3,3-tetrafluorobutane with fluorine

But-2-**y**ne + Fluorine 2,2,3,3-tetrafluorobutane

H3C C = C -CH2 + **2**F2  H3C CF2CF2 - CH3

H H H H H H H H

H C C = C - C -H + F – F H - C- C – C - C- H

H H H H H H

4. But-1,3-di**y**ne should undergo halogenation to form 1,1,2,3,3,4,4 octaiodobutane. The reaction uses **four** moles of iodine molecules/**eight** iodine atoms to break the two(2) triple double bonds at carbon “1” and “2”.

But-1,3-di**e**ne + iodine 1,2,3,4-tetraiodobutane

H C = C C = C H + **4**I2  H C I2 C I2 C I2 C H I2

I I I I

H C C - C = C -H + **4**(I – I) H - C- C – C - C- H

I I I I

**(iii) Reaction with hydrogen halides.**

**Hydrogen halides** reacts with alk**y**ne to form a halogenoalk**e**ne then halogenoalk**a**ne. The triple bond in the alk**y**ne break and form a double then single bond.

The main compound is one which the **hydrogen** atom bond at the carbon with **more** **hydrogen** .

Examples

1. Ethyne reacts with hydrogen bromide to form bromoethane.

Chemical equation

H C = C H + 2HBr H3 C - CH Br2

H H H H

C = C + 2H – Br H - C – C - Br

H Br

Eth**y**ne + Bromine 1,1-dibromoethane

2. Prop**y**ne reacts with hydrogen iodide to form 2,2-diiodopropane (as the main product )

Chemical equation

H3C C = CH + 2HI H3C CHI2 - CH3

Propene + Chlorine 2,2-dichloropropane

Carbon atom with more Hydrogen atoms gets **extra** hydrogen

H H I H

H C C = **C** + 2H – I H - C – C - **C**- H

H H H I H

3. Both But-1-**y**ne and But-2-**y**ne reacts with hydrogen bromide to form 2,2- dibromobutane

Chemical equation

But-1-**e**ne + hydrogen bromide 2,2-dibromobutane

H3C CH2 C = CH + 2HBr H3C CH2CHBr -CH3

H H H H Br H

H C C - C = C + 2H – Br H - C- C – C - C- H

H H H H H Br H

But-2-**y**ne + Hydrogen bromide 2,2-dibromobutane

H3C C = C -CH3 + 2HBr H3C CBr2CH2 - CH3

H H H Br H H

H C C = C - C -H + 2Br – H H - C- C – C - C- H

H H H Br H H

4. But-1,3-di**e**ne react with hydrogen iodide to form 2,3- diiodobutane. The reaction uses **four** moles of hydrogen iodide molecules/**four** iodine atoms and two hydrogen atoms to break the two double bonds.

But-1,3-di**y**ne + iodine 2,2,3,3-tetraiodobutane

H C = C C = C H + **4H**I H3C C I2 C I2 CH3

H H H I I H

H C C - C = C -H + **4**(H – I) H - C- C – C - C- H

H I I H

**B.ALKANOLS(Alcohols)**

**(A) INTRODUCTION.**

Alkanols belong to a homologous series of organic compounds with a general formula **CnH2n +1 OH** and thus **-OH** as the functional group .The 1st ten alkanols include

|  |  |  |  |
| --- | --- | --- | --- |
| n | General / molecular formular | Structural formula | IUPAC name |
| 1 | CH3OH | H – C –O - H  │  H | Methan**ol** |
| 2 | CH3 CH2OH  C2H5 OH | H H  H C – C –O - H  │  H H | Ethan**ol** |
| 3 | CH3 (CH2)2OH  C3H7 OH | H H H  H C – C - C –O - H  │  H H H | Propan**ol** |
| 4 | CH3 (CH2)3OH  C4H9 OH | H H H H  H C – C - C - C –O - H  │  H H H H | Butan**ol** |
| 5 | CH3(CH2)4OH  C5H11 OH | H H H H H  H C – C - C- C- C –O - H  │  H H H H H | Pentan**ol** |
| 6 | CH3(CH2)5OH  C6H13 OH | H H H H H H  H C – C - C- C- C– C - O - H  │  H H H H H H | Hexan**ol** |
| 7 | CH3(CH2)6OH  C7H15 OH | H H H H H H H  H C – C - C- C- C– C –C- O - H  │  H H H H H H H | Heptan**ol** |
| 8 | CH3(CH2)7OH  C8H17 OH | H H H H H H H H  H C – C - C- C- C– C –C- C -O - H  │  H H H H H H H H | Octan**ol** |
| 9 | CH3(CH2)8OH  C9H19 OH | H H H H H H H H H  H C – C - C- C- C– C –C- C –C- O - H  │  H H H H H H H H H | Nonan**ol** |
| 10 | CH3(CH2)9OH  C10H21 OH | H H H H H H H H H H  H C – C - C- C- C– C –C- C –C- C-O - H  │  H H H H H H H H H H | Decan**ol** |

Alkanols like Hydrocarbons( alkanes/alkenes/alkynes) form a homologous series where:

(i)general name is derived from the alkane name then ending with “**-ol**”

(ii)the members have –OH as the fuctional group

(iii)they have the same general formula represented by R-OH where R is an alkyl group.

(iv) each member differ by –CH2 group from the next/previous.

(v)they show a similar and gradual change in their physical properties e.g. boiling and melting points.

(vi)they show similar and gradual change in their chemical properties.

**B. ISOMERS OF ALKANOLS.**

Alkanols exhibit both structural and position isomerism. The isomers are named by using the following basic guidelines:

(i)Like alkanes , identify the **longest** carbon chain to be the parent name.

(ii)Identify the position of the **-OH** functional group to give it the **smallest /lowest** position.

(iii) Identify the type and position of the **side** branches.

**Practice examples of isomers of alkanols**

**(i)Isomers of propanol C3H7OH**

CH3CH2CH2OH - Propan-1-ol

OH

CH3CHCH3 - Propan-2-ol

Propan-2-ol and Propan-1-ol are position isomers because only the position of the –OH functional group changes.

**(ii)Isomers of Butanol C4H9OH**

CH3 CH2 CH3 CH2 OH Butan-1-ol

CH3 CH2 CH CH3

OH Butan-2-ol

CH3

CH3 CH3 CH3

OH 2-methylpropan-2-ol

Butan-2-ol and Butan-1-ol are position isomers because only the position of the -OH functional group changes.

2-methylpropan-2-ol is both a structural and position isomers because both the position of the functional group and the arrangement of the atoms in the molecule changes.

**(iii)Isomers of Pentanol C5H11OH**

CH3 CH2 CH2CH2CH2 OH Pentan-1-ol (Position isomer)

CH3 CH2 CH CH3

OH Pentan-2-ol (Position isomer)

CH3 CH2 CH CH2 CH3

OH Pentan-3-ol (Position isomer)

CH3

CH3 CH2 CH2 C CH3

OH 2-methylbutan-2-ol (Position /structural isomer)

CH3

CH3 CH2 CH2 C CHOH

CH3 2,2-dimethylbutan-1-ol (Position /structural isomer)

CH3

CH3 CH2 CH C CH3

CH3 OH 2,3-dimethylbutan-1-ol (Position /structural isomer)

**(iv)1,2-dichloropropan-2-ol**

CClH2 CCl CH3

OH

**(v)1,2-dichloropropan-1-ol**

CClH2 CHCl CH2

OH

**(vi) Ethan1,2-diol**

H H

HOCH2CH2OH H-O - C - C – O-H

H H

**(vii) Propan1,2,3-triol** HOH H

HOCH2CHOHCH2OH H-O - C- C – C – O-H

H H H

**C. LABORATORY PREPARATION OF ALKANOLS.**

For decades the world over, people have been fermenting grapes juice, sugar, carbohydrates and starch to produce ethanol as a social drug for relaxation.

In large amount, drinking of ethanol by mammals /human beings causes mental and physical lack of coordination.

Prolonged intake of ethanol causes permanent mental and physical lack of coordination because it damages vital organs like the liver.

Fermentation is the reaction where sugar is converted to alcohol/alkanol using biological catalyst/enzymes in **yeast.**

It involves **three** processes:

(i)Conversion of starch to maltose using the enzyme **diastase.**

(C6H10O5)n (s) + H2O(l) --diastase enzyme --> C12H22O11(aq)

(Starch) (Maltose)

(ii)Hydrolysis of Maltose to glucose using the enzyme **maltase.**

C12H22O11(aq)+ H2O(l) -- maltase enzyme -->2 C6H12O6(aq)

(Maltose) (glucose)

(iii)Conversion of glucose to ethanol and carbon(IV)oxide gas using the enzyme **zymase.**

C6H12O6(aq) -- zymase enzyme --> 2 C2H5OH(aq) + 2CO2(g)

(glucose) (Ethanol)

At concentration greater than 15% by volume, the ethanol produced kills the yeast enzyme stopping the reaction.

To increases the concentration, fractional distillation is done to produce spirits (e.g. Brandy=40% ethanol).

Methanol is much more poisonous /toxic than ethanol.

Taken large quantity in small quantity it causes instant blindness and liver, killing the consumer victim within hours.

**School laboratory preparation of ethanol from fermentation of glucose**

Measure 100cm3 of pure water into a conical flask.

Add about five spatula end full of glucose.

Stir the mixture to dissolve.

Add about one spatula end full of yeast.

Set up the apparatus as below.

Conical flask

Lime water

Sugar solution and yeast

Preserve the mixture for about **three** days.

**D.PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOLS**

Use the prepared sample above for the following experiments that shows the characteristic properties of alkanols

1. **Role of yeast**

Yeast is a single cell fungus which contains the enzyme maltase and zymase that catalyse the fermentation process.

1. **Observations in lime water.**

A white precipitate is formed that dissolve to a colourless solution later. Lime water/Calcium hydroxide reacts with carbon(IV)0xide produced during the fermentation to form insoluble calcium carbonate and water.

More carbon (IV)0xide produced during fermentation react with the insoluble calcium carbonate and water to form soluble calcium hydrogen carbonate.

Ca(OH)2(aq)+ CO2 (g) -> CaCO3(s)

H2O(l) + CO2 (g) + CaCO3(s) -> Ca(HCO3) 2 (aq)

**(c)Effects on litmus paper**

Experiment

Take the prepared sample and test with both blue and red litmus papers.

Repeat the same with pure ethanol and methylated spirit.

Sample Observation table

|  |  |
| --- | --- |
| **Substance/alkanol** | **Effect on litmus paper** |
| Prepared sample | Blue litmus paper remain blue  Red litmus paper remain red |
| Absolute ethanol | Blue litmus paper remain blue  Red litmus paper remain red |
| Methylated spirit | Blue litmus paper remain blue  Red litmus paper remain red |

Explanation

Alkanols are neutral compounds/solution that have characteristic sweet smell and taste.

They have no effect on both blue and red litmus papers.

**(d)Solubility in water.**

Experiment

Place about 5cm3 of prepared sample into a clean test tube Add equal amount of distilled water.

Repeat the same with pure ethanol and methylated spirit.

Observation

**No layers** formed between the two liquids.

Explanation

Ethanol is **miscible** in water.Both ethanol and water are polar compounds .

The solubility of alkanols decrease with increase in the alkyl chain/molecular mass.

The alkyl group is insoluble in water while –OH functional group is soluble in water.

As the molecular chain becomes **longer** ,the effect of the **alkyl** group **increases** as the effect of the functional group **decreases**.

**e)Melting/boiling point.**

Experiment

Place pure ethanol in a long boiling tube .Determine its boiling point.

Observation

Pure ethanol has a boiling point of 78oC at sea level/one atmosphere pressure.

Explanation

The melting and boiling point of alkanols increase with increase in molecular chain/mass .

This is because the intermolecular/van-der-waals forces of attraction between the molecules increase.

More heat energy is thus required to weaken the longer chain during melting and break during boiling.

**f)Density**

Density of alkanols increase with increase in the intermolecular/van-der-waals forces of attraction between the molecule, making it very close to each other.

This reduces the volume occupied by the molecule and thus increase the their mass per unit volume (density).

**Summary table showing the trend in physical properties of alkanols**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Alkanol | Melting point  (oC) | Boiling point  (oC) | Density  gcm-3 | Solubility in water |
| Methanol | -98 | 65 | 0.791 | soluble |
| Ethanol | -117 | 78 | 0.789 | soluble |
| Propanol | -103 | 97 | 0.803 | soluble |
| Butanol | -89 | 117 | 0.810 | Slightly soluble |
| Pentanol | -78 | 138 | 0.814 | Slightly soluble |
| Hexanol | -52 | 157 | 0.815 | Slightly soluble |
| Heptanol | -34 | 176 | 0.822 | Slightly soluble |
| Octanol | -15 | 195 | 0.824 | Slightly soluble |
| Nonanol | -7 | 212 | 0.827 | Slightly soluble |
| Decanol | 6 | 228 | 0.827 | Slightly soluble |

**g)Burning**

Experiment

Place the prepared sample in a watch glass. Ignite. Repeat with pure ethanol and methylated spirit.

Observation/Explanation

Fermentation produce ethanol with a lot of water(about a ratio of 1:3)which prevent the alcohol from igniting.

Pure ethanol and methylated spirit easily catch fire / highly flammable.

They burn with an almost colourless non-sooty/non-smoky **blue** flame to form **carbon(IV) oxide** (in excess air/oxygen)or **carbon(II) oxide** (limited air) and **water**.

Ethanol is thus a **saturated** compound like alkanes.

Chemica equation

C2 H5OH(l) + 3O2 (g) -> 3H2O(l) + 2CO2 (g) ( excess air)

C2 H5OH(l) + 2O2 (g) -> 3H2O(l) + 2CO (g) ( limited air)

2CH3OH(l) + 3O2 (g) -> 4H2O(l) + 2CO2 (g) ( excess air)

2 CH3OH(l) + 2O2 (g) -> 4H2O(l) + 2CO (g) ( limited air)

2C3 H7OH(l) + 9O2 (g) -> 8H2O(l) + 6CO2 (g) ( excess air)

C3 H7OH(l) + 3O2 (g) -> 4H2O(l) + 3CO (g) ( limited air)

2C4 H9OH(l) + 13O2 (g) -> 20H2O(l) + 8CO2 (g) ( excess air)

C4 H9OH(l) + 3O2 (g) -> 4H2O(l) + 3CO (g) ( limited air)

Due to its flammability, ethanol is used;

1. as a fuel in spirit lamps
2. as gasohol when blended with gasoline

**(h)Formation of alkoxides**

Experiment

Cut a very small piece of sodium. Put it in a beaker containing about 20cm3 of the prepared sample in a beaker.

Test the products with litmus papers. Repeat with absolute ethanol and methylated spirit.

Sample observations

|  |  |
| --- | --- |
| Substance/alkanol | Effect of adding sodium |
| Fermentation prepared sample | (i)effervescence/fizzing/bubbles  (ii)colourless gas produced that extinguish burning splint with explosion/ “Pop” sound  (iii)colourless solution formed  (iv)blue litmus papers remain blue  (v)red litmus papers turn blue |
| Pure/absolute ethanol/methylated spirit | (i)**slow** effervescence/fizzing/bubbles  (ii)colourless gas **slowly**  produced that extinguish burning splint with explosion/ “Pop” sound  (iii)colourless solution formed  (iv)blue litmus papers remain blue  (v)red litmus papers turn blue |

Explanations

Sodium/potassium reacts slowly with alkanols to form basic solution called **alkoxides** and producing **hydrogen** gas.

If the alkanol has some water the metals react faster with the water to form **soluble hydroxides/alkalis** i.e.

Sodium + Alkanol -> Sodium **alk**oxides + Hydrogen gas

Potassium + Alkanol -> Potassium **alk**oxides + Hydrogen gas

Sodium + Water -> Sodium **hydr**oxides + Hydrogen gas

Potassium + Water -> Potassium **hydr**oxides + Hydrogen gas

Examples

1.Sodium metal reacts with ethanol to form sodium **eth**oxide

Sodium metal reacts with water to form sodium **Hydr**oxide

2CH3CH2OH(l) + 2Na(s) -> 2CH3CH2ONa (aq) + H2 (s)

2H2O(l) + 2Na(s) -> 2NaOH (aq) + H2 (s)

2.Potassium metal reacts with ethanol to form Potassium **eth**oxide

Potassium metal reacts with water to form Potassium **Hydr**oxide

2CH3CH2OH(l) + 2K(s) -> 2CH3CH2OK (aq) + H2 (s)

2H2O(l) + 2K(s) -> 2KOH (aq) + H2 (s)

3.Sodium metal reacts with propanol to form sodium **prop**oxide

Sodium metal reacts with water to form sodium **Hydr**oxide

2CH3CH2 CH2OH(l) + 2Na(s) -> 2CH3CH2 CH2ONa (aq) + H2 (s)

2H2O(l) + 2Na(s) -> 2NaOH (aq) + H2 (s)

4.Potassium metal reacts with propanol to form Potassium **prop**oxide

Potassium metal reacts with water to form Potassium **Hydr**oxide

2CH3CH2 CH2OH(l) + 2K(s) -> 2CH3CH2 CH2OK (aq) + H2 (s)

2H2O(l) + 2K(s) -> 2KOH (aq) + H2 (s)

5.Sodium metal reacts with butanol to form sodium **but**oxide

Sodium metal reacts with water to form sodium **Hydr**oxide

2CH3CH2 CH2 CH2OH(l) + 2Na(s) -> 2CH3CH2 CH2 CH2ONa (aq) + H2 (s)

2H2O(l) + 2Na(s) -> 2NaOH (aq) + H2 (s)

6.Sodium metal reacts with pentanol to form sodium **pent**oxide

Sodium metal reacts with water to form sodium **Hydr**oxide

2CH3CH2 CH2 CH2 CH2OH(l)+2Na(s) -> 2CH3CH2 CH2 CH2 CH2ONa (aq) + H2 (s)

2H2O(l) + 2Na(s) -> 2NaOH (aq) + H2 (s)

**(i)Formation of Esters/Esterification**

Experiment

Place 2cm3 of ethanol in a boiling tube.

Add equal amount of ethanoic acid.To the mixture add carefully 2drops of concentrated sulphuric(VI)acid.

Warm/Heat gently.

Pour the mixture into a beaker containing about 50cm3 of cold water.

Smell the products.

Repeat with methanol

Sample observations

|  |  |
| --- | --- |
| **Substance/alkanol** | **Effect on adding equal amount of ethanol/concentrated sulphuric(VI)acid** |
| Absolute ethanol | Sweet fruity smell |
| Methanol | Sweet fruity smell |

Explanation

Alkanols react with alkanoic acids to form a group of homologous series of sweet smelling compounds called esters and water. This reaction is catalyzed by concentrated sulphuric(VI)acid in the laboratory.

Alkanol + Alkanoic acid –Conc. H2SO4-> Ester + water

Naturally esterification is catalyzed by sunlight. Each ester has a characteristic smell derived from the many possible combinations of alkanols and alkanoic acids that create a variety of known natural(mostly in fruits) and synthetic(mostly in juices) esters .

Esters derive their names from the alkanol first then alkanoic acids. The alkanol “becomes” an **alkyl** group and the alkanoic acid “becomes” **alkanoate** hence **alkylalkanoate**. e.g.

Ethanol + Ethanoic acid -> Ethylethanoate + Water

Ethanol + Propanoic acid -> Ethylpropanoate + Water

Ethanol + Methanoic acid -> Ethylmethanoate + Water

Ethanol + butanoic acid -> Ethylbutanoate + Water

Propanol + Ethanoic acid -> Propylethanoate + Water

Methanol + Ethanoic acid -> Methyethanoate + Water

Methanol + Decanoic acid -> Methyldecanoate + Water

Decanol + Methanoic acid -> Decylmethanoate + Water

During the formation of the ester, the “O” joining the alkanol and alkanoic acid comes from the alkanol.

R1 -CO**OH** + R2 –O**H** -> **R1 -COO –R2** + **H2O**

e.g.

1. Ethanol reacts with ethanoic acid to form the ester ethyl ethanoate and water.

Ethanol + Ethanoic acid --Conc. H2SO4 -->Ethylethanoate + Water

C2H5OH (l) + CH3COOH(l) --Conc. H2SO4 --> CH3COO C2H5(**aq**)+H2O(l)

CH3CH2OH (l)+ CH3COOH(l) --Conc. H2SO4 --> CH3COOCH2CH3(**aq**)+H2O(l)

2. Ethanol reacts with propanoic acid to form the ester ethylpropanoate and water.

Ethanol + Propanoic acid --Conc. H2SO4 -->Ethylethanoate + Water

C2H5OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->CH3CH2COO C2H5(**aq**)+H2O(l)

CH3CH2OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->

CH3 CH2COOCH2CH3(**aq**)+H2O(l)

3. Methanol reacts with ethanoic acid to form the ester methyl ethanoate and water.

Methanol + Ethanoic acid --Conc. H2SO4 -->Methylethanoate + Water

CH3OH (l) + CH3COOH(l) --Conc. H2SO4 --> CH3COO CH3(**aq**)+H2O(l)

4. Methanol reacts with propanoic acid to form the ester methyl propanoate and water.

Methanol + propanoic acid --Conc. H2SO4 -->Methylpropanoate + Water

CH3OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 --> CH3 CH2COO CH3(**aq**)+H2O(l)

5. Propanol reacts with propanoic acid to form the ester propylpropanoate and water.

Propanol + Propanoic acid --Conc. H2SO4 -->Ethylethanoate + Water

C3H7OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->CH3CH2COO C3H7(**aq**)+H2O(l)

CH3CH2 CH2OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->

CH3 CH2COOCH2 CH2CH3(**aq**)+H2O(l)

**(j)Oxidation**

Experiment

Place 5cm3 of absolute ethanol in a test tube.Add three drops of acidified potassium manganate(VII).Shake thoroughly for one minute/warm.Test the solution mixture using pH paper. Repeat by adding acidified potassium dichromate(VII).

Sample observation table

|  |  |  |  |
| --- | --- | --- | --- |
| Substance/alkanol | Adding acidified KMnO4/K2Cr2O7 | pH of resulting solution/mixture | Nature of resulting solution/mixture |
| Pure ethanol | (i)Purple colour of KMnO4decolorized  (ii) Orange colour of K2Cr2O7turns green. | pH= 4/5/6  pH = 4/5/6 | Weakly acidic  Weakly acidic |

Explanation

Both acidified KMnO4 and K2Cr2O7 are oxidizing agents(add oxygen to other compounds. They oxidize alkan**o**ls to a group of homologous series called alkan**a**ls then further oxidize them to alkan**oic** acids.The oxidizing agents are themselves reduced hence changing their colour:

(i) Purple KMnO4 is reduced to colourless Mn2+

(ii)Orange K2Cr2O7is reduced to green Cr3+

The pH of alkanoic acids show they have few H+ because they are weak acids i.e

Alkanol + [O] -> Alkanal + [O] -> alkanoic acid

**NB** The [O] comes from the oxidizing agents acidified KMnO4 or K2Cr2O7

Examples

1.When ethanol is warmed with three drops of acidified KMnO4 there is decolorization of KMnO4

Ethanol + [O] -> Ethanal + [O] -> Ethanoic acid

CH3CH2OH + [O] -> CH3CH2O + [O] -> CH3COOH

2.When methanol is warmed with three drops of acidified K2Cr2O7 ,the orange colour of acidified K2Cr2O7 changes to green.

methanol + [O] -> methanal + [O] -> methanoic acid

CH3OH + [O] -> CH3O + [O] -> HCOOH

3.When propanol is warmed with three drops of acidified K2Cr2O7 ,the orange colour of acidified K2Cr2O7 changes to green.

Propanol + [O] -> Propanal + [O] -> Propanoic acid

CH3CH2 CH2OH + [O] -> CH3CH2 CH2O + [O] -> CH3 CH2COOH

4.When butanol is warmed with three drops of acidified K2Cr2O7 ,the orange colour of acidified K2Cr2O7 changes to green.

Butanol + [O] -> Butanal + [O] -> Butanoic acid

CH3CH2 CH2 CH2OH + [O] ->CH3CH2 CH2CH2O+[O] -> CH3 CH2COOH

Air slowly oxidizes ethanol to dilute ethanoic acid commonly called **vinegar.** If beer is not tightly corked, a lot of carbon(IV)oxide escapes and there is slow oxidation of the beer making it “flat”.

**(k)Hydrolysis /Hydration and Dehydration**

**I. Hydrolysis/Hydration** is the reaction of a compound/substance with water.

Alkenes react with water vapour/steam at high temperatures and high pressures in presence of phosphoric acid catalyst to form alkanols.i.e.

Alkenes + Water - H3PO4 catalyst-> Alkanol

Examples

(i)Ethene is mixed with steam over a phosphoric acid catalyst at 300oC temperature and 60 atmosphere pressure to form ethanol

Ethene + water ---60 atm/300oC/ H3PO4 --> Ethanol

H2C =CH2 (g) + **H2O**(l)--60 atm/300oC/ H3PO4 --> CH3 C**H2O**H(l)

This is the main method of producing large quantities of ethanol instead of fermentation

(ii) Propene + water ---60 atm/300oC/ H3PO4 --> Propanol

CH3C =CH2 (g) + **H2O**(l)--60 atm/300oC/ H3PO4 --> CH3 CH2 C**H2O**H(l)

(iii) Butene + water ---60 atm/300oC/ H3PO4 --> Butanol

CH3 CH2 C=CH2 (g) + **H2O**(l)--60 atm/300oC/ H3PO4 --> CH3 CH2 CH2 C**H2O**H(l)

**II. Dehydration** is the process which concentrated sulphuric(VI)acid (**dehydrating agent**) removes water from a compound/substances.

Concentrated sulphuric(VI)acid dehydrates alkanols to the corresponding alkenes at about 180oC. i.e

Alkanol --Conc. H2 SO4/180oC--> Alkene + Water

Examples

1. At 180oC and in presence of Concentrated sulphuric(VI)acid, ethanol undergoes dehydration to form ethene.

Ethanol ---180oC/ H2SO4 --> Ethene + Water

CH3 C**H2O**H(l)--180oC/ H2SO4 --> H2C =CH2 (g) + **H2O**(l)

2. Propanol undergoes dehydration to form propene.

Propanol ---180oC/ H2SO4 --> Propene + Water

CH3 CH2 C**H2O**H(l)--180oC/ H2SO4 --> CH3CH =CH2 (g) + **H2O**(l)

3. Butanol undergoes dehydration to form Butene.

Butanol ---180oC/ H2SO4 --> Butene + Water

CH3 CH2 CH2C**H2O**H(l)--180oC/ H2SO4 --> CH3 CH2C =CH2 (g) + **H2O**(l)

3. Pentanol undergoes dehydration to form Pentene.

Pentanol ---180oC/ H2SO4 --> Pentene + Water

CH3 CH2 CH2 CH2 C**H2O**H(l)--180oC/ H2SO4-->CH3 CH2 CH2C =CH2 (g)+**H2O**(l)

**(l)Similarities of alkanols with Hydrocarbons**

**I.** Similarity with alkanes

Both alkanols and alkanes burn with a **blue non-sooty flame** to form carbon(IV)oxide(in excess air/oxygen)/carbon(II)oxide(in limited air) and water. This shows they are saturated with high C:H ratio. e.g.

Both ethanol and ethane ignite and burns in air with a **blue non-sooty flame** to form carbon(IV)oxide(in excess air/oxygen)/carbon(II)oxide(in limited air) and water.

CH2 CH2OH(l) + 3O2(g) -Excess air-> 2CO2 (g) + 3H2 O(l)

CH2 CH2OH(l) + 2O2(g) -Limited air-> 2CO(g) + 3H2 O(l)

CH3 CH3(g) + 3O2(g) -Excess air-> 2CO2 (g) + 3H2 O(l)

2CH3 CH3(g) + 5O2(g) -Limited air-> 4CO(g) + 6H2 O(l)

**II.** Similarity with alkenes/alkynes

Both alkanols(R-OH) and alkenes/alkynes(with = C = C = double and – C = C- triple ) bond:

(i)decolorize acidified KMnO4

(ii)turns Orange acidified K2Cr2O7 to green.

Alkan**o**ls(R-**OH**) are oxidized to alkan**a**ls(R-**O**) ant then alkan**oic** acids(R-**OOH**).

Alkenes are oxidized to alkanols with duo/double functional groups.

Examples

1.When ethanol is warmed with three drops of acidified K2Cr2O7 the orange of acidified K2Cr2O7 turns to green. Ethanol is oxidized to ethanol and then to ethanoic acid.

Ethan**o**l + [O] -> Ethan**a**l + [O] -> Ethan**oic** acid

CH3CH2OH + [O] -> CH3CH2O + [O] -> CH3COOH

2.When ethene is bubbled in a test tube containing acidified K2Cr2O7 ,the orange of acidified K2Cr2O7 turns to green. Ethene is oxidized to ethan-1,2-diol.

Ethene + [O] -> Ethan-1,2-diol.

H2C=CH2 + [O] -> HOCH2 -CH2OH

**III.** Differences with alkenes/alkynes

Alkanols do not decolorize bromine and chlorine water.

Alkenes decolorizes bromine and chlorine water to form halogenoalkanols

Example

When ethene is bubbled in a test tube containing bromine water,the bromine water is decolorized. Ethene is oxidized to bromoethanol.

Ethene + Bromine water -> Bromoethanol.

H2C=CH2 + HOBr -> BrCH2 -CH2OH

**IV.** Differences in melting and boiling point with Hydrocarbons

Alkanos have higher melting point than the corresponding hydrocarbon (alkane/alkene/alkyne)

This is because most alkanols exist as **dimer.**A dimer is a molecule made up of two other molecules joined usually by van-der-waals forces/hydrogen bond or dative bonding.

Two alkanol molecules form a dimer joined by hydrogen bonding.

Example

In Ethanol the oxygen atom attracts/pulls the shared electrons in the covalent bond more to itself than Hydrogen.

This creates a partial negative charge (δ-) on oxygen and partial positive charge(δ+) on hydrogen.

Two ethanol molecules attract each other at the partial charges through Hydrogen bonding forming a **dimme**r.

H H H

Hydrogen bonds

Covalent bonds

H C C O H H

H H H O C C H

H H

Dimerization of alkanols means more energy is needed to break/weaken the Hydrogen bonds before breaking/weakening the intermolecular forces joining the molecules of all organic compounds during boiling/melting.

E.**USES OF SOME ALKANOLS**

(a)Methanol is used as industrial alcohol and making methylated spirit

(b)Ethanol is used:

1. as alcohol in alcoholic drinks e.g Beer, wines and spirits.

2.as antiseptic to wash woulds

3.in manufacture of vanishes, ink ,glue and paint because it is volatile and thus easily evaporate

4.as a fuel when blended with petrol to make gasohol.

**B.ALKANOIC ACIDS (Carboxylic acids)**

**(A) INTRODUCTION.**

Alkanoic acids belong to a homologous series of organic compounds with a general formula **CnH2n +1 COOH** and thus **-COOH** as the functional group .The 1st ten alkanoic acids include:

|  |  |  |  |
| --- | --- | --- | --- |
| n | General /molecular  formular | Structural formula | IUPAC name |
| 0 | HCOOH | H – C –O - H  │  O | Methanoic acid |
| 1 | CH3 COOH | H  H – C – C – O - H  │  H O | Ethanoic acid |
| 2 | CH3 CH2 COOH  C2 H5 COOH | H H  H-C – C – C – O – H    H H O | Propanoic acid |
| 3 | CH3 CH2 CH2 COOH  C3 H7 COOH | H H H  H- C - C – C – C – O – H    H H H O | Butanoic acid |
| 4 | CH3CH2CH2CH2 COOH  C4 H9 COOH | H H H H  H - C – C - C – C – C – O – H    H H H H O | Pentanoic acid |
| 5 | CH3CH2 CH2CH2CH2 COOH  C5 H11 COOH | H H H H H  H C - C – C - C – C – C – O – H    H H H H H O | Hexanoic acid |
| 6 | CH3CH2 CH2 CH2CH2CH2 COOH  C6 H13 COOH | H H H H H H  H C C - C – C - C – C – C – O – H    H H H H H H O | Pentanoic acid |

Alkanoic acids like alkanols /alkanes/alkenes/alkynes form a homologous series where:

(i)the general name of an alkanoic acids is derived from the alkane name then ending with “**–oic**” acid as the table above shows.

(ii) the members have R-**COOH**/R C-O-H as the functional group.

O

(iii)they have the same general formula represented by R-COOH where R is an alkyl group.

(iv)each member differ by –CH2- group from the next/previous.

(v)they show a similar and gradual change in their physical properties e.g. boiling and melting point.

(vi)they show similar and gradual change in their chemical properties.

(vii) since they are acids they show similar properties with mineral acids.

**(B) ISOMERS OF ALKANOIC ACIDS.**

Alkanoic acids exhibit both structural and position isomerism. The isomers are named by using the following basic guidelines

(i)Like alkanes. identify the longest carbon chain to be the parent name.

(ii)Identify the position of the -C-O-H functional group to give it the smallest

O

/lowest position.

(iii)Identify the type and position of the side group branches.

**Practice examples on isomers of alkanoic acids**

1.Isomers of butanoic acid C3H7COOH

CH3 CH2 CH2 COOH

Butan-1-oic acid

CH3

H2C C COOH 2-methylpropan-1-oic acid

2-methylpropan-1-oic acid and Butan-1-oic acid are structural isomers because the position of the functional group does not change but the arrangement of the atoms in the molecule does.

2.Isomers of pentanoic acid C4H9COOH

CH3CH2CH2CH2 COOH pentan-1-oic acid

CH3

CH3CH2CH COOH 2-methylbutan-1-oic acid

CH3

H3C C COOH 2,2-dimethylpropan-1-oic acid

CH3

3.Ethan-1,2-dioic acid

O O

HOOC- COOH // H - O – C - C – O – H

4.Propan-1,3-dioic acid

O H O

HOOC- CH2COOH // H - O – C – C - C – O – H

H

5.Butan-1,4-dioic acid

O H H O

HOOC CH2 CH2 COOH H- O – C – C - C – C –O – H

H H

6.2,2-dichloroethan-1,2-dioic acid

HOOCCHCl2  Cl

H – O - C – C – Cl

O H

**(C) LABORATORY AND INDUSTRIAL PREPARATIONOF ALKANOIC ACIDS.**

**In a school laboratory**, alkanoic acids can be prepared by adding an oxidizing agent (H+/KMnO4 or H+/K2Cr2O7)to the corresponding alkanol then warming.

The oxidation converts the alkanol first to an alkanal the alkanoic acid.

**NB** Acidified KMnO4 is a stronger oxidizing agent than acidified K2Cr2O7

General equation:

R- CH2 – **OH** + [O] --H+/KMnO4--> R- CH –**O** + H2O(l)

(alkanol) (alkanal)

R- CH – **O** + [O] --H+/KMnO4--> R- C –**OOH**

(alkanal) (alkanoic acid)

Examples

1.Ethan**o**l on warming in acidified KMnO4 is oxidized to ethan**a**l then ethan**oic** acid .

CH3- CH2 – **OH** + [O] --H+/KMnO4--> CH3- CH –**O** + H2O(l)

(ethanol) (ethanal)

CH3- CH – **O** + [O] --H+/KMnO4--> CH3- C –**OOH**

(ethanal) (ethanoic acid)

2Propan**o**l on warming in acidified KMnO4 is oxidized to propan**a**l then propan**oic** acid

CH3- CH2 CH2 – **OH** + [O] --H+/KMnO4--> CH3- CH2 CH –**O** + H2O(l)

(propanol) (propanal)

CH3- CH – **O** + [O] --H+/KMnO4--> CH3- C –**OOH**

(propanal) (propanoic acid)

Industrially,large scale manufacture of alkanoic acid like ethanoic acid is obtained from:

(a)Alk**e**nes reacting with steam at high temperatures and pressure in presence of phosphoric(V)acid catalyst and undergo hydrolysis to form alkanols. i.e.

Alkenes + Steam/water -- H2PO4 Catalyst--> Alkanol

The alkanol is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the alkanoic acid.

Alkanol + Air -- MnSO4 Catalyst/5 atm pressure--> Alkanoic acid

Example

Ethene is mixed with steam over a phosphoric(V)acid catalyst,300oC temperature and 60 atmosphere pressure to form ethanol.

CH2=CH2 + H2O -> CH3 CH2OH

(Ethene) (Ethanol)

This is the industrial large scale method of manufacturing ethanol

Ethanol is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the ethanoic acid.

CH3 CH2OH + [O] -- MnSO4 Catalyst/5 atm pressure--> CH3 COOH

(Ethanol) (Ethanoic acid)

(b)Alk**y**nes react with liquid water at high temperatures and pressure in presence of Mercury(II)sulphate(VI)catalyst and 30% concentrated sulphuric(VI)acid to form alkan**a**ls.

Alk**y**ne + Water -- Mercury(II)sulphate(VI)catalyst--> Alkan**a**l

The alkan**a**l is then oxidized by air at 5 atmosphere pressure with Manganese (II) sulphate(VI) catalyst to form the alkan**oic** acid.

Alkan**a**l + air/oxygen -- Manganese(II)sulphate(VI)catalyst--> Alkan**oic acid**

Example

Eth**y**ne react with liquid water at high temperature and pressure with Mercury (II) sulphate (VI)catalyst and 30% concentrated sulphuric(VI)acid to form ethan**a**l.

CH = CH + H2O --HgSO4--> CH3 CH2O

(Ethyne) (Ethan**a**l)

This is another industrial large scale method of manufacturing ethanol from large quantities of ethyne found in natural gas.

Ethan**a**l is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the ethanoic acid.

CH3 CH2O + [O] -- MnSO4 Catalyst/5 atm pressure--> CH3 COOH

(Ethan**a**l) (Oxygen from air) (Ethanoic acid)

**(D) PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOIC ACIDS.**

**I.Physical properties of alkanoic acids**

The table below shows some physical properties of alkanoic acids

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Alkanol | Melting point(oC) | Boiling point(oC) | Density(gcm-3) | Solubility in water |
| Methanoic acid | 18.4 | 101 | 1.22 | soluble |
| Ethanoic acid | 16.6 | 118 | 1.05 | soluble |
| Propanoic acid | -2.8 | 141 | 0.992 | soluble |
| Butanoic acid | -8.0 | 164 | 0.964 | soluble |
| Pentanoic acid | -9.0 | 187 | 0.939 | Slightly soluble |
| Hexanoic acid | -11 | 205 | 0.927 | Slightly soluble |
| Heptanoic acid | -3 | 223 | 0.920 | Slightly soluble |
| Octanoic acid | 11 | 239 | 0.910 | Slightly soluble |
| Nonanoic acid | 16 | 253 | 0.907 | Slightly soluble |
| Decanoic acid | 31 | 269 | 0.905 | Slightly soluble |

From the table note the following:

1. Melting and boiling point decrease as the carbon chain increases due to increase in intermolecular forces of attraction between the molecules requiring more energy to separate the molecules.
2. The density decreases as the carbon chain increases as the intermolecular forces of attraction increases between the molecules making the molecule very close reducing their volume in unit mass.
3. Solubility decreases as the carbon chain increases as the soluble –COOH end is shielded by increasing insoluble alkyl/hydrocarbon chain.
4. Like alkanols ,alkanoic acids exist as dimmers due to the hydrogen bonds within the molecule. i.e..





**II Chemical properties of alkanoic acids**

The following experiments shows the main chemical properties of ethanoic (alkanoic) acid.

**(a)Effect on litmus papers**

Experiment

Dip both blue and red litmus papers in ethanoic acid. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute nitric(V)acid.

Sample observations

|  |  |  |
| --- | --- | --- |
| Solution/acid | Observations/effect on litmus papers | Inference |
| Ethanoic acid | Blue litmus paper turn red  Red litmus paper remain red | H3O+/H+(aq)ion |
| Succinic acid | Blue litmus paper turn red  Red litmus paper remain red | H3O+/H+(aq)ion |
| Citric acid | Blue litmus paper turn red  Red litmus paper remain red | H3O+/H+(aq)ion |
| Oxalic acid | Blue litmus paper turn red  Red litmus paper remain red | H3O+/H+(aq)ion |
| Tartaric acid | Blue litmus paper turn red  Red litmus paper remain red | H3O+/H+(aq)ion |
| **Nitric(V)acid** | Blue litmus paper turn red  Red litmus paper remain red | H3O+/H+(aq)ion |

Explanation

All acidic solutions contains H+/H3O+(aq) ions. The H+ /H3O+ (aq) ions is responsible for turning blue litmus paper/solution to red

**(b)pH**

Experiment

Place 2cm3 of ethaoic acid in a test tube. Add 2 drops of universal indicator solution and determine its pH. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI)acid.

Sample observations

|  |  |  |
| --- | --- | --- |
| Solution/acid | pH | Inference |
| Ethanoic acid | 4/5/6 | Weakly acidic |
| Succinic acid | 4/5/6 | Weakly acidic |
| Citric acid | 4/5/6 | Weakly acidic |
| Oxalic acid | 4/5/6 | Weakly acidic |
| Tartaric acid | 4/5/6 | Weakly acidic |
| **Sulphuric(VI)acid** | **1/2/3** | **Strongly acidic** |

Explanations

Alkanoic acids are weak acids that partially/partly dissociate to release few H+ ions in solution. The pH of their solution is thus 4/5/6 showing they form weakly acidic solutions when dissolved in water.

All alkanoic acid dissociate to releases the **“H”** at the functional group in -COO**H** to form the **alkanoate ion;** –COO**-**

Mineral acids(Sulphuric(VI)acid, Nitric(V)acid and Hydrochloric acid) are strong acids that wholly/fully dissociate to release many H+ ions in solution. The pH of their solution is thus 1/2/3 showing they form strongly acidic solutions when dissolved in water.i.e

Examples

1. CH3COO**H**(aq) CH3COO-(aq) + H+(aq)

(ethanoic acid) (ethanoate ion) (few H+ ion)

1. CH3 CH2COO**H**(aq) CH3 CH2COO-(aq) + H+(aq)

(propanoic acid) (propanoate ion) (few H+ ion)

1. CH3 CH2 CH2COO**H**(aq) CH3 CH2 CH2COO-(aq) + H+(aq)

(Butanoic acid) (butanoate ion) (few H+ ion)

1. HOO**H**(aq) HOO-(aq) + H+(aq)

(methanoic acid) (methanoate ion) (few H+ ion)

1. H2 SO4 (aq) SO42- (aq) + 2H+(aq)

(sulphuric(VI) acid) (sulphate(VI) ion) (**many** H+ ion)

1. HNO3 (aq) NO3- (aq) + H+(aq)

(nitric(V) acid) (nitrate(V) ion) (**many** H+ ion)

(c)Reaction with metals

Experiment

Place about 4cm3 of ethanoic acid in a test tube. Put about 1cm length of polished magnesium ribbon. Test any gas produced using a burning splint. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

|  |  |  |
| --- | --- | --- |
| Solution/acid | Observations | Inference |
| Ethanoic acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that burn with “pop” sound/explosion | H3O+/H+(aq)ion |
| Succinic acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that burn with “pop” sound/explosion | H3O+/H+(aq)ion |
| Citric acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that burn with “pop” sound/explosion | H3O+/H+(aq)ion |
| Oxalic acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that burn with “pop” sound/explosion | H3O+/H+(aq)ion |
| Tartaric acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that burn with “pop” sound/explosion | H3O+/H+(aq)ion |
| **Nitric(V)acid** | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that burn with “pop” sound/explosion | H3O+/H+(aq)ion |

Explanation

Metals higher in the reactivity series displace the hydrogen in all acids to evolve/produce hydrogen gas and form a salt. Alkanoic acids react with metals with metals to form alkanoates salt and produce/evolve hydrogen gas .Hydrogen extinguishes a burning splint with a pop sound/explosion. Only the “H”in the functional group -COO**H** is /are displaced and not in the alkyl hydrocarbon chain.

Alkanoic acid + Metal -> Alkanoate + Hydrogen gas. i.e.

Examples

1. For a monovalent metal with monobasic acid

2R – COO**H** + 2M -> 2R- COO**M** + 2H2(g)

2.For a divalent metal with monobasic acid

2R – COO**H** + M -> (R- COO) 2**M** + H2(g)

3.For a divalent metal with dibasic acid

**H**OOC-R-COO**H**+ M -> **M**OOC-R-COO**M** + H2(g)

4.For a monovalent metal with dibasic acid

**H**OOC-R-COO**H**+ 2M -> **M**OOC-R-COO**M** + H2(g)

5 For mineral acids

(i)Sulphuric(VI)acid is a dibasic acid

**H2** SO4 (aq) + 2M -> M2 SO4 (aq) + H2(g)

**H2** SO4 (aq) + M -> MSO4 (aq) + H2(g)

(ii)Nitric(V) and hydrochloric acid are monobasic acid

**H**NO3 (aq) + 2M -> 2MNO3 (aq) + H2(g)

**H**NO3 (aq) + M -> M(NO3 ) 2 (aq) + H2(g)

Examples

1.Sodium reacts with ethanoic acid to form sodium ethanoate and produce. hydrogen gas.

**Caution**: This reaction is explosive.

CH3COO**H** (aq) + Na(s) -> CH3COO**Na** (aq) + H2(g)

(Ethanoic acid) (Sodium ethanoate)

2.Calcium reacts with ethanoic acid to form calcium ethanoate and produce. hydrogen gas.

2CH3COO**H** (aq) + Ca(s) -> (CH3COO) 2**Ca** (aq) + H2(g)

(Ethanoic acid) (Calcium ethanoate)

3.Sodium reacts with ethan-1,2-dioic acid to form sodium ethan-1,2-dioate and produce. hydrogen gas.

**H**OOC-COO**H**+ 2Na -> **Na**OOC - COO**Na** + H2(g)

(ethan-1,2-dioic acid) (sodium ethan-1,2-dioate)

Commercial name of ethan-1,2-dioic acid is oxalic acid. The salt is sodium oxalate.

4.Magnesium reacts with ethan-1,2-dioic acid to form magnesium ethan-1,2-dioate and produce. hydrogen gas.

**H**OOC-R-COO**H**+ Mg -> ( OOC - COO)**Mg** + H2(g)

(ethan-1,2-dioic acid) (magnesium ethan-1,2-dioate)

5.Magnesium reacts with

(i)Sulphuric(VI)acid to form Magnesium sulphate(VI)

**H2** SO4 (aq) + Mg -> MgSO4 (aq) + H2(g)

(ii)Nitric(V) and hydrochloric acid are monobasic acid

2**H**NO3 (aq) + Mg -> M(NO3 ) 2 (aq) + H2(g)

**(d)Reaction with hydrogen carbonates and carbonates**

Experiment

Place about 3cm3 of ethanoic acid in a test tube. Add about 0.5g/ ½ spatula end full of sodium hydrogen carbonate/sodium carbonate. Test the gas produced using lime water. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

|  |  |  |
| --- | --- | --- |
| Solution/acid | Observations | Inference |
| Ethanoic acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that forms a white precipitate with lime water | H3O+/H+(aq)ion |
| Succinic acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that forms a white precipitate with lime water | H3O+/H+(aq)ion |
| Citric acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that forms a white precipitate with lime water | H3O+/H+(aq)ion |
| Oxalic acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that forms a white precipitate with lime water | H3O+/H+(aq)ion |
| Tartaric acid | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that forms a white precipitate with lime water | H3O+/H+(aq)ion |
| **Nitric(V)acid** | (i)effervescence, fizzing, bubbles  (ii)colourless gas produced that forms a white precipitate with lime water | H3O+/H+(aq)ion |

All acids react with hydrogen carbonate/carbonate to form salt ,water and evolve/produce bubbles of carbon(IV)oxide and water.

Carbon(IV)oxide forms a white precipitate when bubbled in lime water/extinguishes a burning splint.

Alkanoic acids react with hydrogen carbonate/carbonate to form alkanoates ,water and evolve/produce bubbles of carbon(IV)oxide and water.

Alkanoic acid + hydrogen carbonate -> alkanoate + water + carbon(IV)oxide

Alkanoic acid + carbonate -> alkanoate + water + carbon(IV)oxide

Examples

1. Sodium hydrogen carbonate reacts with ethanoic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.

CH3COO**H** (aq) + NaHCO3 (s) -> CH3COO**Na** (aq) + H2O(l) + CO2 (g)

(Ethanoic acid) (Sodium ethanoate)

2.Sodium carbonate reacts with ethanoic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.

2CH3COO**H** (aq) + Na2CO3 (s) -> 2CH3COO**Na** (aq) + H2O(l) + CO2 (g)

(Ethanoic acid) (Sodium ethanoate)

3.Sodium carbonate reacts with ethan-1,2-dioic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.

**H**OOC-COO**H**+ Na2CO3 (s) -> **Na**OOC - COO**Na** + H2O(l) + CO2 (g)

(ethan-1,2-dioic acid) (sodium ethan-1,2-dioate)

4.Sodium hydrogen carbonate reacts with ethan-1,2-dioic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.

**H**OOC-COO**H**+ 2NaHCO3 (s) -> **Na**OOC - COO**Na** + H2O(l) + 2CO2 (g)

(ethan-1,2-dioic acid) (sodium ethan-1,2-dioate)

**(e)Esterification**

Experiment

Place 4cm3 of ethanol acid in a boiling tube.

Add equal volume of ethanoic acid. To the mixture, add 2 drops of concentrated sulphuric(VI)acid **carefully**. Warm/heat gently on Bunsen flame.

Pour the mixture into a beaker containing 50cm3 of water. Smell the products. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

|  |  |
| --- | --- |
| Solution/acid | Observations |
| Ethanoic acid | Sweet fruity smell |
| Succinic acid | Sweet fruity smell |
| Citric acid | Sweet fruity smell |
| Oxalic acid | Sweet fruity smell |
| Tartaric acid | Sweet fruity smell |
| **Dilute sulphuric(VI)acid** | **No sweet fruity smell** |

Explanation

Alkanols react with alkanoic acid to form the sweet smelling homologous series of esters and water.The reaction is catalysed by concentrated sulphuric(VI)acid in the laboratory but naturally by sunlight /heat.Each ester has a characteristic smell derived from the many possible combinations of alkanols and alkanoic acids.

Alkanol + Alkanoic acids -> Ester + water

Esters derive their names from the alkanol first then alkanoic acids. The alkanol “becomes” an **alkyl** group and the alkanoic acid “becomes” **alkanoate** hence **alkylalkanoate**. e.g.

Ethanol + Ethanoic acid -> Ethylethanoate + Water

Ethanol + Propanoic acid -> Ethylpropanoate + Water

Ethanol + Methanoic acid -> Ethylmethanoate + Water

Ethanol + butanoic acid -> Ethylbutanoate + Water

Propanol + Ethanoic acid -> Propylethanoate + Water

Methanol + Ethanoic acid -> Methyethanoate + Water

Methanol + Decanoic acid -> Methyldecanoate + Water

Decanol + Methanoic acid -> Decylmethanoate + Water

During the formation of the ester, the “O” joining the alkanol and alkanoic acid comes from the alkanol.

R1 -CO**OH** + R2 –O**H** -> **R1 -COO –R2** + **H2O**

Examples

1. Ethanol reacts with ethanoic acid to form the ester ethyl ethanoate and water.

Ethanol + Ethanoic acid --Conc. H2SO4 -->Ethylethanoate + Water

C2H5OH (l) + CH3COOH(l) --Conc. H2SO4 --> CH3COO C2H5(**aq**)+H2O(l)

CH3CH2OH (l)+ CH3COOH(l) --Conc. H2SO4 --> CH3COOCH2CH3(**aq**)+H2O(l)

2. Ethanol reacts with propanoic acid to form the ester ethylpropanoate and water.

Ethanol + Propanoic acid --Conc. H2SO4 -->Ethylethanoate + Water

C2H5OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->CH3CH2COO C2H5(**aq**)+H2O(l)

CH3CH2OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->

CH3 CH2COOCH2CH3(**aq**)+H2O(l)

3. Methanol reacts with ethanoic acid to form the ester methyl ethanoate and water.

Methanol + Ethanoic acid --Conc. H2SO4 -->Methylethanoate + Water

CH3OH (l) + CH3COOH(l) --Conc. H2SO4 --> CH3COO CH3(**aq**)+H2O(l)

4. Methanol reacts with propanoic acid to form the ester methyl propanoate and water.

Methanol + propanoic acid --Conc. H2SO4 -->Methylpropanoate + Water

CH3OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 --> CH3 CH2COO CH3(**aq**)+H2O(l)

5. Propanol reacts with propanoic acid to form the ester propylpropanoate and water.

Propanol + Propanoic acid --Conc. H2SO4 -->Ethylethanoate + Water

C3H7OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->CH3CH2COO C3H7(**aq**)+H2O(l)

CH3CH2 CH2OH (l)+ CH3 CH2COOH(l) --Conc. H2SO4 -->

CH3 CH2COOCH2 CH2CH3(**aq**)+H2O(l)

**C. DETERGENTS**

Detergents are cleaning agents that improve the cleaning power /properties of water.A detergent therefore should be able to:

(i)dissolve substances which water can not e.g grease ,oil, fat

(ii)be washed away after cleaning.

There are two types of detergents:

(a)Soapy detergents

(b)Soapless detergents

1. **SOAPY DETERGENTS**

Soapy detergents usually called soap is long chain salt of organic alkanoic acids.Common soap is sodium octadecanoate .It is derived from reacting concentrated sodium hydroxide solution with octadecanoic acid(18 carbon alkanoic acid) i.e.

Sodium hydroxide + octadecanoic acid -> Sodium octadecanoate + water

**Na**OH(aq) + CH3 (CH2) 16 COOH(aq) -> CH3 (CH2) 16 COO – **Na**+ (aq) +H2 O(l)

Commonly ,soap can thus be represented ;

R-COO – **Na**+where;

R is a long chain alkyl group and -COO – **Na**+ is the alkanoate ion.

In a school laboratory and at industrial and domestic level,soap is made by reacting concentrated sodium hydroxide solution with esters from (animal) **fat** and **oil**. The process of making soap is called **saponification.** During saponification ,the ester is **hydrolyzed** by the alkali to form sodium salt /soap and **glycerol/propan-1,2,3-triol** is produced.

Fat/oil(ester)+sodium/potassium hydroxide->sodium/potassium salt(soap)+ glycerol

Fats/Oils are esters with fatty acids and glycerol parts in their structure;

C17H35COO**CH2**

C17H35COO**CH**

C17H35COO**CH2**

When boiled with concentrated sodium hydroxide solution NaOH;

(i)NaOH ionizes/dissociates into **Na**+ and **OH**- ions

(ii)fat/oil split into **three** C17H35COO- and **one** CH2 CH CH2

(iii) the three **Na**+ combine with the three C17H35COO- to form the salt C17H35COO- **Na**+

(iv)the three **OH**-ions combine with the CH2 CH CH2 to form an alkanol with three functional groups CH2 **OH** CH **OH** CH2 **OH(propan-1,2,3-triol)**

C17H35COO**CH2 CH2OH**

C17H35COO**CH +NaOH -> 3** C17H35COO- **Na**+ **+ CHOH**

C17H35COO**CH2 CH2OH**

**Ester Alkali Soap glycerol**

Generally:

CnH2n+1COO**CH2 CH2OH**

CnH2n+1COO**CH +NaOH -> 3** CnH2n+1COO- **Na**+ **+ CHOH**

CnH2n+1COO**CH2 CH2OH**

**Ester Alkali Soap glycerol**

R - COO**CH2 CH2OH**

R - COO**CH +NaOH -> 3**R-COO- **Na**+ **+ CHOH**

R- COO**CH2 CH2OH**

**Ester Alkali Soap glycerol**

During this process a little sodium chloride is added to **precipitate** the soap by reducing its solubility. This is called **salting out**.

The soap is then added colouring agents ,perfumes and herbs of choice.

**School laboratory preparation of soap**

Place about 40 g of fatty (animal fat)beef/meat in 100cm3 beaker .Add about 15cm3 of 4.0M sodium hydroxide solution. Boil the mixture for about 15minutes.Stir the mixture .Add about 5.0cm3 of distilled water as you boil to make up for evaporation. Boil for about another 15minutes.Add about four spatula end full of pure sodium chloride crystals. Continue stirring for another five minutes. Allow to cool. Filter of /decant and wash off the residue with distilled water .Transfer the clean residue into a dry beaker. Preserve.

**The action of soap**

Soapy detergents:

(i)act by reducing the surface tension of water by forming a thin layer on top of the water.

(ii)is made of a **non-polar** alkyl /hydrocarbon tail and a **polar**  -COO-Na+ head. The non-polar alkyl /hydrocarbon tail is **hydrophobic** (water hating) and thus does not dissolve in water .It dissolves in non-polar solvent like grease, oil and fat. The polar -COO-Na+ head is **hydrophilic** (water loving)and thus dissolve in water. When washing with soapy detergent, the non-polar tail of the soapy detergent surround/dissolve in the dirt on the garment /grease/oil while the polar head dissolve in water.

Through **mechanical agitation**/stirring/sqeezing/rubbing/beating/kneading, some grease is dislodged/lifted of the surface of the garment. It is immediately surrounded by more soap molecules It float and spread in the water as tiny droplets that scatter light in form of emulsion making the water cloudy and shinny. It is removed from the garment by rinsing with fresh water.The repulsion of the soap head prevent /ensure the droplets do not mix.Once removed, the dirt molecules cannot be redeposited back because it is surrounded by soap molecules.

**Advantages and disadvantages of using soapy detergents**

Soapy detergents are biodegradable. They are acted upon by bacteria and rot.They thus do not cause environmental pollution.

Soapy detergents have the diadvatage in that:

(i)they are made from fat and oils which are better eaten as food than make soap.

(ii)forms an insoluble precipitate with hard water called **scum.** Scum is insoluble calcium octadecanoate and Magnesium octadecanoate formed when soap reacts with Ca2+ and Mg2+ present in hard water.

Chemical equation

2C17H35COO- **Na**+ (aq)**+** Ca2+(aq) -> **(**C17H35COO- **)**Ca2+ (s) + **2Na**+(aq)

(insoluble Calcium octadecanote/scum)

2C17H35COO- **Na**+ (aq)**+** Mg2+(aq) -> **(**C17H35COO- **)**Mg2+ (s) + **2Na**+(aq)

(insoluble Magnesium octadecanote/scum)

This causes wastage of soap.

Potassium soaps are better than Sodium soap. Potassium is more expensive than sodium and thus its soap is also more expensive.

**(b)SOAPLESS DETERGENTS**

Soapless detergent usually called detergent is a long chain salt fromed from by-products of fractional distillation of crude oil.Commonly used soaps include:

(i)washing agents

(ii)toothpaste

(iii)emulsifiers/wetting agents/shampoo

Soapless detergents are derived from reacting:

(i)concentrated sulphuric(VI)acid with a long chain alkanol e.g. Octadecanol(18 carbon alkanol) to form alkyl hydrogen sulphate(VI)

Alkanol + Conc sulphuric(VI)acid -> alkyl hydrogen sulphate(VI) + Water

R –O**H** + **H**2SO4 -> R –O-SO3H + **H2**O

(ii)the alkyl hydrogen sulphate(VI) is then neutralized with sodium/potassium hydroxide to form sodium/potassium alkyl hydrogen sulphate(VI)

Sodium/potassium alkyl hydrogen sulphate(VI) is the soapless detergent.

alkyl hydrogen + Potassium/sodium -> Sodium/potassium + Water

sulphate(VI) hydroxide alkyl hydrogen sulphate(VI)

R –O-SO3H + **Na**OH -> R –O-SO3- Na+ + **H2**O

Example

Step I : Reaction of Octadecanol with Conc.**H**2SO4

C17H35CH2OH(aq) + H2SO4 -> C17H35CH2-**O- SO3- H+**(aq) + H2O(l)

octadecanol + sulphuric(VI)acid -> Octadecyl hydrogen sulphate(VI) + water

Step II: Neutralization by an alkali

C17H35CH2-**O- SO3- H+**(aq) + NaOH-> C17H35CH2-**O- SO3- Na+**(aq) + H2O(l)

Octadecyl hydrogen + sodium/potassium -> sodium/potassium octadecyl+Water

sulphate(VI) hydroxide hydrogen sulphate(VI)

**School laboratory preparation of soapless detergent**

Place about 20g of olive oil in a 100cm3 beaker. Put it in a trough containing ice cold water.

Add dropwise carefully 18M concentrated sulphuric(VI)acid stirring continuously into the olive oil until the oil turns brown.Add 30cm3 of 6M sodium hydroxide solution.Stir.This is a soapless detergent.

**The action of soapless detergents**

The action of soapless detergents is similar to that of soapy detergents.The soapless detergents contain the hydrophilic head and a long hydrophobic tail. i.e.

vvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvv-COO-Na+

vvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvv-O-SO3- Na+

(long hydrophobic /non-polar alkyl tail) (hydrophilic/polar/ionic head)

The tail dissolves in fat/grease/oil while the ionic/polar/ionic head dissolves in water.

The tail stick to the dirt which is removed by the attraction of water molecules and the polar/ionic/hydrophilic head by mechanical agitation /squeezing/kneading/ beating/rubbing/scrubbing/scatching.

The suspended dirt is then surrounded by detergent molecules and repulsion of the anion head preventing the dirt from sticking on the material garment.

The tiny droplets of dirt emulsion makes the water cloudy. On rinsing the cloudy emulsion is washed away.

**Advantages and disadvantages of using soapless detergents**

Soapless detergents are non-biodegradable unlike soapy detergents.

They persist in water during sewage treatment by causing foaming in rivers ,lakes and streams leading to marine /aquatic death.

Soapless detergents have the advantage in that they:

(i)do not form scum with hard water.

(ii)are cheap to manufacture/buying

(iii)are made from petroleum products but soapis made from fats/oil for human consumption.

**Sample revision questions**

1. Study the scheme below

KOH

Fat/oil

Residue X

Filtrate Y

Filtration

Sodium Chloride

Boiling

**(a)Identify the process**

Saponification

**(b)Fats and oils are esters. Write the formula of the a common structure of ester**

C17H35COO**CH2**

C17H35COO**CH**

C17H35COO**CH2**

**(c)Write a balanced equation for the reaction taking place during boiling**

C17H35COOCH2 CH2OH

C17H35COOCH +3NaOH -> 3 C17H35COO- Na+ + CHOH

C17H35COOCH2 CH2OH

Ester Alkali Soap glycerol

**(d)Give the IUPAC name of:**

**(i)Residue X**

Potassium octadecanoate

**(ii)Filtrate Y**

Propan-1,2,3-triol

**(e)Give one use of fitrate Y**

Making paint

**(f)What is the function of sodium chloride**

To reduce the solubility of the soap hence helping in precipitating it out

**(g)Explain how residue X helps in washing.**

Has a non-polar hydrophobic tail that dissolves in dirt/grease /oil/fat

Has a polar /ionic hydrophilic head that dissolves in water.

From mechanical agitation,the dirt is plucked out of the garment and surrounded by the tail end preventing it from being deposited back on the garment.

**(h)State one:**

**(i)advantage of continued use of residue X on the environment**

Is biodegradable and thus do not pollute the environment

**(ii)disadvantage of using residue X**

Uses fat/oil during preparation/manufacture which are better used for human consumption.

**(i)Residue X was added dropwise to some water.The number of drops used before lather forms is as in the table below.**

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Water sample** | | |
| **A** | **B** | **C** |
| **Drops of residue X** | **15** | **2** | **15** |
| **Drops of residue X in boiled water** | **2** | **2** | **15** |

**(i)State and explain which sample of water is:**

**I. Soft**

Sample B .Very little soap is used and no effect on amount of soap even on boiling/heating.

**II. Permanent hard**

Sample C . A lot of soap is used and no effect on amount of soap even on boiling/heating. Boiling does not remove permanent hardness of water.

**III. Temporary hard**

Sample A . A lot of soap is used before boiling. Very little soap is used on boiling/heating. Boiling remove temporary hardness of water.

**(ii)Write the equation for the reaction at water sample C.**

Chemical equation

2C17H35COO- **K**+ (aq)**+** CaSO4(aq) -> **(**C17H35COO- **)**Ca2+ (s) + **K**2SO4(aq)

(insoluble Calcium octadecanote/scum)

Ionic equation

2C17H35COO- **K**+ (aq)**+** Ca2+(aq) -> **(**C17H35COO- **)**Ca2+ (s) + **2K**+(aq)

(insoluble Calcium octadecanote/scum)

Chemical equation

2C17H35COO- **K**+ (aq)**+** MgSO4(aq) -> **(**C17H35COO- **)**Mg2+ (s) + **K**2SO4(aq)

(insoluble Calcium octadecanote/scum)

Ionic equation

2C17H35COO- **K**+ (aq)**+** Mg2+(aq) -> **(**C17H35COO- **)**Mg2+ (s) + **2K**+(aq)

(insoluble Magnesium octadecanote/scum)

**(iii)Write the equation for the reaction at water sample A before boiling.**

Chemical equation

2C17H35COO- **K**+ (aq)**+** Ca(HCO3)(aq) ->**(**C17H35COO- **)**Ca2+ (s) +2KHCO3 (aq)

(insoluble Calcium octadecanote/scum)

Ionic equation

2C17H35COO- **K**+ (aq)**+** Ca2+(aq) -> **(**C17H35COO- **)**Ca2+ (s) +2K+(aq)

(insoluble Calcium octadecanote/scum)

Chemical equation

2C17H35COO- **K**+ (aq)**+** Mg(HCO3)(aq) ->**(**C17H35COO- **)**Mg2+ (s) +2KHCO3 (aq)

(insoluble Calcium octadecanote/scum)

Ionic equation

2C17H35COO- **K**+ (aq)**+** Mg2+(aq) -> **(**C17H35COO- **)**Mg2+ (s) +2K+(aq)

(insoluble Magnesium octadecanote/scum)

**(iv)Explain how water becomes hard**

Natural or rain water flowing /passing through rocks containing calcium (chalk, gypsum, limestone)and magnesium compounds (dolomite)dissolve them to form soluble Ca2+  and Mg2+ ions that causes water hardness.

**(v)State two useful benefits of hard water**

-Used in bone and teeth formation

-Coral polyps use hard water to form coral reefs

-Snails use hard water to make their shells

**2.Study the scheme below and use it to answer the questions that follow.**

Conc. H2SO4

Substance B

6M sodium hydroxide

Brown solid A

Ice cold water

Olive oil

**(a)Identify :**

**(i)brown solid A**

Alkyl hydrogen sulphate(VI)

**(ii)substance B**

Sodium alkyl hydrogen sulphate(VI)

**(b)Write a general formula of:**

**(i)Substance A.**

**O**

R-O-S O3 **H** //R- O - S - O  **- H**

**O**

**(ii)Substance B O**

R-O-S O3**- Na+** R- O - S - O **- Na+**

**O**

**(c)State one**

**(i) advantage of continued use of substance B**

**-**Does not form scum with hard water

-Is cheap to make

-Does not use food for human as a raw material.

**(ii)disadvantage of continued use of substance B.**

Is non-biodegradable therefore do not pollute the environment

**(d)Explain the action of B during washing.**

Has a non-polar hydrocarbon long tail that dissolves in dirt/grease/oil/fat.

Has a polar/ionic hydrophilic head that dissolves in water

Through mechanical agitation the dirt is plucked /removed from the garment and surrounded by the tail end preventing it from being deposited back on the garment.

**(e) Ethene was substituted for olive oil in the above process. Write the equation and name of the new products A and B.**

Product A

Ethene + Sulphuric(VI)acid -> Ethyl hydrogen sulphate(VI)

H2C=CH2 + H2SO4 –> H3C – CH2 –O-SO3H

Product B

Ethyl hydrogen sulphate(VI) + sodium hydroxide -> sodium Ethyl + Water

hydrogen sulphate(VI)

H3C – CH2 –O-SO3H + **Na**OH -> H3C – CH2 –O-SO3-**Na+**+ H2O

(f)Ethanol can also undergo similar reactions forming new products A and B.Show this using a chemical equation.

Product A

Ethanol + Sulphuric(VI)acid ->Ethyl hydrogen sulphate(VI) + water

H3C-CH2OH + H2SO4 –> H3C – CH2 –O-SO3H + H2O

Product B

Ethyl hydrogen sulphate(VI) + sodium hydroxide -> sodium Ethyl + Water

hydrogen sulphate(VI)

H3C – CH2 –O-SO3H + **Na**OH -> H3C – CH2 –O-SO3-**Na+**+ H2O

**3.Below is part of a detergent**

**H3C – (CH2 )16 – O - SO3 - K +**

**(a)Write the formular of the polar and non-polar end**

Polar end

H3C – (CH2 )16 –

Non-polar end

– O - SO3 - K +

**(b)Is the molecule a soapy or saopless detergent?**

Soapless detergent

**(c)State one advantage of using the above detergent**

-does not form scum with hard water

-is cheap to manufacture

**D. POLYMERS AND FIBRES**

Polymers and fibres are giant molecules of organic compounds. Polymers and fibres are formed when **small** molecules called monomers join together to form **large** molecules called polymers at high temperatures and pressures. This process is called polymerization.

Polymers and fibres are either:

(a)**Natural** polymers and fibres

(b)**Synthetic** polymers and fibres

Natural polymers and fibres are found in living things(plants and animals) Natural polymers/fibres include:

-proteins/polypeptides making amino acids in animals

-cellulose that make cotton,wool,paper and silk

-Starch that come from glucose

-Fats and oils

-Rubber from latex in rubber trees.

Synthetic polymers and fibres are man-made. They include:

-polyethene

-polychloroethene

-polyphenylethene(polystyrene)

-Terylene(Dacron)

-Nylon-6,6

-Perspex(artificial glass)

Synthetic polymers and fibres have the following characteristic advantages over natural polymers

1. They are light and portable

2. They are easy to manufacture.

3. They can easily be molded into shape of choice.

4. They are resistant to corrosion, water, air , acids, bases and salts.

5. They are comparatively cheap, affordable, colourful and aesthetic

Synthetic polymers and fibres however have the following disadvantages over natural polymers

1. They are non-biodegradable and hence cause environmental pollution during disposal
2. They give out highly poisonous gases when burnt like chlorine/carbon(II)oxide
3. Some on burning produce Carbon(IV)oxide. Carbon(IV)oxide is a green house gas that cause global warming.
4. Compared to some metals, they are poor conductors of heat,electricity and have lower tensile strength.

To reduce environmental pollution from synthetic polymers and fibres, the followitn methods of disposal should be used:

1.Recycling: Once produced all synthetic polymers and fibres should be recycled to a new product. This prevents accumulation of the synthetic polymers and fibres in the environment.

2.Production of biodegradable synthetic polymers and fibres that **rot** away.

There are two types of polymerization:

(a)addition polymerization

(b)condensation polymerization

**(a)addition polymerization**

Addition polymerization is the process where a small unsaturated monomer (alkene ) molecule join together to form a large saturated molecule. Only alkenes undergo addition polymerization.

Addition polymers are named from the alkene/monomer making the polymer and adding the prefix “**poly**” before the name of monomer to form a **polyalkene**

During addition polymerization

(i)the double bond in alkenes break

(ii)free radicals are formed

(iii)the free radicals collide with each other and join to form a larger molecule. The more collisions the larger the molecule.

**Examples of addition polymerization**

1.Formation of Polyethene

Polyethene is an addition polymer formed when ethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting paticles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H H H H H H H H

Ethene + Ethene + Ethene + Ethene + **…**

(ii)the double bond joining the ethane molecule break to free readicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H H H H H H H H

Ethene radical + Ethene radical + Ethene radical + Ethene radical + **…**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

•C – C - C – C - C – C - C - C• + **…**

H H H H H H H H

Lone pair of electrons can be used to join more monomers to form longer polyethene.

Polyethene molecule can be represented as:

H H H H H H H H extension of

molecule/polymer

- C – C - C – C - C – C - C – C- + **…**

H H H H H H H H

Since the molecule is a repetition of one monomer, then the polymer is:

H H

( C – C )**n**

H H

Where n is the number of monomers in the polymer. The number of monomers in the polymer can be determined from the molar mass of the polymer and monomer from the relationship:

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

Examples

**Polythene has a molar mass of 4760.Calculate the number of ethene molecules in the polymer(C=12.0, H=1.0 )**

Number of monomers/repeating units in polyomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C2H4 )= 28 Molar mass polyethene = 4760

Substituting 4760 = 170 ethene molecules

28

The **commercial** name of polyethene is **polythene**. It is an elastic, tough, transparent and durable plastic. Polythene is used:

(i)in making plastic bag

(ii)bowls and plastic bags

(iii)packaging materials

2.Formation of Polychlorethene

Polychloroethene is an addition polymer formed when chloroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

chloroethene + chloroethene + chloroethene + chloroethene + **…**

(ii)the double bond joining the chloroethene molecule break to free radicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

•C – C - C – C - C – C - C - C• + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

Lone pair of electrons can be used to join more monomers to form longer polychloroethene.

Polychloroethene molecule can be represented as:

H H H H H H H H extension of

molecule/polymer

- C – C - C – C - C – C - C – C- + **…**

H **Cl** H **Cl** H **Cl** H **Cl**

Since the molecule is a repetition of one monomer, then the polymer is:

H H

( C – C )**n**

H **Cl**

Examples

**Polychlorothene has a molar mass of 4760.Calculate the number of chlorethene molecules in the polymer(C=12.0, H=1.0,Cl=35.5 )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C2H3Cl )= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**

62.5

The **commercial** name of polychloroethene is **polyvinylchloride(PVC)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope

(ii)water pipes

(iii)crates and boxes

3.Formation of Polyphenylethene

Polyphenylethene is an addition polymer formed when phenylethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H **C6H5** H **C6H5** H **C6H5** H **C6H5**

phenylethene + phenylethene + phenylethene + phenylethene + **…**

(ii)the double bond joining the phenylethene molecule break to free radicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H **C6H5** H **C6H5** H **C6H5** H **C6H5**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

• C – C - C – C - C – C - C - C • + **…**

H **C6H5**  H **C6H5** H **C6H5** H **C6H5**

Lone pair of electrons can be used to join more monomers to form longer polyphenylethene.

Polyphenylethene molecule can be represented as:

H H H H H H H H

- C – C - C – C - C – C - C - C -

H **C6H5**  H **C6H5** H **C6H5** H **C6H5**

Since the molecule is a repetition of one monomer, then the polymer is:

H H

( C – C )**n**

H **C6H5**

Examples

**Polyphenylthene has a molar mass of 4760.Calculate the number of phenylethene molecules in the polymer(C=12.0, H=1.0, )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C8H8 )= 104 Molar mass polyethene = 4760

Substituting 4760 = 45.7692 =>45 polyphenylethene molecules**(whole number)**

104

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

(iii)clothe linings

4.Formation of Polypropene

Polypropene is an addition polymer formed when propene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

H H H H H H H H

C = C + C = C + C = C + C = C + **…**

H **CH3** H **CH3** H **CH3** H **CH3**

propene + propene + propene + propene + **…**

(ii)the double bond joining the phenylethene molecule break to free radicals

H H H H H H H H

•C – C• + •C - C• + •C - C• + •C - C• + **…**

H **CH3** H **CH3** H **CH3** H **CH3**

(iii)the free radicals collide with each other and join to form a larger molecule

H H H H H H H H lone pair of electrons

• C – C - C – C - C – C - C - C • + **…**

H **CH3**  H **CH3** H **CH3** H **CH3**

Lone pair of electrons can be used to join more monomers to form longer propene.

propene molecule can be represented as:

H H H H H H H H

- C – C - C – C - C – C - C - C -

H **CH3**  H **CH3** H **CH3** H **CH3**

Since the molecule is a repetition of one monomer, then the polymer is:

H H

( C – C )**n**

H **CH3**

Examples

**Polypropene has a molar mass of 4760.Calculate the number of propene molecules in the polymer(C=12.0, H=1.0, )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass propene (C3H8 )= 44 Molar mass polyethene = 4760

Substituting 4760 = 108.1818 =>108 propene molecules**(whole number)**

44

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

(iii)clothe linings

5.Formation of Polytetrafluorothene

Polytetrafluorothene is an addition polymer formed when tetrafluoroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

F F F F F F F F

C = C + C = C + C = C + C = C + **…**

F F F F F F F F

tetrafluoroethene + tetrafluoroethene+ tetrafluoroethene+ tetrafluoroethene + **…**

(ii)the double bond joining the tetrafluoroethene molecule break to free radicals

F F F F F F F F

•C – C• + •C - C• + •C - C• + •C - C• + **…**

F F F F F F F F

(iii)the free radicals collide with each other and join to form a larger molecule

F F F F F F F F lone pair of electrons

•C – C - C – C - C – C - C - C• + **…**

F F F FF F F F

Lone pair of electrons can be used to join more monomers to form longer polytetrafluoroethene.

polytetrafluoroethene molecule can be represented as:

F F F F F F F F extension of

molecule/polymer

- C – C - C – C - C – C - C – C- + **…**

F F F F F F F F

Since the molecule is a repetition of one monomer, then the polymer is:

F F

( C – C )**n**

F F

Examples

**Polytetrafluorothene has a molar mass of 4760.Calculate the number of tetrafluoroethene molecules in the polymer(C=12.0, ,F=19 )**

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C2F4 )= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**

62.5

The **commercial** name of polytetrafluorethene(**P.T.F.E**) is **Teflon(P.T.F.E)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope

(ii)water pipes

(iii)crates and boxes

5.Formation of rubber from Latex

Natural rubber is obtained from rubber trees.

During harvesting an incision is made on the rubber tree to produce a milky white substance called **latex.**

Latex is a mixture of rubber and lots of water.

The latex is then added an acid to coagulate the rubber.

Natural rubber is a polymer of 2-methylbut-1,3-diene ;

H CH3 H H

CH2=C (CH3) CH = CH2 H - C  **=** C – C **=** C - H

During natural polymerization to rubber, one double C=C bond break to self add to another molecule.The double bond remaining move to carbon “2” thus;

H CH3 H H H CH3 H H

- C - C  **=** C  **-** C - C - C **=** C **-** C -

H H H H

Generally the structure of rubber is thus;

H CH3 H H

-(- C - C **=** C - C -)**n**-

H H

Pure rubber is soft and sticky.It is used to make erasers, car tyres. Most of it is vulcanized.Vulcanization is the process of heating rubber with sulphur to make it harder/tougher.

During vulcanization the sulphur atoms form a cross link between chains of rubber molecules/polymers. This decreases the number of C=C double bonds in the polymer.

H CH3 H H H CH3 H H

Sulphur atoms make cross link between polymers

- C - C  **-** C  **-** C - C - C **-** C **-** C -

H S H H S H

H CH3 S H H CH3 S H

- C - C  **-**  C  **-** C - C - C **-** C **-** C -

H H H H H H

Vulcanized rubber is used to make **tyres**, **shoes** and **valves**.

6.Formation of synthetic rubber

Synthetic rubber is able to resist action of oil,abrasion and organic solvents which rubber cannot.

Common synthetic rubber is a polymer of 2-chlorobut-1,3-diene ;

H Cl H H

CH2=C (Cl CH = CH2 H - C  **=** C – C **=** C - H

During polymerization to synthetic rubber, one double C=C bond is broken to self add to another molecule. The double bond remaining move to carbon “2” thus;

H Cl H H H Cl H H

- C - C  **=** C  **-** C - C - C **=** C **-** C -

H H H H

Generally the structure of rubber is thus;

H Cl H H

-(- C - C **=** C - C -)**n**-

H H

Rubber is thus strengthened through vulcanization and manufacture of synthetic rubber.

**(b)Condensation polymerization**

Condensation polymerization is the process where two or more small monomers join together to form a larger molecule by elimination/removal of a simple molecule. (usually water).

Condensation polymers acquire a different name from the monomers because the two monomers are two different compounds

During condensation polymerization:

(i)the two monomers are brought together by high pressure to reduce distance between them.

(ii)monomers realign themselves at the functional group.

(iii)from each functional group an element is removed so as to form simple molecule (of usually H2O/HCl)

(iv)the two monomers join without the simple molecule of H2O/HCl

**Examples of condensation polymerization**

1.Formation of Nylon-6,6

**Method 1**: Nylon-6,6 can be made from the condensation polymerization of hexan-1,6-dioic acid with hexan-1,6-diamine.Amines are a group of homologous series with a general formula R-NH2 and thus -NH2 as the functionalgroup.

During the formation of Nylon-6,6:

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.

O O H H

H- O - C – (CH2 ) 4 – C – O - H + H –N – (CH2) 6 – N – H

(iii)from each functional group an element is removed so as to form a molecule of H2O and the two monomers join at the linkage .

O O H H

H- O - C – (CH2 ) 4 – C **–** N – (CH2) 6 – N – H + H 2O

.

Polymer bond linkage

Nylon-6,6 derive its name from the two monomers each with **six** carbon chain

**Method 2**: Nylon-6,6 can be made from the condensation polymerization of hexan-1,6-dioyl dichloride with hexan-1,6-diamine.

Hexan-1,6-dioyl dichloride belong to a group of homologous series with a general formula R-OCland thus -OCl as the functionalgroup.

The R-OCl is formed when the “OH” in R-O**OH**/alkanoic acid is replaced by Cl/chlorine/Halogen

During the formation of Nylon-6,6:

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.

O O H H

Cl - C – (CH2 ) 4 – C – Cl + H –N – (CH2) 6 – N – H

(iii)from each functional group an element is removed so as to form a molecule of HCl and the two monomers join at the linkage .

O O H H

Cl - C – (CH2 ) 4 – C **–** N – (CH2) 6 – N – H + HCl

.

Polymer bond linkage

The two monomers each has **six** carbon chain hence the name “nylon-6,6”

The commercial name of Nylon-6,6 is **Nylon** It is a a tough, elastic and durable plastic. It is used to make **clothes**, **plastic ropes** and **carpets.**

2.Formation of Terylene

**Method 1**: Terylene can be made from the condensation polymerization of ethan-1,2-diol with benzene-1,4-dicarboxylic acid.

Benzene-1,4-dicarboxylic acid a group of homologous series with a general formula R-COOHwhere R is a ring of six carbon atom called Benzene ring .The functionalgroup is -COOH.

During the formation of Terylene:

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.

O O

H- O - C – C6H5 – C – O - H + H –O – CH2 CH2 – O – H

(iii)from each functional group an element is removed so as to form a molecule of H2O and the two monomers join at the linkage .

O O

H- O - C – C6H5  – C **–** O – (CH2) 6 – N – H + H 2O

.

Polymer bond linkage of terylene

**Method 2**: Terylene can be made from the condensation polymerization of benzene-1,4-dioyl dichloride with ethan-1,2-diol.

Benzene-1,4-dioyl dichloride belong to a group of homologous series with a general formula R-OCland thus -OCl as the functionalgroup and R as a benzene ring.

The R-OCl is formed when the “OH” in R-O**OH** is replaced by Cl/chlorine/Halogen

During the formation of Terylene

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.

O O

Cl - C – C5H5 – C – Cl + H –O – CH2 CH2 – O - H

(iii)from each functional group an element is removed so as to form a molecule of HCl and the two monomers join at the linkage .

O O

Cl - C – C5H5  – C **–** O – CH2 CH2 – O – H + HCl

.

Polymer bond linkage of terylene

The commercial name of terylene is **Polyester /polyster** It is a a tough, elastic and durable plastic. It is used to make **clothes**, **plastic ropes and sails** and  **plastic model kits.**

ENERGY CHANGES

**1.Introduction to Energy changes**

Energy is the capacity to do **work**. There are many/various **forms** of energy like heat, electric, mechanical, and/ or chemical energy.There are two types of energy:

(i)Kinetic Energy(KE) ;the energy in motion.

(ii)Potential Energy(PE); the stored/internal energy.

Energy like matter , is **neither created nor destroyed** but can be transformed /changed from one form to the other/ is interconvertible. This is the **princ]iple of conservation** of energy. e.g. Electrical energy into heat through a filament in bulb. Chemical and physical processes take place with **absorption** or **evolution**/production of energy mainly in form of **heat**

The study of energy changes that accompany physical/chemical reaction/changes is called **Thermochemistry.** Physical/chemical reaction/changes that involve energy changes are called **thermochemical reactions.** The SI unit of energy is the **Joule**(**J**).Kilo Joules(**kJ**)and megaJoules(**MJ**) are also used. The Joule(J) is defined as the:

1. quantity of energy transferred when a force of one **newton** acts through a distance of one **metre**.
2. quantity of energy transferred when one **coulomb** of electric charge is passed through a potential difference of one **volt**.

All thermochemical reactions should be carried out at standard conditions of:

(i) **298K /25oC** temperature

(ii)**101300Pa/101300N/m2 /760mmHg/1** atmosphere pressure.

**2.Exothermic and endothermic processes/reactions**

Some reactions **/** processes take place with evolution/production of energy. They are said to be **exothermic** while others take place with **absorption** of energy. They are said to be **endothermic.**

Practically exothermic reactions **/** processes cause a rise in temperature (by a rise in thermometer reading/mercury or alcohol level rise)

Practically endothermic reactions **/** processes cause a fall in temperature (by a fall in thermometer reading/mercury or alcohol level decrease)

**To demonstrate/illustrate exothermic and endothermic processes/reactions**

a) Dissolving Potassium nitrate(V)/ammonium chloride crystals

**Procedure:**

Measure 20cm3 of water in a beaker. Determine and record its temperature T1.Put about 1.0g of Potassium nitrate(V) crystals into the beaker. Stir the mixture carefully and note the highest temperature rise /fall T2.Repeat the whole procedure by using ammonium chloride in place of Potassium nitrate (V) crystals.

Sample results

|  |  |  |
| --- | --- | --- |
| **Temperture (oC)** | **Using Potassium nitrate(V) crystals** | **Using Ammonium**  **chloride crystals** |
| T2(Final temperature) | 21.0 | 23.0 |
| T1 (Initial temperature) | 25.0 | 26.0 |
| Change in temperature(T2 –T1) | **4.0** | **3.0** |

Note:

(i)Initial(**T1**) temperature of dissolution of both potassium nitrate(V) crystals and ammonium chloride crystals is **higher** than the final temperature(**T2**) (ii) Change in temperature(T2 –T1) is **not** a mathematical “-**4.0”** or“-**3.0”**. (iii)Dissolution of both potassium nitrate(V) and ammonium chloride crystals isan **endothermic** process becauseinitial(**T1**) temperatureis **higher** than the final temperature(**T2**) thus causes a **fall/drop** in temperature.

b) Dissolving concentrated sulphuric(VI) acid/sodium hydroxide crystals

**Procedure:**

Measure 20cm3 of water in a beaker. Determine and record its temperature T1.**Carefully** put about 1.0g/four pellets of sodium hydroxide crystals into the beaker. Stir the mixture carefully and note the highest temperature rise /fall T2.Repeat the whole procedure by using 2cm3 of concentrated sulphuric(VI) acid in place of sodium hydroxide crystals.

**CAUTION:**

(i)Sodium hydroxide crystals are **caustic** and cause painful blisters on contact with skin.

(ii) Concentrated sulphuric (VI) acid is **corrosive** and cause painful wounds on contact with skin.

Sample results

|  |  |  |
| --- | --- | --- |
| **Temperture (oC)** | **Using Sodium hydroxide pellets** | **Using Concentrated sulphuric(VI) acid** |
| T2(Final temperature) | 30.0 | 32.0 |
| T1 (Initial temperature) | 24.0 | 25.0 |
| Change in temperature(T2 –T1) | **6.0** | **7.0** |

Note:

(i)Initial (**T1**) temperature of dissolution of both concentrated sulphuric (VI) acid and sodium hydroxide pellets is **lower** than the final temperature (**T2**). (ii)Dissolution of both Sodium hydroxide pellets and concentrated sulphuric (VI) acid isan **exothermic** process becausefinal (**T2**) temperatureis **higher** than the initial temperature (**T1**) thus causes a **rise** in temperature.

The above reactions show heat loss **to** and heat gain **from** the surrounding as illustrated by a **rise** and **fall** in temperature/thermometer readings.

Dissolving both potassium nitrate(V) and ammonium chloride crystals causes heat **gain** **from** the **surrounding** that causes **fall** in thermometer reading.

Dissolving both Sodium hydroxide pellets and concentrated sulphuric (VI) acid causes heat **loss to** the **surrounding** that causes **rise** in thermometer reading.

At the same temperature and pressure ,heat absorbed and released is called enthalpy/ heat content denoted **H.**

Energy change is measured from the heat content/enthalpy of the **final** and **initial** products. It is denoted **∆H(**delta H**)**.i.e.

Enthalpy/energy/ change in heat content ∆H = H**final** – H**initial**

For chemical reactions:

∆H = H**products** – H**reactants**

For exothermic reactions, the heat contents of the **reactants** is **more** than/**higher** than the heat contents of **products**, therefore the ∆H is negative (**-∆H**) For endothermic reactions, the heat contents of the **reactants** is **less** than/**lower** than the heat contents of **products**, therefore the ∆H is negative (**+∆H**)

Graphically, in a **sketch** energy level diagram:

(i)For endothermic reactions the heat content of the reactants should be relatively/slightly **lower** than the heat content of the products

(ii)For exothermic reactions the heat content of the reactants should be relatively/slightly **higher** than the heat content of the products

Sketch energy level diagrams for endothermic dissolution

Energy

)

kJ

(

H

2

KNO

3

(

aq

)

**+**

∆

H = H

**2**

–

H

**1**

H

1

KNO

3

)

(

s

Reaction path/coordinate/progress

Ske

tch energy level diagrams for exothermic dissolution

H

2

NaOH

)

s

(

Energy(kJ)

-

∆

H = H

**2**

–

H

**1**

H

1

NaOH (aq)

Reaction path/coordinate/progress

|  |  |
| --- | --- |
|  |  |
|  |  |

H

2

H

2

SO

4

)

l

(

Energy

(

kJ

)

-

∆

H = H

**2**

–

H

**1**

H

1

H

2

SO

4

(

aq

)

Reaction

path/coordinate/progress

**3.Energy changes in physical processes**

Energy

)

kJ

(

H

2

NH

4

Cl

)

(

aq

**+**

∆

H = H

**2**

–

H

**1**

H

1

NH

4

Cl

)

(

s

Reaction

path/coordinate/progress

**Melting/freezing/fusion/solidification** and **boiling/vaporization/evaporation** are the two physical processes. Melting /freezing point of pure substances is fixed /constant. The boiling point of pure substance depend on **external** atmospheric **pressure**.

Melting/fusion is the physical change of a **solid** to **liquid**. Freezing is the physical change of a **liquid** to **solid**.

Melting/freezing/fusion/solidification are therefore two **opposite** but **same** reversible physical processes. i.e

1. (**s**) ========A(**l**)

Boiling/vaporization/evaporation is the physical change of a **liquid** to **gas/vapour**.

Condensation/liquidification is the physical change of **gas/vapour** to **liquid**. Boiling/vaporization/evaporation and condensation/liquidification are therefore two **opposite** but **same** reversible physical processes. i.e

1. (**l**) ========B(**g**)

Practically

(i) Melting/liquidification/fusion involves **heating** a solid to **weaken** the strong bonds holding the solid particles together. Solids are made up of very strong bonds holding the particles **very close** to each other (**Kinetic Theory of matter**).On heating these particles gain energy/heat from the surrounding heat source to form a liquid with **weaker** bonds holding the particles close together but with some degree of **freedom**. Melting/freezing/fusion is an **endothermic** (**+**∆H)process that require/absorb energy from the surrounding.

(ii)Freezing/fusion/solidification involves cooling a a liquid to reform /rejoin the very strong bonds to hold the particles **very close** to each other as solid and thus lose their degree of **freedom** (**Kinetic Theory of matter**). Freezing /fusion / solidification is an **exothermic** (**-**∆H)process that require particles holding the liquid together to lose energy to the surrounding.

(iii)Boiling/vaporization/evaporation involves **heating** a liquid to completely **break/free** the bonds holding the liquid particles together. Gaseous particles have high degree of **freedom** (**Kinetic Theory of matter**). Boiling /vaporization / evaporation is an **endothermic** (**+**∆H) process that require/absorb energy from the surrounding.

(iv)Condensation/liquidification is **reverse** process of boiling /vaporization / evaporation.It involves gaseous particles losing energy to the surrounding to form a liquid.It is an **exothermic**(**+**∆H) process.

The quantity of energy required to **change** one mole of a solid **to** liquid or to **form** one mole of a solid **from** liquid at constant temperature is called **molar enthalpy/latent heat of fusion**. e.g.

H2O(**s**) -> H2O(**l**) ∆H = **+6.0kJ mole-1** (endothermic process)

H2O(**l**) -> H2O(**s**) ∆H = **-6.0kJ mole-1** (exothermic process)

The quantity of energy required to **change** one mole of a liquid **to** gas/vapour or to **form** one mole of a liquid **from** gas/vapour at constant temperature is called **molar enthalpy/latent heat of vapourization**. e.g.

H2O(**l**) -> H2O(**g**) ∆H = **+44.0kJ mole-1** (endothermic process)

H2O(**g**) -> H2O(**l**) ∆H = **-44.0kJ mole-1** (exothermic process)

The following experiments illustrate/demonstrate practical determination of melting and boiling

**a) To determine the boiling point of water Procedure:**

Measure 20cm3 of tap water into a 50cm3 glass beaker. Determine and record its temperature.Heat the water on a strong Bunsen burner flame and record its temperature after every thirty seconds for four minutes. Sample results

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Time(seconds) | 0 | 30 | 60 | 90 | 120 | 150 | 180 | 210 | 240 |
| Temperature(oC) | 25.0 | 45.0 | 85.0 | 95.0 | 96.0 | 96.0 | 96.0 | 97.0 | 98.0 |

Questions

**1.Plot a graph of temperature against time(y-axis)**

Sketch graph of temperature against time

boiling point

96 oC

Temperature(0C)

25oC

time(seconds)

**2.From the graph show and determine the boiling point of water** Note:

Water boils at 100oC at sea level/one atmosphere pressure/101300Pa **but** boils at **below** 100oC at **higher** altitudes. The sample results above are from Kiriari Girls High School-Embu County on the slopes of Mt Kenya in Kenya. Water here boils at 96oC.

**3.Calculate the molar heat of vaporization of water.(H= 1.0,O= 16.O)** Working:

Mass of water = density x volume => (20 x 1) /1000 = **0.02kg**

Quantity of heat produced

= mass of water x specific heat capacity of water x temperature change

=>0.02kg x 4.2 x ( 96 – 25 ) = **5.964kJ**

Heat of vaporization of one mole H2O = Quantity of heat

Molar mass of H2O

=>5.964kJ = **0.3313 kJ mole -1**

18

**To determine the melting point of candle wax**

**Procedure**

Weigh exactly 5.0 g of candle wax into a boiling tube. Heat it on a strongly Bunsen burner flame until it completely melts. Insert a thermometer and remove the boiling tube from the flame. Stir continuously. Determine and record the temperature after every 30seconds for four minutes. **Sample results**

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Time(seconds) | 0 | 30 | 60 | 90 | 120 | 150 | 180 | 210 | 240 |
| Temperature(oC) | 93.0 | 85.0 | 78.0 | 70.0 | 69.0 | 69.0 | 69.0 | 67.0 | 65.0 |

Questions

**1.Plot a graph of temperature against time(y-axis)**

Sketch graph of temperature against time

9

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Temperat

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C)

melt

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69

o

C

time(seconds

)

2.From the graph show and determine the melting point of the candle wax

**4.Energy changes in chemical processes**

Thermochemical reactions measured at **standard** conditions of 298K(25oC) and 101300Pa/101300Nm2/ 1 atmospheres/760mmHg/76cmHg produce standard **enthalpies** denoted **∆Hᶿ.**

Thermochemical reactions are named from the type of reaction producing the energy change. Below are some thermochemical reactions:

1. Standard enthalpy/heat of reaction **∆Hᶿr**
2. Standard enthalpy/heat of combustion **∆Hᶿc**
3. Standard enthalpy/heat of displacement **∆Hᶿd** (d) Standard enthalpy/heat of neutralization **∆Hᶿn**

(e) Standard enthalpy/heat of solution/dissolution **∆Hᶿs** (f) Standard enthalpy/heat of formation **∆Hᶿf**

**(a)Standard enthalpy/heat of reaction ∆Hᶿr**

The molar standard enthalpy/heat of reaction may be defined as the energy/heat change when one mole of products is formed at standard conditions

1. chemical reaction involves the reactants forming products. For the reaction to take place the bonds holding the reactants must be broken so that new bonds of the products are formed. i.e.

**A**-**B** + C-D -> **A**-C + **B**-D

Old Bonds **broken** A-B and C-D on reactants

New Bonds **formed** A-C and B-D on products

The energy required to break one mole of a (covalent) bond is called **bond dissociation energy**. The SI unit of bond dissociation energy is **kJmole-1** The higher the bond dissociation energy the stronger the (covalent)bond

**Bond dissociation energies of some (covalent)bonds**

|  |  |  |  |
| --- | --- | --- | --- |
| Bond | Bond dissociation energy  (kJmole-1) |  | Bond dissociation energ y (kJmole-1) |
| H-H | 431 | I-I | 151 |
| C-C | 436 | C-H | 413 |
| C=C | 612 | O-H | 463 |
| C = C | 836 | C-O | 358 |
| N = N | 945 | H-Cl | 428 |
| N-H | 391 | H-Br | 366 |
| F-F | 158 | C-Cl | 346 |
| Cl-Cl | 239 | C-Br | 276 |
| Br-Br | 193 | C-I | 338 |
| H-I | 299 | O=O | 497 |
| Si-Si | 226 | C-F | 494 |

The molar enthalpy of reaction can be calculated from the bond dissociation energy by:

(i)**adding** the total bond dissociation energy of the **reactants**(endothermic process/**+∆H**) and total bond dissociation energy of the **products**(exothermic process/**-∆H**).

(ii)**subtracting** total bond dissociation energy of the **reactants** **from** the total bond dissociation energy of the **products**(exothermic process/**-∆H less/minus** endothermic process/**+∆H**).

**Practice examples/Calculating ∆Hr**

**1.Calculate ∆Hr from the following reaction:**

**a) H2(g)+ Cl2(g)-> 2HCl(g)**

Working

Old bonds broken (endothermic process/**+∆H** )

= (H-H + Cl-Cl) => (+431 + (+ 239)) = **+ 670kJ**

New bonds broken (exothermic process/**-∆H** )

= (2(H-Cl ) => (- 428 x 2)) = **-856kJ**

∆Hr =(+ 670kJ + -856kJ) = 186 kJ = **-93kJ mole-1**

2

The above reaction has negative **-∆H** enthalpy change and is therefore practically exothermic.

The thermochemical reaction is thus:

½ H2(g)+ ½ Cl2(g)-> HCl(g) ∆Hr = -93kJ

**b) CH4(g)+ Cl2(g)-> CH3Cl + HCl(g)**

Working

Old bonds broken (endothermic process/**+∆H** )

= (4(C-H) + Cl-Cl)

=> ((4 x +413) + (+ 239)) = **+ 1891kJ**

New bonds broken (exothermic process/**-∆H** )

= (3(C-H + H-Cl + C-Cl)

=> (( 3 x - 413) + 428 + 346) = -**2013 kJ**

∆Hr =( + 1891kJ + -2013 kJ) = **-122 kJ mole-1**

The above reaction has negative **-∆H** enthalpy change and is therefore practically exothermic.

The thermochemical reaction is thus:

CH4(g)+ Cl2(g)-> CH3Cl(g) + HCl(g) ∆H = -122 kJ

**c) CH2CH2(g)+ Cl2(g)-> CH3Cl CH3Cl (g)**

Working

Old bonds broken (endothermic process/**+∆H** )

= (4(C-H) + Cl-Cl + C=C)

=> ((4 x +413) + (+ 239) +(612)) = **+ 2503kJ**

New bonds broken (exothermic process/**-∆H** )

= (4(C-H + C-C + 2(C-Cl) )

=> (( 3 x - 413) + -436 +2 x 346 = -**2367 kJ**

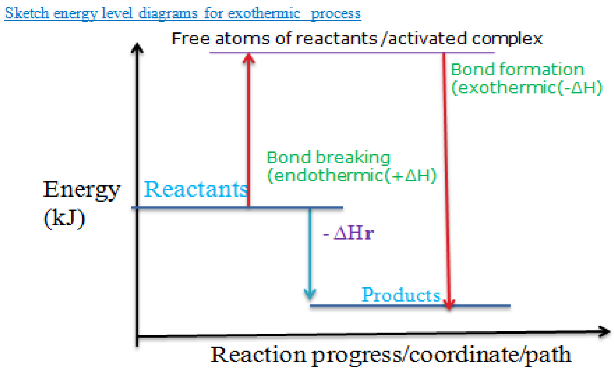
∆Hr =( + 2503kJ + -2367 kJ) = **+136 kJ mole-1**

The above reaction has negative **+∆H** enthalpy change and is therefore practically endothermic.

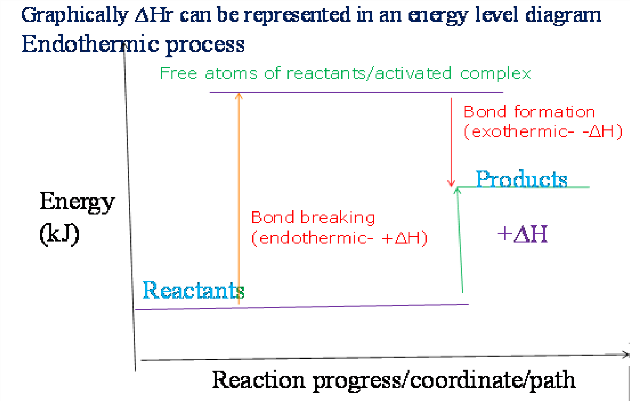
The thermochemical reaction is thus:

CH2CH2(g)+ Cl2(g)-> CH3Cl CH3Cl (g) ∆H = +136 kJ Note that:

1. a reaction is **exothermic** if the bond dissociation energy of **reactants** is **more** than bond dissociation energy of products.
2. a reaction is **endothermic** if the bond dissociation energy of **reactants** is **less** than bond dissociation energy of products.



**(b)Standard enthalpy/heat of combustion ∆Hᶿc**



The molar standard enthalpy/heat of **combustion(∆Hᶿc**) is defined as the energy/heat change when one mole of a substance is burnt in oxygen/excess air at standard conditions.

Burning is the reaction of a substance with oxygen/air. It is an exothermic process producing a lot of energy in form of heat.

A substance that undergoes burning is called a **fuel.** A fuel is defined as the combustible substance which burns in air to give heat energy for domestic or industrial use. A fuel may be **solid** (e.g coal, wood, charcoal) **liquid** (e.g petrol, paraffin, ethanol, kerosene) or **gas** (e.g liquefied petroleum gas/LPG, Water gasCO2/H2, biogas-methane, Natural gas-mixture of hydrocarbons)

To determine the molar standard enthalpy/heat of **combustion(∆Hᶿc**) of ethanol

**Procedure**

Put 20cm3 of distilled water into a 50cm3 beaker. Clamp the beaker. Determine the temperature of the water T1.Weigh an empty burner(empty tin with wick). Record its mass M1.Put some ethanol into the burner. Weigh again the burner with the ethanol and record its mass M2. Ignite the burner and place it below the clamped 50cm3 beaker. Heat the water in the beaker for about one minute. Put off the burner. Record the highest temperature rise of the water, T2. Weigh the burner again and record its mass M3

**Sample results:**

|  |  |
| --- | --- |
| Volume of water used | 20cm3 |
| Temperature of the water before heating T1 | 25.0oC |
| Temperature of the water after heating T2 | 35.0oC |
| Mass of empty burner M1 | 28.3g |
| Mass of empty burner + ethanol before igniting M2 | 29.1g |
| Mass of empty burner + ethanol after igniting M3 | 28.7g |

**Sample calculations:**

**1.Calculate:**

1. **∆T the change in temperature**

**∆T =** T2 – T1  => (35.0oC – 25.0oC) = **10.0oC**

1. **the mass of ethanol used in burning**

mass of ethanol used = M2 – M1 => 29.1g – 28.7g = **0.4g**

1. **the number of moles of ethanol used in burning**  moles of ethanol = mass used => 0.4 = **0.0087 /8.7 x 10-3**moles molar mass of ethanol 46

**2. Given that the specific heat capacity of water is 4.2 kJ-1kg-1K-1,determine the heat produced during the burning.**

Heat produced **∆H** = mass of water(**m)** x specific heat capacity (**c**)x **∆T**

=> 20 x 4.2 x 10 = **840 J**oules = **0.84 kJ**

1000

**3.Calculate the molar heat of combustion of ethanol**

Molar heat of combustion ∆Hc = Heat produced ∆H

Number of moles of fuel

=> 0.84 kJ = **96.5517 kJmole-1**

0.0087 /8.7 x 10-3 moles

**4.List two sources of error in the above experiment.**

(i)Heat loss to the surrounding **lowers** the practical value of the molar heat of combustion of ethanol.

A draught shield tries to minimize the loss by protecting wind from wobbling the flame.

(ii) Heat gain by reaction vessels/beaker **lowers** ∆T and hence ∆Hc

**5.Calculate the heating value of the fuel.**

Heating value = molar heat of combustion => 96.5517 kJmole-1 = **2.0989 kJg-1**

Molar mass of fuel 46 g

**Heating value is the enrgy produced when a unit mass/gram of a fuel is completely burnt**

**6.Explain other factors used to determine the choice of fuel for domestic and industrial use.**

(i) **availability and affordability**-some fuels are more available cheaply in rural than in urban areas at a lower cost.

(ii)**cost of storage and transmission**-a fuel should be easy to transport and store safely. e.g LPG is very convenient to store and use. Charcoal and wood are bulky.

(iii)**environmental effects –**Most fuels after burning produce carbon(IV) oxide gas as a byproduct. Carbon(IV) oxide gas is green house gas that causes global warming. Some other fuel produce acidic gases like sulphur(IV) oxide ,and nitrogen(IV) oxide. These gases cause acid rain. Internal combustion engines exhaust produce lead vapour from leaded petrol and diesel. Lead is carcinogenic.

(iv)**ignition point-**The temperature at which a fuel must be heated before it burns in air is the ignition point. Fuels like petrol have very low ignition point, making it highly flammable. Charcoal and wood have very high ignition point.

**7.Explain the methods used to reduce pollution from common fuels.**

**(i)Planting trees-**Plants absorb excess carbon(IV)oxide for photosynthesis and release oxygen gas to the atmosphere.

**(ii)using catalytic converters** in internal combustion engines that convert harmful/toxic/poisonous gases like carbon(II)oxide and nitrogen(IV)oxide to harmless non-poisonous carbon(IV)oxide, water and nitrogen gas by using platinum-rhodium catalyst along the engine exhaust pipes.

Further practice calculations

**1.Calculate the heating value of methanol CH3OH given that 0.87g of the fuel burn in air to raise the temperature of 500g of water from 20oC to 27oC.(C12.0,H=1.0 O=16.0).**

Moles of methanol used = Mass of methanol used => 0.87 g = **0.02718 moles**

Molar mass of methanol 32

Heat produced **∆H** = mass of water(**m)** x specific heat capacity (**c**)x **∆T**

=> 500 x 4.2 x 7 = 14700 **J**oules = **14.7 kJ**

1000

Molar heat of combustion ∆Hc = Heat produced ∆H

Number of moles of fuel

=> 14.7 kJ = **540.8389 kJmole-1**

0.02718 moles

Heating value = molar heat of combustion => 540.8389 kJmole-1 = **16.9012 kJg-1**

Molar mass of fuel 32 g

**2. 1.0 g of carbon burn in excess air to raise the temperature of 400g of water by 18oC.Determine the molar heat of combustion and hence the heating value of carbon(C-12.0,).**

Moles of carbon used = Mass of carbon used => 1.0 g = **0.0833 moles**

Molar mass of carbon 12

Heat produced **∆H** = mass of water(**m)** x specific heat capacity (**c**)x **∆T**

=> 400 x 4.2 x 18 = 30240 **J**oules = **30.24 kJ**

1000

Molar heat of combustion ∆Hc = Heat produced ∆H

Number of moles of fuel

=> 30.24 kJ = **363.0252 kJmole-1**

0.0833 moles

Heating value = molar heat of combustion => 363.0252 kJmole-1= **30.2521 kJg-1**

Molar mass of fuel 12 g

**(c)Standard enthalpy/heat of displacement ∆Hᶿd**

The molar standard enthalpy/heat of **displacement ∆Hᶿd** is defined as the energy/heat change when one mole of a substance is displaced from its solution. A displacement reaction takes place when a more reactive element/with less electrode potential **Eᶿ** /negative Eᶿ /higher in the reactivity/electrochemical series remove/displace another with less reactive element/with higher electrode potential **Eᶿ** /positive Eᶿ /lower in the reactivity/electrochemical series **from** its solution.e.g.

(i)Zn(s) + CuSO4(aq) -> Cu(s) + ZnSO4(aq)

Ionically: Zn(s) + Cu2+(aq) -> Cu(s) + Zn2+ (aq)

(ii)Fe(s) + CuSO4(aq) -> Cu(s) + FeSO4(aq)

Ionically: Fe(s) + Cu2+(aq) -> Cu(s) + Fe2+ (aq)

(iii)Pb(s) + CuSO4(aq) -> Cu(s) + PbSO4(**s**)

This reaction stops after some time as insoluble PbSO4(**s**) coat/cover unreacted lead.

(iv)Cl2(g) + 2NaBr(aq) -> Br2(aq) + 2NaCl(aq)

Ionically: Cl2(g)+ 2Br- (aq) -> Br2(aq) + 2Cl- (aq)

Practically, a displacement reaction takes place when a known amount /volume of a solution is added excess of a more reactive metal.

To determine the molar standard enthalpy/heat of **displacement(∆Hᶿd**) of copper **Procedure**

Place 20cm3 of 0.2M copper(II)sulphate(VI)solution into a 50cm3 plastic beaker/calorimeter. Determine and record the temperature of the solution T1.Put all the Zinc powder provided into the plastic beaker. Stir the mixture using the thermometer. Determine and record the highest temperature change to the nearest

0.5oC- T2 . Repeat the experiment to complete table 1 below

Table 1

|  |  |  |
| --- | --- | --- |
| Experiment | I | II |
| Final temperature of solution(T2) | 30.0oC | 31.0oC |
| Final temperature of solution(T1) | 25.0oC | 24.0oC |
| Change in temperature(∆T) | 5.0 | 6.0 |

**Questions**

**1.(a) Calculate:**

**(i)average ∆T**

Average∆T = change in temperature in experiment I and II

=>5.0 + 6.0 = **5.5oC**

2

**(ii)the number of moles of solution used**

Moles used = molarity x volume of solution = 0.2 x 20 = **0.004 moles**

1000 1000

**(iii)the enthalpy change ∆H for the reaction**

Heat produced **∆H** = mass of solution(**m)** x specific heat capacity (**c**)x **∆T**

=> 20 x 4.2 x 5.5 = 462 **J**oules = **0.462 kJ**

1000

**(iv)State two assumptions made in the above calculations.**

Density of solution = density of water = 1gcm-3

Specific heat capacity of solution=Specific heat capacity of solution=4.2 kJ-1kg-1K This is because the solution is assumed to be **infinite dilute.**

**2. Calculate the enthalpy change for one mole of displacement of Cu2+ (aq) ions.**

Molar heat of displacement ∆Hd = Heat produced ∆H

Number of moles of fuel

=> 0.462 kJ = **115.5 kJmole-1**

0.004

**3.Write an ionic equation for the reaction taking place.**

Zn(s) + Cu2+(aq) -> Cu(s) + Zn2+(aq)

**4.State the observation made during the reaction.**

Blue colour of copper(II)sulphate(VI) fades/becomes less blue/colourless. Brown solid deposits are formed at the bottom of reaction vessel/ beaker.

**5.Illustrate the above reaction using an energy level diagram.**

Z

n(s) + Cu

2+

(

aq

)

Energy

∆

H

=

-

kJmole

115.5

-

1

(

kJ

)

Cu(s) + Zn

2+

(

aq

)

Reaction progress/path/co

o

rdinates

**6. Iron is less reactive than Zinc. Explain the effect of using iron instead of Zinc on the standard molar heat of displacement ∆Hd of copper(II)sulphate (VI) solution.**

No effect.Cu2+ (aq) are displaced from their solution.The element used to displace it does not matter.The reaction however faster if a more reactive metal is used.

**7.(a)If the standard molar heat of displacement ∆Hd of copper(II)sulphate (VI) solution is 209kJmole-1 calculate the temperature change if 50cm3 of 0.2M solution was displaced by excess magnesium.**

Moles used = molarity x volume of solution = 0.2 x 50 = **0.01 moles**

1000 1000

Heat produced ∆H = Molar heat of displacement ∆Hd x Number of moles

=>209kJmole-1x 0.01 moles = **2.09 kJ**

**∆**T (change in temperature) **=**  Heat produced ∆H

Molar heat of displacement ∆Hd x Number of moles

=>2.09 kJ = **9.9524K**elvin

0.01 moles

**(b)Draw an energy level diagram to show the above energy changes**

Mg(s) + Cu

2+

(

aq

)

Energy

∆

H

=

-

209

kJmole

-

1

(

kJ

)

Cu(s) + Mg

2+

(

aq

)

Reaction progress/path/coordinates

1. **The enthalpy of displacement ∆Hd of copper(II)sulphate (VI) solution is 12k6kJmole-1.Calculate the molarity of the solution given that 40cm3 of this solution produces 2.204kJ of energy during a displacement reaction with excess iron filings.**

Number of moles **=**  Heat produced ∆H

Molar heat of displacement ∆Hd

=>2.204 kJ = **0.0206moles**

126 moles

Molarity of the solution = moles x 1000

Volume of solution used

= 0.0206moles x 1000 = **0.5167 M**

40

1. **If the molar heat of displacement of Zinc(II)nitrate(V)by magnesium powder is 25.05kJmole-1 ,calculate the volume of solution which must be added 0.5 moles solution if there was a 3.0K rise in temperature.**

Heat produced ∆H = Molar heat of displacement ∆Hd x Number of moles

=>25.08kJmole-1x 0.5 moles = **1.254 kJ** x 1000 **=1254J**

Mass of solution (**m) =** Heat produced **∆H**

specific heat capacity (**c**)x **∆T**

=> 1254J = **99.5238 g**

4.2 x 3

Volume = mass x density = 99.5238 g x 1 = **99.5238cm3**

Note: The solution assumes to be too dilute /infinite dilute such that the density and specific heat capacity is assumed to be that of water.

**Graphical determination of the molar enthalpy of displacement of copper Procedure:**

Place 20cm3 of 0.2M copper(II)sulphate (VI) solution into a calorimeter/50cm3 of plastic beaker wrapped in cotton wool/tissue paper.

Record its temperature at time T= 0.

Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds .

Place all the (1.5g) Zinc powder provided.

Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds for five minutes.

Determine the highest temperature change to the nearest 0.5oC.

**Sample results**

|  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Time oC | 0.0 | 30.0 | 60.0 | 90.0 | 120.0 | 150.0 | 180.0 | 210.0 | 240.0 | 270.0 |
| Temperature | 25.0 | 25.0 | 25.0 | 25.0 | 25.0 | xxx | 36.0 | 35.5 | 35.0 | 34.5 |

**Sketch graph of temperature against time**

36.5

Extrapolation

Te

mperature

point

∆

T

o

C

130Time(seconds)

**Questions**

1. **Show and determine the change in temperature ∆T**

From a well constructed graph ∆T= T2 –T1 at **150** second by **extrapolation**

∆T = 36.5 – 25.0 = 11.5oC

**2.Calculate the number of moles of copper(II) sulphate(VI)used given the molar heat of displacement of Cu2+ (aq)ions is 125kJmole-1**

Heat produced **∆H** = mass of solution(**m)** x specific heat capacity (**c**)x **∆T**

=> 20 x 4.2 x 11.5 = 966 **J**oules = **0.966 kJ**

1000

Number of moles **=**  Heat produced ∆H

Molar heat of displacement ∆Hd

=>.966 kJ = **0.007728moles**

125 moles **7.728 x 10-3moles**

1. **What was the concentration of copper(II)sulphate(VI) in moles per litre.**

Molarity = moles x 1000 => 7.728 x 10-3moles x 1000 = **0.3864M**

Volume used 20

**4.The actual concentration of copper(II)sulphate(VI) solution was**

**0.4M.Explain the differences between the two.**

Practical value is **lower** than theoretical. Heat/energy **loss** to the surrounding and that absorbed by the reaction vessel **decreases** ∆T hence **lowering** the practical number of moles and molarity against the theoretical value

5.a) In an experiment to determine the molar heat of reaction when magnesium displaces copper ,0.15g of magnesium powder were added to 25.0cm3 of 2.0M copper (II) chloride solution. The temperature of copper (II) chloride solution was 25oC.While that of the mixture was 43oC.

i)Other than increase in temperature, state and explain the observations which were made during the reaction.(3mks)

ii)Calculate the heat change during the reaction (specific heat capacity of the solution = 4.2jg-1k-1and the density of the solution = 1g/cm3(2mks)

iii)Determine the molar heat of displacement of copper by magnesium.(Mg=24.0).

iv)Write the ionic equation for the reaction.(1mk)

v)Sketch an energy level diagram for the reaction.(2mks)

**(c)Standard enthalpy/heat of neutralization ∆Hᶿn**

The molar standard enthalpy/heat of **neutralization ∆Hᶿn** is defined as the energy/heat change when one mole of a H+ (H3O+)ions react completely with one mole of OH- ions to form one mole of H2O/water.

Neutralization is thus a reaction of an acid /H+ (H3O+)ions with a base/alkali/ OH- ions to form salt and water only.

Strong acids/bases/alkalis are completely dissociated to **many** free ions(H+ /H3O+ and OH- ions).

Weak acids/bases/alkalis are partially dissociated to **few** free ions(H+ (H3O+ and OH- ions) and exist **more** as molecules.

Neutralization is an exothermic(-∆H) process.The enrgy produced during neutralization depend on the amount of free ions (H+ H3O+ and OH-)ions existing in the acid/base/alkali reactant:

(i)for weak acid-base/alkali neutralization,some of the **energy** is used to dissociate /ionize the **molecule** into free H+ H3O+ and OH- ions therefore the overall energy evolved is comparatively **lower/lesser/smaller** than strong acid / base/ alkali neutralizations.

(ii) (i)for strong acid/base/alkali neutralization, no **energy** is used to dissociate /ionize since **molecule** is wholly/fully dissociated/ionized into free H+ H3O+ and OH- ions.The overall energy evolved is comparatively **higher/more** than weak acid-base/ alkali neutralizations. For strong acid-base/alkali neutralization, the enthalpy of neutralization is **constant** at about **57.3**kJmole-1 irrespective of the acid-base used. This is because ionically:

**OH-(aq)+ H+(aq) -> H2O(l)** for any wholly dissociated acid/base/alkali Practically ∆Hᶿn can be determined as in the examples below:

To determine the molar enthalpy of neutralization ∆Hn of Hydrochloric acid **Procedure**

Place 50cm3 of 2M hydrochloric acid into a calorimeter/200cm3 plastic beaker wrapped in cotton wool/tissue paper. Record its temperature T1.Using a clean measuring cylinder, measure another 50cm3 of 2M sodium hydroxide. Rinse the bulb of the thermometer in distilled water. Determine the temperature of the sodium hydroxide T2.Average T2 andT1 to get the initial temperature of the mixture

T3.

Carefully add all the alkali into the calorimeter/200cm3 plastic beaker wrapped in cotton wool/tissue paper containing the acid. Stir vigorously the mixture with the thermometer.

Determine the highest temperature change to the nearest 0.5oC T4 as the final temperature of the mixture. Repeat the experiment to complete table 1.

Table I . **Sample results**

|  |  |  |
| --- | --- | --- |
| Experiment | I | II |
| Temperature of acid T1 (oC) | 22.5 | 22.5 |
| Temperature of base T2 (oC) | 22.0 | 23.0 |
| Final temperature of solution T4(oC) | 35.5 | 36.0 |
| Initial temperature of solution T3(oC) | 22.25 | 22.75 |
| Temperature change( T5) | 13.25 | 13.75 |

**(a)Calculate T6 the average temperature change**

T6 = 13.25 +13.75 = 13.5 oC

2

**(b)Why should the apparatus be very clean?**

Impurities present in the apparatus reacts with acid /base lowering the overall temperature change and hence ∆Hᶿn.

**(c)Calculate the:**

**(i)number of moles of the acid used**

number of moles = molarity x volume => 2 x 50 = **0.1moles**

1000 1000

**(ii)enthalpy change** ∆H **of neutralization.**

∆H = (**m**)mass of solution(acid+base) x (**c**)specific heat capacity of solution x

**∆T**(T6) => (50 +50) x 4.2 x 13.5 = **5670J**oules = **5.67kJ**

**(iii) the molar heat of neutralization the acid.**

∆Hn = Enthalpy change ∆H => 5.67kJ = **56.7kJ mole-1**

Number of moles 0.1moles

**(c)Write the ionic equation for the reaction that takes place**

OH-(aq)+ H+(aq) -> H2O(l)

**(d)The theoretical enthalpy change is 57.4kJ. Explain the difference with the results above.**

The theoretical value is higher

Heat/energy loss to the surrounding/environment lowers **∆T**/T6 and thus ∆Hn Heat/energy is absorbed by the reaction vessel/calorimeter/plastic cup lowers ∆T and hence ∆Hn

**(e)Compare the ∆Hn of the experiment above with similar experiment repeated with neutralization of a solution of:**

1. **potassium hydroxide with nitric(V) acid** The results would be the same/similar.

Both are neutralization reactions of strong acids and bases/alkalis that are fully /wholly dissociated into many free H+ / H3O+ and OH- ions.

1. **ammonia with ethanoic acid**

The results would be lower/∆Hn would be less.

Both are neutralization reactions of weak acids and bases/alkalis that are partially /partly dissociated into few free H+ / H3O+ and OH- ions. Some energy is used to ionize the molecule.

**(f)Draw an energy level diagram to illustrate the energy changes**

H

**2**

H

+

(

aq

)

+

OH

-

(

aq

)

Energy

)

kJ

(

∆

H =

**-**

**56.7**

**kJ**

H

1

H

2

O (

l

)

Reaction path/coordinate/progress

**Theoretical examples**

**1.The molar enthalpy of neutralization was experimentary shown to be 51.5kJ per mole of 0.5M hydrochloric acid and 0.5M sodium hydroxide. If the volume of sodium hydroxide was 20cm3, what was the volume of hydrochloric acid used if the reaction produced a 5.0oC rise in temperature?**

Working:

Moles ofsodium hydroxide = molarity x volume => 0.5 M x 20cm3 = **0.01 moles**

1000 1000

Enthalpy change∆H **=**  ∆Hn => 51.5 = **0.515kJ**

Molessodium hydroxide 0.01 moles

Mass of base + acid = Enthalpy change∆H in Joules

Specific heat capacity x ∆T

=> 0.515kJ x 1000 = **24.5238**g

4.2 x 5

Mass/volume of HCl = Total volume – volume of NaOH

=>24.5238 - 20.0 = **4.5238 cm3**

**3. ∆Hn of potassium hydroxide was practically determined to be**

**56.7kJmole-1.Calculate the molarity of 50.0 cm3 potassium hydroxide used to neutralize 25.0cm3 of dilute sulphuric(VI) acid raising the temperature of the solution from 10.0oC to 16.5oC.**

∆H = (**m**)mass of solution(acid+base) x (**c**)specific heat capacity of solution x **∆T** => (50 +25) x 4.2 x 6.5 = **2047.5J**oules

Molespotassium hydroxide =Enthalpy change∆H

∆Hn

2047.5Joules = **0.0361 moles**

56700Joules

Molarity of KOH = moles x 1000 => 0.0361 moles x 1000 = **0.722M**

Volume used 50cm3

**3.Determine the specific heat capacity of a solution of a solution mixture of**

**50.0cm3 of 2M potassium hydroxide neutralizing 50.0cm3 of 2M nitric(V)**

**acid if a 13.25oC rise in temperature is recorded.(1mole of potassium hydroxide produce 55.4kJ of energy)**

Moles ofpotassium hydroxide = molarity KOH x volume

1000

=> 2 M x 50cm3 = **0.1 moles**

1000

Enthalpy change∆H  **=** ∆Hn x Molespotassium hydroxide

=> 55.4kJ x 0.1 moles = 5.54kJ x 1000=**5540J**oules

Specific heat capacity = Enthalpy change∆H in Joules

Mass of base + acid x ∆T

=> 5540 =  **4.1811J-1g-1K-1**

(50+50) x 13.25

Graphically ∆Hn can be determined as in the example below:

**Procedure**

Place 8 test tubes in a test tube rack .Put 5cm3 of 2M sodium hydroxide solution into each test tube.

Measure 25cm3 of 1M hydrochloric acid into 100cm3 plastic beaker.

Record its initial temperature at volume of base =0. Put one portion of the base into the beaker containing the acid.

Stir carefully with the thermometer and record the highest temperature change to the nearest 0.5oC.

Repeat the procedure above with other portions of the base to complete table 1 below

**Table 1:Sample results.**

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| olume of acid(cm3) | 25.0 | 25.0 | 25.0 | 25.0 | 25.0 | 25.0 | 25.0 | 25.0 | 25.0 |
| Volume of alkali(cm3) | 0 | 5.0 | 10.0 | 15.0 | 20.0 | 25.0 | 30.0 | 35.0 | 40.0 |
| Final temperature(oC) | **22.0** | **24.0** | **26.0** | **28.0** | **28.0** | **27.0** | **26.0** | **25.0** | **24.0** |
| Initial temperature(oC) | **22.0** | **22.0** | **22.0** | **22.0** | **22.0** | **22.0** | **22.0** | **22.0** | **22.0** |
| Change in temperature | **0.0** | **2.0** | **4.0** | **6.0** | **6.0** | **5.0** | **4.0** | **3.0** | **2.0** |

(a)Complete the table to determine the change in temperature.

(b)Plot a graph of volume of sodium hydroxide against temperature change.

**6**

**.7=**

T

2

∆T (oC)

**0**  **=T1**

**From the graph show and determine :**

**(i)the highest temperature change ∆T**

∆T =T2-T1 => highest temperature-T2 (from extrapolating a correctly plotted graph) less lowest temperature at volume of base=0 :T1

=>∆T = 6.7 – 0.0 = **6.70C**

**(ii)the volume of sodium hydroxide used for complete neutralization** From a correctly plotted graph – **16.75cm3**

**(c)Calculate the number of moles of the alkali used**

Moles NaOH = molarity x volume =>2M x 16.75cm3 = **0.0335 moles**

1000 1000

**(d)Calculate ∆H for the reaction**

∆H = mass of solution(acid+base) x c x ∆T

=>(25.0 + 16.75) x 4.2 x 6.7 = 1174.845 J = **1.174845kJ**

1000

**(e)Calculate the molar enthalpy of neutralization of the alkali.**

∆Hn  = ∆Hn  = 1.174845kJ =  **35.0701kJ**

Number of moles 0.0335

**(d)Standard enthalpy/heat of solution ∆Hᶿs**

The standard enthalpy of solution ∆Hᶿsis defined as the energy change when one mole of a substance is dissolve in excess distilled water to form an infinite dilute solution. An infinite dilute solution is one which is **too** dilute to be diluted further.

Dissolving a solid involves two processes:

(i) **breaking** the **crystal** of the solid into **free ions**(cations and anion).This process is the **opposite** of the **formation** of the crystal itself. The energy required to form one mole of a crystal structure from its **gaseous ions** is called **Lattice energy**/heat/enthalpy of lattice (**∆Hl**).Lattice energy /heat/enthalpy of lattice (**∆**Hl) is an endothermic process (**+∆Hl**).

The table below shows some **∆Hl** in kJ for the process MX(s) -> M+ (g) + X- (g)

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Li | Na | K | Ca | Mg |
| F | +1022 | +900 | +800 | +760 | +631 |
| Cl | +846 | +771 | +690 | +2237 | +2493 |
| Br | +800 | +733 | +670 | +2173 | +2226 |

(ii)**surrounding** the free ions by polar **water** molecules. This process is called **hydration.** The energy produced when one mole of ions are completely hydrated is called **hydration energy**/ heat/enthalpy of hydration(**∆Hh**).Hydration energy /enthalpy of hydration(∆Hh) is an **exothermic** process(∆Hh).

The table below shows some ∆Hhin kJ for some ions;

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| ion | Li+ | Na+ | K+ | Mg2+ | Ca2+ | F- | Cl- | Br- |
| ∆Hh | -1091 | -406 | -322 | -1920 | -1650 | -506 | -364 | -335 |

The **sum** of the **lattice** energy **+∆Hl**(endothermic) and **hydration** energy **-∆Hh**

(exothermic) gives the heat of solution**-∆Hs**

**∆Hs = ∆Hl +∆Hh Note**

Since ∆Hl is an endothermic process and ∆Hh is an exothermic process then ∆Hs is**:** (i)**exothermic** if **∆Hl** is **less** than **∆Hh** and hence a solid **dissolve** easily in water.

(ii)**endothermic** if **∆Hl** is **more** than **∆Hh** and hence a solid does not **dissolve** easily in water.

(a)Dissolving sodium chloride crystal/s:

(i) **NaCl -**---breaking the crystal into free ions--**-> Na +(g)+ Cl-(g) ∆**Hl **=+771 kJ** (ii) Hydrating the ions;

Na +(g) + **aq** -> Na(aq) **∆Hh = - 406 kJ**

Cl-(g) + aq -> Cl-(aq) **∆Hh = - 364 kJ**

**∆Hs =∆Hh** +**∆Hs** ->(- 406 kJ + - 364 kJ) + +771 kJ = **+ 1.0 kJmole-1**

NaCl does not **dissolve** easily in water because overall **∆Hs** is **endothermic**

Solubility of NaCl therefore increases with increase in temperature.

Increase in temperature increases the energy to break the crystal lattice of NaCl to freeNa +(g)+ Cl-(g)

**(b)Dissolving magnesium chloride crystal/s//**  **MgCl2 (s) ->MgCl2 (aq)**

(i) MgCl2 **-**-breaking the crystal into free ions-**->**Mg 2+(g)+ 2Cl-(g) ∆Hl =+2493 kJ (ii) Hydrating the ions;

Mg 2+(g) + aq -> Mg 2+(g) (aq) ∆Hh = - 1920 kJ

2Cl-(g) + aq -> 2Cl-(aq) ∆Hh = (- 364 x 2) kJ

∆Hs =∆Hh +∆Hs -> (- 1920 kJ + (- 364 x 2 kJ)) + +2493 kJ = -**155.0 kJmole-1**

MgCl2 (s) **dissolve** easily in water because overall **∆Hs** is **exothermic** .

Solubility of MgCl2 (s) therefore decreases with increase in temperature.

**(c)Dissolving Calcium floride crystal/s//**  **CaF2 (s) -> CaF2 (aq)**

1. **CaF2 -->**Ca 2+(g)+ 2F-(g) ∆Hl =+760 kJ
2. Hydrating the ions;

Ca 2+(g) + aq -> Ca 2+(g) (aq) ∆Hh = - 1650 kJ

2F-(g) + aq -> 2F-(aq) ∆Hh = (- 506 x 2) kJ

∆Hs =∆Hh +∆Hs -> (- 1650 kJ + (- 506 x 2 kJ)) + +760 kJ = -**1902.0 kJmole-1**

CaF2 (s) **dissolve** easily in water because overall **∆Hs** is **exothermic** .

Solubility of CaF2 (s) therefore decreases with increase in temperature.

**(d)Dissolving magnesium bromide crystal/s//**  **MgBr2 (s) ->MgBr2 (aq)**

1. MgCl2 **-**-breaking the crystal into free ions-**->**Mg 2+(g)+ 2Br-(g) ∆Hl =+2226 kJ
2. Hydrating the ions;

Mg 2+(g) + aq -> Mg 2+(g) (aq) ∆Hh = - 1920 kJ

2Br-(g) + aq -> 2Br-(aq) ∆Hh = (- 335x 2) kJ

∆Hs =∆Hh +∆Hs -> (- 1920 kJ + (- 335 x 2 kJ)) + +2226 kJ = -**364.0 kJmole-1**

MgBr2 (s) **dissolve** easily in water because overall **∆Hs** is **exothermic** .

Solubility of MgBr2(s) therefore decreases with increase in temperature.

Practically the heat of solution can be determined from dissolving known amount /mass/volume of solute in known mass /volume of water/solvent.

From the temperature of solvent **before** and **after** dissolving the change in temperature(∆T) during dissolution is determined.

To determine the ∆Hs ammonium nitrate

Place 100cm3 of distilled water into a plastic beaker/calorimeter. Determine its temperature and record it at time =0 in table I below.

Put all the 5.0g of ammonium nitrate (potassium nitrate/ammonium chloride can also be used)provided into the plastic beaker/calorimeter, stir using a thermometer and record the highest temperature change to the nearest 0.5oCafter every ½ minute to complete table I.

Continue stirring the mixture throughout the experiment. **Sample results: Table I**

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Time (minutes) | 0.0 | ½ | 1 | 1 ½ | 2 | 2 ½ | 3 | 3 ½ |
| Temperature()oC | **22.0** | **21.0** | **20.0** | **19.0** | **19.0** | **19.5** | **20.0** | **20.5** |

(a)Plot a graph of temperature against time(x-axis)

2

2.0

**=**

T

1

temperature(

o

C)

∆

T

Time (minutes

)

18.7.

o

C

T

1

**(b)From the graph show and determine the highest temperature change ∆T**  ∆T =T2-T1 => lowest temperature-T2 (from extrapolating a correctly plotted graph) less highest temperature at volume of base=0 :T1

=>∆T =18.7 – 22.0 = **3.30C**

**(c)Calculate the number of moles of ammonium nitrate(V) used**

Moles NH4NO3 = mass used => 5.0 = **0.0625 moles**

Molar mass 80

**(d)Calculate ∆H for the reaction**

∆H = mass of water x c x ∆T

->100 x 4.2 x 3.3 = **+**1386 J = **+1.386kJ**

1000

**(e)Calculate the molar enthalpy of dissolution of ammonium nitrate(V).**

∆Hs  = ∆H = +1.386kJ = + 22.176kJ mole-1

Number of moles 0.0625 moles

**(f)What would happen if the distilled water was heated before the experiment was performed.**

The ammonium nitrate(V)would take less time to dissolves. Increase in temperature reduces lattice energy causing endothermic dissolution to be faster

**g**

**(**

**)**

**Illustrate the process above in an energy level diagram**

NH

4

+

(

g

)

NO

+

3

-

(

)

g

+

∆

H

NH

4

+

aq

(

)

+

NO

3

-

(

aq

)

Energy(kJ)

+

∆

H

∆

H =

**-**

**22.176**

**kJ**

NH

4

NO

3

-

s

(

)

Reaction path /progress/coordinate

**(h) 100cm3 of distilled water at 25oC was added carefully 3cm3 concentrated sulphuric(VI)acid of density 1.84gcm-3.The temperature of the mixture rose from 250C to 38oC.Calculate the molar heat of solution of sulphuric(VI)acid (S=32.0,H=1.0,0=16.0)**

Working

Molar mass of H2SO4 = 98g

Mass of H2SO4= Density x volume => 1.84gcm-3  x 3cm3 = **5.52 g**

Mass of H2O = Density x volume => 1.00gcm-3  x 100cm3 = **100 g**

Moles of H2SO4= mass => 5.52 g = **0.0563 moles**

Molar mass of H2SO4  98g

Enthalpy change ∆H= (mass of acid + water) x specific heat capacity of water x ∆T

=> (100 +5.52 g) x 4.2 x 13oC = 5761.392 J = **5.761392 kJ**

1000

∆Hs of H2SO4= ∆H => 5.761392 kJ = **-102.33378kJmoles-1**

Moles of H2SO4  0.0563 moles

**(e)Standard enthalpy/heat of formation ∆Hᶿf**

The molar enthalpy of formation ∆Hᶿf is defined as the energy change when one mole of a compound is formed from its elements at 298K(25oC) and 101325Pa(one atmosphere)pressure. ∆Hᶿf is practically difficult to determine in a school laboratory.

It is determined normally determined by applying Hess’ law of constant heat summation.

Hess’ law of constant heat summation states that **“the total enthalpy/heat/energy change of a reaction is the same regardless of the route taken from reactants**

**to products at the same temperature and pressure”.**

Hess’ law of constant heat summation is as a result of a series of experiments done by the German Scientist Henri Hess(1802-1850).

He found that the total energy change from the reactants to products was the same irrespective of the intermediate products between. i.e.

A(s) --**∆H1**-->C(s) = A(s) --**∆H2**-->B(s)--**∆H3**-->C(s) Applying Hess’ law of constant heat summation then:

A(s) B(s)

**∆**

**H**

**2**

**∆**

**H**

**1**

**∆**

**H**

**3**

C(s)

The above is called an **energy cycle diagram**. It can be used to calculate any of the missing energy changes since:

1. ∆H1 =∆H2 + ∆H3
2. ∆H2 =∆H1 + -∆H3
3. ∆H3 = - ∆H1 + ∆H2

Examples of applying Hess’ law of constant heat summation

**1.Calculate the molar enthalpy of formation of methane (CH4) given that ∆Hᶿc of carbon-graphite is -393.5kJmole-1,Hydrogen is -285.7 kJmole-1 and that of methane is 890 kJmole-1**

Working

Carbon-graphite ,hydrogen and oxygen can react to first form methane.

Methane will then burn in the oxygen present to form carbon(IV)oxide and water.

Carbon-graphite can burn in the oxygen to form carbon(IV)oxide.

Hydrogen can burn in the oxygen to form water.

C(s)+ 2H2 (g)+2O2 (g) --**∆H1**--> CH4(g) +2O2(g) --**∆H2**--> **CO2(g)+2H2O(l)**

C(s)+ 2H2 (g)+2O2 (g) --**∆H3**--> **CO2(g)+2H2O(l)**

Energy cycle diagram

C(s)

+

2

H

2

)

g

(

+

2

O

2

g

)

(

∆

H

1

=

∆

H

ᶿ

c

=

-

kJ

890.4

CH

4

)

g

(

+

2

O

2

g)

(

∆

H

3

=

∆

H

ᶿ

c

=

-

kJ

393.5

∆

H

3

=

∆

H

ᶿ

c

=

-

285.7

kJ x

2

∆

H

2

=

∆H

ᶿ

f

x

=

CO

2

)

g

(

+

2

H

2

O(l)

Substituting:

∆H3 = ∆H1 + ∆H2

-393.5 + (-285.7 x 2) = -890.4kJ + x

x = -74.5 kJ

Heat of formation ∆Hᶿf CH4 = -74.5 kJmole-1

**2. Calculate the molar enthalpy of formation of ethyne (C2H2) given : ∆Hᶿc of carbon-graphite = -394kJmole-1,Hydrogen = -286 kJmole-1 , (C2H2) = -1300 kJmole-1**

Working

Carbon-graphite ,hydrogen and oxygen can react to first form ethyne.

Ethyne will then burn in the oxygen present to form carbon(IV)oxide and water. Carbon-graphite can burn in the oxygen to form carbon(IV)oxide.

Hydrogen can burn in the oxygen to form water.

2C(s)+ H2 (g)+2 ½ O2 (g) --**∆H1**--> C2H2 (g) +2 ½ O2(g) --**∆H2**--> **CO2(g)+H2O(l)**

2C(s)+ H2 (g)+ 2 ½ O2 (g) --**∆H3**--> 2**CO2(g)+H2O(l)**

Energy cycle diagram

2C(s) + H2 (g) +2½O2(g) **∆H1**=∆Hᶿf =x CH +2½O2(g) ∆H3=∆Hᶿc =-394kJx 2 ∆H3=∆Hᶿc =-286kJ ∆H = ∆Hᶿc= -1300kJ

2

2

2

2CO2(g) + H2O(l)

Substituting:

∆H3 = ∆H1 + ∆H2

( -394 x 2) + -286 = -1300kJ + x

x = +244 kJ

Heat of formation ∆Hᶿf CH4 = **+244 kJmole-1**

**3. Calculate the molar enthalpy of formation of carbon(II)oxide (CO) given : ∆Hᶿc of carbon-graphite = -393.5kJmole-1, ∆Hᶿc of carbon(II)oxide (CO)= -283 kJmole-1**

Working

Carbon-graphite reacts with oxygen first to formcarbon (II)oxide (CO).

Carbon(II)oxide (CO) then burn in the excess oxygen to form carbon(IV)oxide. Carbon-graphite can burn in excess oxygen to form carbon (IV) oxide.

C(s)+ ½O2 (g) --**∆H1**--> CO (g) + ½ O2(g) --**∆H2**--> **CO2(g)**

C(s)+ O2 (g) --**∆H3**--> **CO2(g)**

Energy cycle diagram

C(s) + ½O (g) ∆H1=∆Hᶿf =x CO+½O2(g)

2

∆

H

2

=

∆H

ᶿ

c

=

-

283

kJ

∆H3=∆Hᶿc =-393.5kJ

CO2(g)

Substituting:

∆H3 = ∆H1 + ∆H2

-393.5kJ = -283kJ + x

x = -110 kJ

Heat of formation ∆Hᶿf CO = -110 kJmole-1

**4.Study the information below:**

**H2(g) + ½ O2(g)-> H2O(l) ∆H1=-286 kJmole-1**

**C(s) + O2(g)-> CO2(g) ∆H2=-393 kJmole-1**

**2C(s) + H2(g) + ½ O2(g)->C2H5OH(l) ∆H3=-277 kJmole-1 Use the information to calculate the molar enthalpy of combustion ∆H4 of ethanol**

Energy cycle diagram

2

C(s

)

+

3

H

2

(

g)

+3

½O

2

(

g

)

**∆**

**H**

**3**

=

∆

H

ᶿ

f

=

-

227

kJ

C

2

H

5

OH +3

O

2

g)

(

∆

H

2

=

∆H

ᶿ

c

=

-

394

J

k

x 2

∆

H

1

=

∆H

ᶿ

c

=

-

286

kJ

x 3

∆

H

4

=

∆H

ᶿ

c

=

x

**CO**

**2**

**2**

**g) +**

**(**

**3**

**H**

**2**

**O(l)**

Substituting:

∆H1 + ∆H2 = ∆H3 + ∆H4

( -394 x 2) + -286 x 3 = -277 + x

∆H4 = -1369 kJ

Heat of combustion ∆Hᶿc C2H5OH = **-1369 kJmole-1**

**5.Given the following information below:**

**CuSO4(s) + (aq)-> CuSO4(aq) ∆H=-66.1 kJmole-1**

**CuSO4(s) + (aq)+5H2O(l)-> CuSO4 .5H2O (aq) ∆H=-77.4 kJmole-1**

**Calculate ∆Hfor the reaction;**

**CuSO4(aq) + 5H2O-> CuSO4 .5H2O (aq) ∆H=-77.4 kJmole-1**

Working

**CuSO4(s) + (aq)+5H2O(l)-> CuSO4(aq)+ 5H2O(l)-> CuSO4 .5H2O (aq) CuSO4(s) + (aq)+5H2O(l)-> CuSO4 .5H2O (aq)**

Energy cycle diagram

CuSO4(s) + (aq)+5H2O(l) ∆H1=+66.1kJ CuSO (aq)+ 5H2O(l)

4

∆

H

2

=

x

∆H3= =-77.4kJ

CuSO4 .5H2O (aq)

Substituting:

∆H3 = ∆H2 + ∆H1

( -77.4kJ = x + +66.1kJ

∆H4 = **-10.9 kJ**

Heat of dissolution of CuSO4 = **-10.9kJmole-1**

Practically, Hess’ law can be applied practically as in the following examples

**a)Practical example 1**

Determination of the enthalpy of formation of CuSO4.5H2O

**Experiment I**

Weigh accurately 12.5 g of copper(II)sulphate(VI)pentahydrate. Measure 100cm3 of distilled water into a beaker. Determine its temperature T1 .Put all the crystals of the copper(II)sulphate(VI)pentahydrate carefully into the beaker. Stir using a thermometer and determine the highest temperature change T2 Repeat the procedure again to complete table 1.

Table 1:Sample results

|  |  |  |
| --- | --- | --- |
| Experiment | I | II |
| Highest /lowest temperature T2 | 27.0 | 29.0 |
| Initial temperature T1 | 24.0 | 25.0 |
| Change in temperature ∆T | 3.0 | 4.0 |

**Experiment II**

Weigh accurately 8.0g of anhydrous copper(II)sulphate(VI). Measure 100cm3 of distilled water into a beaker. Determine its temperature T1 .Put all the crystals of the copper(II)sulphate(VI)pentahydrate carefully into the beaker. Stir using a thermometer and determine the highest temperature change T2 Repeat the procedure again to complete table II.

Table II :Sample results

|  |  |  |
| --- | --- | --- |
| Experiment | I | II |
| Highest /lowest temperature T2 | 26.0 | 27.0 |
| Initial temperature T1 | 25.0 | 25.0 |
| Change in temperature ∆T | 1.0 | 2.0 |

**Questions**

**(a)Calculate the average ∆T in**

**(i)Table I**

∆T= T2 -T1 => 3.0 +4.0 = **3.5 oC**



**(ii)Table II**

∆T= T2 -T1 => 1.0 +2.0 = **1.5 oC**

2

**(b)Calculate the number of moles of solid used in:**

**(i)Experiment I**

Moles of CuSO4.5H2O = Mass => 12.5 = **0.05 moles**

Molar mass 250

**(ii)Experiment II**

Moles of CuSO4 = Mass => 8.0 = **0.05 moles**

Molar mass 160

**(c)Calculate the enthalpy change for the reaction in:**

**(i)Experiment I**

Enthalpy change of CuSO4.5H2O= mass of Water(**m**) x **c** x **∆T**

=>100cm3 x 4.2 x 3.5 oC = **-1.47kJ**

1000

**(ii)Experiment II**

Enthalpy change of CuSO4 = mass of water(**m**) x **c** x **∆T**

=>100cm3 x 4.2 x 1.5 oC = **-0.63kJ**

1000

**(c)Calculate the molar enthalpy of solution CuSO4 .5H2O (s) form the results in (i)experiment I.**

∆Hs = CuSO4.5H2O= ∆H => -1.47kJ = **29.4kJ**

Number of Moles 0.05 moles

**(ii)experiment II.**

∆Hs = CuSO4= ∆H => -0.63kJ = **12.6kJ**

Number of Moles 0.05 moles

**(d) Using an energy level diagram, calculate the molar enthalpy change for the reaction:**

**CuSO4 .5H2O (s) -> CuSO4(s)+5H2O(l)**

Energy cycle diagram

CuSO4(s) + (aq)+5H2O(l) ∆H1=x CuSO . 5H O (s)+ (aq)

4

2

∆H3= =-29.4kJ ∆H2= -12.6kJ

CuSO4 .5H2O (aq)

∆H3 = ∆H1 +∆H2

=>-29.4kJ = -12.6kJ + x

=>-29.4kJ - (+12.6kJ) = x **x = 16.8kJ**

**b)Practical example II**

Determination of enthalpy of solution of ammonium chloride

Theoretical information.

Ammonium chloride dissolves in water to form ammonium chloride solution. Aqueous ammonia can react with excess dilute hydrochloric acid to form ammonium chloride solution. The heat change taking place can be calculated from the heat of reactions:

1. NH3(aq) + HCl(aq) -> NH4Cl(s)
2. NH4Cl(s) + (aq) -> NH4Cl(aq)
3. NH3(aq) + HCl(aq) -> NH4Cl(aq)

Experiment procedure I

Measure 50cm3 of water into a 100cm3 beaker. Record its temperature T1 as initial temperature to the nearest 0.5oC in table I. Add exactly 5.0g of ammonium chloride crystals weighed carefully into the water. Stir and record the highest temperature change T2 as the final temperature change. Repeat the above procedure to complete table I.

Sample results TableI

|  |  |  |
| --- | --- | --- |
| Experiment | I | II |
| final temperature(oC) | 19.0 | 20.0 |
| initial temperature(oC) | 22.0 | 22.0 |
| temperature change ∆T(oC) | 3.0 | 2.0 |

Experiment procedure II

Measure 25cm3 of 2M aqueous ammonia into a 100cm3 beaker. Record its temperature T1 as initial temperature to the nearest 0.5oC in table II. Measure 25cm3 of 2M hydrochloric acid solution. Add the acid into the beaker containing aqueous ammonia. Stir and record the highest temperature change T2 as the final temperature change. Repeat the above procedure to complete table II.

Sample results:Table II

|  |  |  |
| --- | --- | --- |
| Experiment | I | II |
| final temperature(oC) | 29.0 | 29.0 |
| initial temperature(oC) | 22.0 | 22.0 |
| temperature change ∆T(oC) | 7.0 | 7.0 |

**Sample Calculations: (a)Calculate the average ∆T in**

**(i)Table I**

∆T= T2 -T1 => -3.0 +-2.0 = **2.5 oC**

2

**(ii)Table II**

∆T= T2 -T1 => 7.0 +7.0 = **7.0 oC**

2

**(b)Calculate the enthalpy change for the reaction in:**

**(i)Experiment I**

Enthalpy change ∆H = mass of Water(**m**) x **c** x **∆T**

=>50cm3 x 4.2 x 2.5 oC = **+0.525kJ**

1000

**(ii)Experiment II**

Enthalpy change of CuSO4 = mass of water(**m**) x **c** x **∆T**

=>25+25cm3 x 4.2 x 7 oC = **+1.47kJ**

1000

**(c)Write the equation for the reaction taking place in:**

**(i)Experiment I**

NH4Cl(**s**) + (aq) -> NH4Cl(**aq**)

**(ii)Experiment I**

NH3(aq) + HCl(aq) -> NH4Cl(aq)

**(d)Calculate the enthalpy change ∆H for the reaction:**  NH3(g) + HCl(g) -> NH4Cl(s) given that:

1. NH3(g) + (aq) -> NH3(aq) ∆H= -40.3kJ
2. (aq) + HCl(g) -> HCl(aq) ∆H= -16.45kJ

**(e)Applying Hess’ Law of constant heat summation:**

Energy level diagram

NH4Cl(s) + aq

**+0.525kJ**=∆H4

**- 40.3kJ**=∆H1  **-16.43kJ**=∆H2

N

2

(

g)

+

1½ H

2

(

g)

½ Cl

+

2

**∆**

**H**

**f**

(

aq

)

(

aq

)

NH3 (aq) + HCl(aq) **-1.47kJ**=∆H3 NH4Cl(s)

∆H1 + ∆H2 + ∆H3 = ∆H4 + ∆Hf **- 40.3kJ + -16.43kJ + -1.47kJ = +0.525kJ + ∆Hf**

=>**∆Hf = -58.865kJ.**

Practice theoretical examples:

**1. Using an energy level diagram calculate the ∆Hs  of ammonium chloride crystals given that.**

**∆Hf  of NH3 (aq) = -80.54kJ mole-1**

**∆Hf  of HCl (aq) = -164.46kJ mole-1**

**∆Hf  of NH4Cl (aq) = -261.7483kJ mole-1**

**∆Hs  of NH4Cl (aq) = -16.8517kJ mole-1**

N2(g) + 1½ H2(g) + ½ Cl2 **∆Hf**=-261.7483kJ NH Cl(s) + aq

4

**x**

=

∆

H

s

)

aq

(

)

aq

(

**-**

**80.54**

**kJ**

=

∆

H

1

**-**

**16**

**4.46**

**kJ**

=

∆

H

2

NH3 (aq) + HCl(aq)  **16.8517kJ**=∆H3 NH4Cl(s)

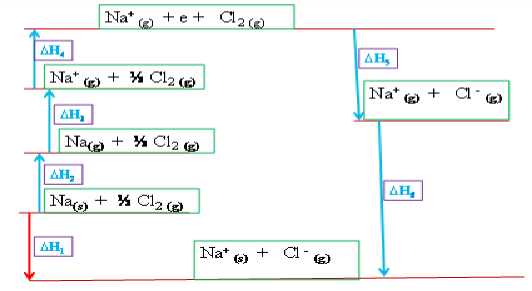
∆H1 + ∆H2 + ∆H3  = ∆H4  + ∆Hf

- 80.54kJ + -164.46kJ + -16.8517kJ = -261.7483kJ + ∆Hf

=>∆Hf = **-33.6kJmole-1**.

Study the energy cycle diagram below and use it to:

(a)Identify the energy changes ∆H1 ∆H2 ∆H3 ∆H4 ∆H5 ∆H6



∆H1  - enthalpy/heat of formation of sodium chloride (∆Hf)

∆H2 -enthalpy/heat of atomization of sodium (∆Hat)

∆H3  -enthalpy/heat of ionization/ionization energy of sodium (∆H i)

∆H4 -enthalpy/heat of atomization of chlorine (∆Hat)

∆H5 -enthalpy/heat of electron affinity of chlorine (∆He)

∆H6 enthalpy/heat of lattice/Lattice energy of sodium chloride(∆H l)

1. **Calculate ∆H1  given that ∆H2 =+108kJ , ∆H3=+500kJ, ∆H4 =+121kJ ,∆H5**

**=-364kJ and ∆H6 =-766kJ**

Working:

∆H1 =∆H2 +∆H3 +∆H4 +∆H5 +∆H6

Substituting:

∆H1= +108kJ + +500kJ + +121kJ +-364kJ + -766kJ

∆H1= **-401kJmole-1**

1. **Given the that:**

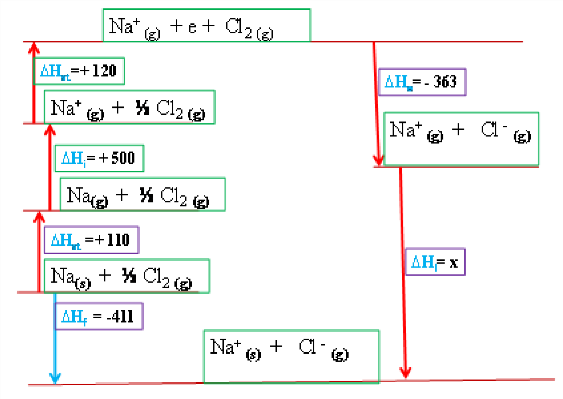
**(i) Ionization energy of sodium = + 500kJmole-1**

**(ii)∆Hat of sodium = + 110kJmole1**

**(iii) Electron affinity of chlorine = - 363kJmole-1**

**(iv)∆Hat of chlorine = + 120kJmole-1**

**(v) ∆Hf  of sodium chloride = -411kJ , calculate the lattice energy of sodium chloride using an energy cycle diagram.**



Working:

Applying Hess law then: ∆Hf =∆Ha  +∆Hi +∆Ha +∆He +∆Hl  Substituting:

-411= +108kJ + +500kJ + +121kJ +-364kJ + x -411 + -108kJ + -500kJ + -121kJ + +364kJ = x x= **-776kJmole-1**

