CHEMISTRY NOTES

FORM 2 NOTES

STRUCTURE OF THE ATOM AND THE PERIODIC TABLE

Specific objectives

By the end of the topic, the learners should be able to:

a) Describe the structure of the atom and write electron configuration.

b) Define atomic number, mass number, isotope and relative atomic mass.

c) Determine relative mass from isotopic composition.

d) Use electron configuration to determine the position of an element in the periodic table the type of ion it forms.

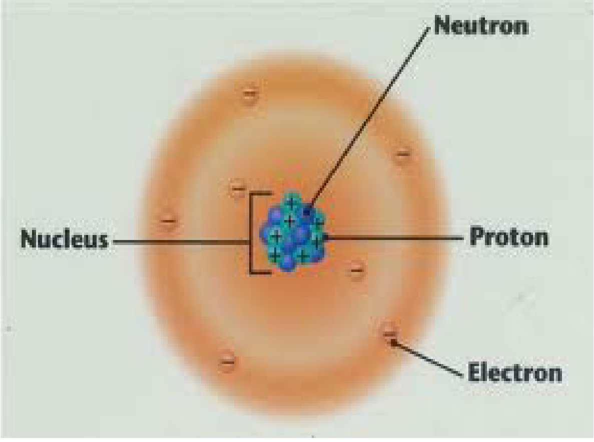
e) Define valency and oxidation number of an element.

f) Use valency to drive chemical formulae of compounds.

g) Write balanced chemical equations.

The structure of an atom

- Matter is composed of the simplest substance, the elements, which cannot be broken down into anything simpler by any chemical process.
- The smallest particle of an element is called the atom.
- Atoms are extremely small, far too small to be seen.
- An atom has a radius of about 10^{-8}cm with hydrogen atoms being the smallest.



Simplest structure of the atom

- Each atom is made up of two regions;
  - (i) The nucleus
  - (ii) The shells {energy levels}

[Diagram pg 8, NM patel bk2]

- The nucleus is very small compared to the volume of the atom with approximately 10^{-5} of the diameter of the atom.
- The atom is made up of three different sub-atomic particles:
  - (i) Protons
  - (ii) Neutrons
  - (iii) Electrons

(i) Protons

- They are present in the nucleus of every atom.
- They are positively charged particles. {one unit positive charge, +1}
- Has a relative mass of 1.

(ii) Neutrons

- They are present in the nucleus of all atoms except hydrogen.
- They are particles with no charge.
- They have a relative mass of 1.

NB: The particles (protons and neutrons) within the nucleus are collectively called nucleons.

(iii) Electrons

- The energy levels contain the negatively charged particles called electrons.
- These particles spin around the nucleus in definite orbits called energy levels.
- They are very light compared to the nucleus and have a relative mass of $\frac{1}{1840}$ (almost 0).
- They have a unit negative charge (-1).

- The table below summarizes the characteristics of the subatomic particles;

Particles	**Symbol**	**Mass relative to hydrogen atom**	**Relative charge**
Proton	P	1	+1
Neutron	n	1	0
Electron	e	$\frac{1}{1840}$ (almost 0)	-1

Electron energy levels in an atom

- Different elements have different chemical properties because;
    (i) They are different number of electrons.
    (ii) Electrons are arranged differently around the nucleus of their atoms.
- Electrons are arranged in energy levels around the nucleus of an atom.
- The energy levels or shells are represented as circles around the nucleus and are labeled 1, 2, 3 etc.

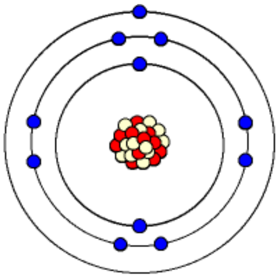
Number of electrons in an energy level.

- Each energy level can hold a certain maximum number of electrons.
- It can hold less than the maximum but never more.
- The first shell, nearest to the nucleus, is the lowest energy level and electrons fill this energy level first.

- A summary of the maximum electrons that can be held by each energy level is given in the table below;

Energy levels	**Maximum electrons**
1	2
2	8
3	18

- In general, the number of maximum electrons that an energy level can hold is given by $2n^2$, where n is the number of the energy level e.g. for energy level 3, n=3 and therefore the maximum electrons that can reside in it are $2(3)^2$=18electrons e.g. the atom of sodium with atomic number 11;



Filling energy levels

- The electrons fill the energy levels starting from the lowest energy level i.e. 1 which can hold a maximum of 2 electrons.

  Examples

1. Show diagrammatically the electron filling in the following atoms;
   a) Hydrogen, H with 1electron.
   b) Helium, He with 2 electrons.
   c) Potassium, K with 19electrons.
   d) Neon, Ne with 10electrons.

NB:

(i) The 3rd energy level can hold a maximum of 18electrons. However, the 4th energy level begins to fill before the 3rd one is full. Elements with atomic number 1-20 can hold a maximum of 8 electrons in the 3rd energy level.

(ii) From the examples above Ne (2.8), it can be seen that the outermost energy level contain a maximum of 8electrons. This shows a stable electron arrangement.

Atomic characteristics

Atomic number

- This is the number of protons in an atom and is denoted by **Z.**
- A neutral atom has equal number of protons and electrons and the Z refers also tom the number of electrons in that atom e.g. H has 1 proton and hence its atomic number is 1; Potassium has 19 protons and hence Z=19.

Mass number

- The total number of protons and neutrons in the nucleus of an atom is called the mass number and is denoted by **A** i.e.
  - Mass number, A=number of protons + number of neutrons.
  - A=atomic number + number of neutrons.
  - A=Z+N {where N denotes the number of neutrons.}

Examples

1. A neutral has mass number 39 and the number of electrons is 20. What is the atomic number? Give the number of protons and neutrons in the nucleus.

Solution

- A=Z+N; 39=20+N => N=19neutrons,protons=20,electrons=20

2. Copy and complete the following table.

Element	**Symbol**	**Number of electrons**	**Number of protons**	**Number of 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	4	4	
Boron	B	5	5			11

Carbon	C	6				12
Nitrogen	N		7	7		
Oxygen	O				8	16
Fluorine	F				9	19
Neon	Ne	10				20
Sodium	Na			12	11	
Magnesium	Mg		12	12		
Aluminium	Al	13		14		
Silicon	Si			14		28
Phosphorous	P		15	16		
Sulphur	S	16				32
Chlorine	Cl	17		18		
Argon	Ar		18		18	40
Potassium	K	19		20		
Calcium	Ca		20	20		

Symbols

- The atomic mass number is written as a superscript on the top left hand corner of the symbol for the atom i.e.

$^{\text{Mass number}}\text{Symbols}$ e.g. ^{12}C

- The atomic number is written as a subscript on the bottom left hand corner of the symbol for the atom i.e.

$_{\text{Atomic number}}\text{Symbol}$ e.g. $_{6}C$

- A combination of these two numbers is given as;

$$^{\text{Mass number}}_{\text{Atomic number}}\text{Symbol} \rightarrow {}^{A}_{Z}\text{Symbol} \rightarrow {}^{12}_{6}C$$

Examples

1. What information can you get from the symbol ${}^{31}_{15}\mathrm{P}$ about the structure of the phosphorous atom?

Solution

Mass number=31

Atomic number=15

Number of protons=15

Number of electrons=15

Number of neutrons=31-15=16

Electron arrangement {Electron configuration}

- This is the distribution of electrons in energy levels of an atom.
- This can be represented as; 1 for hydrogen, 2 for helium, 2.1 for lithium and 2.8 for neon.

Example

1. Write the electron arrangement of the atoms with the following atomic numbers.
   a) K=19 =>2.8.8.1
   b) C=6 =>2.4
   c) O=8 =>2.6
   d) Ca=20 =>2.8.8.2

NB:

(i) One can tell from the electron arrangement the number of energy levels an element has.

(ii) Elements with the same number of occupied energy levels belong to the same period e.g. in the above example; K and Ca belong to the period 4 while C and O belong to period 2.

Isotopes

- These are atoms which have the same atomic number but different mass number i.e. they have the same number of protons but different number of neutrons.

Example

1. Which of the following elements X, Y and Z are isotopes?

$^{39}_{19}X$, $^{40}_{19}Y$, $^{39}_{20}Z$

Solution

⇨ X and Y

Relative atomic mass

- This is the average mass of a large number of atoms of an element to that of a carbon-12 atom.
- It is normally calculated from the relative abundance of isotope of an element.

Examples

1. Calculate the R.A.M of chlorine from it isotope given that a mixture of chlorine had two isotopes, ^{35}Cl and ^{37}Cl in the ratio 3:1.
2. Naturally occurring Gallium consists of the isotope ^{69}Ga and ^{71}Ga in the ratio 3:2. Calculate the R.A.M of Gallium.

Periodic table

- The periodic table is a grid of horizontal and vertical lines of elements.
- The horizontal lines are called periods represented by numerical letters 1, 2 etc and vertical column called groups are represented by Roman numerals, I to VII.

[Get a periodic table]

- The modern periodic table is based on the atomic structure i.e. the number of proton in the atom of an element.

Building up of the periodic table for the first twenty elements

- Elements with the same number of electrons in the outermost energy level form a group e.g. H, Li, Na, K have one electrons in the outermost energy hence belong to group I.
- Elements with the same energy level can be arranged in the same vertical column which represents a group.
- There are eight groups numbered in Roman numerals.
- The first twenty elements are arranged into their respective groups with group VIII being referred to as group zero because these elements have little tendency to lose or gain

electrons during reaction.

Group I 1st column	**Group II 2nd column**	**Group III 3rd column**	**Group IV 4th column**	**Group V 5th column**	**Group VI 6th column**	**Group VII 7th column**	**Group VIII 8th column**
H	Be	B	C	N	O	F	He
Li	Mg	Al	Si	P	S	Cl	Ne
Na	Ca						Ar
K							

- The simplified form of the periodic table is as shown below;

H								**He**
Li	**Be**	**Transition metals**	**B**	**C**	**N**	**O**	**F**	**Ne**
Na	**Mg**		**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
K	**Ca**							

Note

(i) A) **Groups:** Elements which are arranged vertical in the periodic table form groups I to VIII (or 0). Elements of the same group have the same number of electrons in the outer shell and they have similar chemical properties.

B) **Periods:** Elements which are arranged horizontally from the left to the right across the table form the periods. Elements in the same period have the same number of energy levels (shells)

(ii) Between group II and III lies a group of elements referred to as the Transition metals

Which have unique characteristics that make them not to fit in the eight groups of the periodic table? [The real periodic table]

Ion formation

- Atoms with electron configuration of 2, 2.8 and 2.8.8 are said to be stable.
- Atoms without this electron configuration tend to attain this stability by either losing or

gaining electrons.

- Sodium for instance, with atomic number 11 has an electron configuration 2.8.1 can acquire the stable electron arrangement of 2.8 by losing one electron or 2.8.8 by gaining seven electrons.
- It requires less energy to lose one electron than gaining seven and therefore sodium forms a stable configuration by losing an electron.
- The resulting sodium particle will have only ten electrons while the nucleus still has 11 protons hence the net charge is +1 written as Na^+.
- The ion formation of sodium can be represented diagrammatically as follows;

$$Na \rightarrow Na^+ + e^-$$

Further examples

1. Show ion formation of the following atoms
    a) Cl=17
    b) Mg=12
    c) O=8

- The charged particles formed when atoms gain or lose electrons are called ions.
- The positively charged ions are called cations while the negatively charged ions are called anions.
- Metals mostly lose electrons to form positively charged ions (cations) while non-metals gain electrons to form negatively charged ions (anions).
- During chemical reactions atoms of metals react by losing electrons from their outermost energy level while non-metals generally gain electrons into their outermost energy level.
- The number of electrons an atom loses or gains during a chemical reaction is the valency of the atom and is also known as the combining power of an element.
- The valency generally correspond to the group in which the element is placed in the periodic table e.g. group I and VII have a valency of 1 and group II and VI have a valency of 2.

NB: Some elements have variable valences e.g. Iron which has a valency of either 2 or 3, Copper has 1 or 2 and lead has 2 or 4.

Radicals

- These are groups of atoms with a net charge that exist and react as a unit during chemical reactions.
- The valency of a radical is the same as the value of its charge e.g. Sulphate ion (SO_4^{2-}) has a valency of 2, a Nitrate ion (NO_3^-) has a valency of 1, an ammonium ion (NH_4^+)has a valency of 1.
- The table below summarizes the valency of some common elements and radicals

a) Metals/ Non-metals

	Valency I	**Valency 2**	**Valency 3**
Metals	Sodium	Calcium	
	Potassium	Barium	
		Magnesium	
		Zinc	Aluminium
		Iron	Iron
		Lead	
		Copper	
Non-metals	Nitrogen		
	Chlorine	Nitrogen	Nitrogen
	Fluorine	Oxygen	Phosphorous
	Hydrogen	Sulphur	

b) Radicals

	Valency 1	**Valency 2**	**Valency 3**
Radicals	Ammonium (NH_4^+)	Carbonate (CO_3^{2-})	Phosphate(PO_4^{3-})
	Hydroxide (OH^-)	Sulphate (SO_4^{2-})	
	Nitrate (NO_3^-)	Sulphite (SO_3^{2-})	

	Hydrogen carbonate (HCO_3^-) Hydrogen Sulphate (HSO_4^-)		

Oxidation number

- The oxidation number of an element shows the number of electrons which have been removed or added to it to get the present state.
- There are positive and negative oxidation numbers.
- Since atoms are electrically neutral, they are assigned an oxidation of zero.
- The table below shows the oxidation number of some ions.

Particles	**Oxidation number**
Mg^{2+}	+2
Fe^{2+}	+2
Fe^{3+}	+3
H^+	+1
Cu^+	+1
Cu^{2+}	+2
Cl^-	-1
Cu	0
Mg	0
H	0

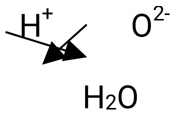
NB: Oxidation number is written with the sign coming before the number while charge is written as a superscript with the number coming before the positive or negative sign e.g. Mg^{2+} or N^{3-}.

Chemical formulae

- A chemical formula is a representation of a chemical substance using chemical symbols.
- It shows the constituent elements and the proportions in which they are combined.

- In order to write a correct formula, it is necessary to know the symbols and valencies of the elements which form the compound.
- When writing the formula, always start with the element which is more likely to lose an electron or electrons then follow with the one that gains e.g.

The formula of water;



Examples

1. Write the chemical formula for the following compounds

   (a) Calcium fluoride =>CaF_2

   (b) Carbon (II) oxide => CO

   (c) Carbon (IV) oxide=>CO_2

   (d) Lead (IV) oxide=>PbO

   (e) Aluminium carbonate=>$Al_2(CO_3)_3$

   (f) Potassium nitrate=>KNO_3

   (g) Calcium hydrogen sulphate=> $Ca(HCO_3)_2$

Chemical equations

- A chemical equation represents a chemical change by use of symbols and formulae.
- The formula of the reactants on the left hand side and those of the products on the right hand side e.g. copper metal reacts with oxygen to form copper (II)n oxide; The word equation is give as;

Copper metal + oxygen gas → Copper (II) oxide

⇨ The equation which represents the reaction can be written in form of symbols as follows;-

$2Cu+O_2 \rightarrow 2CuO$

Balancing chemical equations

- An equation is balanced when the number of atoms of each type of reactants is equal to that of the products side.

- The following guidelines should be observed when balancing chemical equations.

  (i) Write the chemical equation in words.

  (ii) Write the correct chemical formulae for both reactants and products.

  (iii) Check whether the number of atoms of each element on the reactant side is equal to that on the product side.

  (iv) If the number is not equal, multiply the chemical formula containing the unbalanced atoms with the lowest common multiple.

  (v) Check again to ensure that all the atoms are balanced. **<u>NB:</u>** The chemical formulae of the reactants and products should never change during balancing of the equations.

  (vi) A chemical equation is correct and complete if the physical states of the reactants and the products are indicated. The following symbols represent the physical states which are written in brackets after each reactant or product;

  Solid (s), liquid (l), in solution/aqueous (aq), gas (g)

  e.g. $2Cu_{(s)} + O_{2(g)} \rightarrow 2CuO_{(s)}$

  <u>Examples</u>

1. Write a balanced chemical equation by completing the following equations;

   a) Zinc granules + dilute hydrochloric acid →

   b) Sodium hydroxide + dilute hydrochloric acid →

   c) Zinc oxide + dilute Sulphuric (VI) acid →

   d) Zinc metal + dilute nitric (V) acid →

   e) Calcium hydroxide + Carbon (IV) oxide →

   f) Sodium + water →

2. Balance the following equations;

   a) $Mg_{(s)} + HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

   b) $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2(g)}$

   c) $NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$

d) $CuCO_{3(s)} + HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

e) $H_2S_{(g)} + O_{2(g)} \rightarrow SO_{2(g)} + H_2O_{(l)}$

f) $C_2H_{6(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

3. Write a balanced equation for each of the following reactions;

   a) Heating of sodium metal in oxygen gas to obtain sodium oxide.

   b) Magnesium metal with steam.

   c) Action of dilute hydrochloric acid on Calcium carbonate.

   d) Calcium metal and water.

   e) Copper (II) oxide solid with dilute Sulphuric (VI) acid.

CHEMICAL FAMILIES;PATTERNS AND PROPERTIES

Specific objectives

By the end of this topic, the learners should be able to:

a) Indentify and write electron arrangement of alkali metals, alkaline earth metals, halogens and noble gases.

b) State and explain the trends in physical properties of elements in group I, II, VII and VIII.

c) State and explain the trends in reactivity of elements in group I, II, VII and VIII.

d) Explain the similarities in chemical formulae of compounds of the elements in a group.

e) Explain the reactive nature of group VII elements.

f) Indentify and write electron arrangement of period 3 elements.

g) State and explain the trends in physical and chemical properties of the elements in period 3.

Introduction

- Element in the same group are said to belong to the same chemical family, they include;

    a) Alkali metals {Group I}

b) Alkaline earth metals{Group II}

c) Halogens {Group VII}

d) Noble gases{Group VIII}

a) Alkali metals

- These are elements in group I of the periodic table.
- They include;
  - Lithium [Li]
  - Sodium [Na]
  - Potassium [K]
  - Rubidium [Rb]
  - Caesium [Cs]
  - Francium [Fr]
- The electron arrangements of the first three alkali metals are as follows;

$_{3}Li=2.1$

$_{11}Na=2.8.1$

$_{19}K=2.8.8.1$

- Each alkali metal atom has one electron in the outermost energy level.
- Down the group there is an increase in the number of occupied energy levels.

Trend in properties of alkali metals

1. Gradation in size of the atom and ions

- The table below summarizes the atomic and ionic radius of 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

- From the table above, the following can be noted;

i. Atomic radius

- **Definition:** This is the distance between the centre of the nucleus of an atom and the outermost energy level occupied by an electron or electrons.
- It increase down the group; **This is because each alkali metal has one more occupied energy level than the preceding member in the group** i.e. Li has two energy levels, Na=3,K=4.Therefore the outmost electron in a sodium atom is further from the nucleus than the outermost electron in a lithium atom.

ii. Ionic radius

- **Definition**: This is the distance between the centre of the nucleus of an ion and the outermost energy level occupied by an electron or electrons.
- It also increases down the group; ***This is because there is an increase in energy level just like the atoms.***

NB: The atomic radius is greater than ionic radius of the same element; **This is because when an atom loses an electron to form positively charged ion, the remaining electrons experience greater nuclear attraction. The remaining energy levels move closer to the nucleus resulting in a reduction in the radius.**

2. Physical properties

- The table shows some physical properties of alkali metals.

Elements	**Appearance**	**Ease of cutting**	**Melting point (^{0}C)**	**Boiling point (^{0}C)**	**Electrical conductivity**	**Atomic radius (nm)**	**1st ionization energy (Kj.mol^{-1})**
Lithium	Silvery white	Slightly hard	180	1330	Good	0.133	520
Sodium	Shiny grey	Easy	98	890	Good	0.157	496
Potassium	Shiny grey	Easy	64	774	Good	0.203	419

- From the table the following physical properties can be deduced;

a) The alkali metals have a shiny metallic luster when freshly cut. However, the surface quickly tarnished because of reacting with air.

b) The alkali metals are soft and easy to cut. The softness and ease to cut increase down the group. ***This is due to the decrease in the strength of the forces holding the atoms together as you move down the group.***

c) The alkali metals have relatively low melting and boiling points. They decrease down the group**; This is due to the weakening of the forces holding the atoms together.**

d) Alkali metals are good conductors of both heat and electricity. ***Conductivity in this chemical family is due to the presence of delocalized electrons in the structure of the metal.*** Since they all have one electron in their outermost energy level, their conductivity is similar.

e) **Ionization energy** which is the minimum energy required to remove an electron from the outermost energy level of an atom in the gaseous state.

- The ionization energy decrease down the group. This means that less energy is required to remove an electron from the outermost energy level of potassium than lithium; ***This is because the effective force of attraction on the outermost electron by the positive nucleus decrease with increasing atomic size and distance from the nucleus.***

3. Chemical properties

i. Group one elements react vigorously with both air and water with the reactivity increasing down the group i.e. potassium is more reactive than sodium and lithium.

- When sodium is exploded in air, it reacts with moisture in air to form sodium hydroxide.

$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

- The NaOH formed further reacts with CO_2 in the air to form Na_2CO_3;

$$2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_3.H_2O_{(s)}$$

- Sodium burns in air with **a yellow flame** to form a yellowish white solid which is mainly sodium oxide;

$$4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$$

- When sodium burns in air enriched with oxygen it forms mainly sodium peroxide;

$$2Na_{(s)} + O_{2(g)} \rightarrow Na_2O_{2(s)}$$

- Potassium burns in air with a **lilac flame** to form a white solid which is potassium oxide;

$$4K_{(s)} + O_{2(g)} \rightarrow 2K_2O_{(s)}$$

- The table below shows the reaction of alkali metals with water.

Metal	**Observation when the metal reacts with water**	**Rate of reaction**
Lithium	Lithium floats in water. Colourless gas is produced. The gas does not ignite spontaneously. The resulting solution turns red litmus paper blue.	Vigorous
Sodium	Sodium metal darts on the surface of water as it melts into a silvery ball. A hissing sound is produced. A colourless gas is produced which may ignite spontaneously. The solution formed is alkaline.	More vigorous
Potassium	The metal darts about on the surface of water as it melts into a silvery ball. A colourless gas is produced which spontaneously burst into flame. Potassium vapour burns with a lilac flame. The resulting solution is alkaline.	Explosive

<u>NB:</u> - Alkali metals react with water to form alkaline solution and hydrogen gas i.e.

Alkali metals + water → metal hydroxide + hydrogen gas

e.g. $2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$

$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$

$2K(s) + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$

- Of the three alkali metals, potassium is the most reactive followed by sodium and then lithium**; This is because the electron in the outermost energy level is loosely held by the nucleus and therefore is easily removed during a reaction.**

ii. Alkali metals react with chlorine gas to form the corresponding metal chlorides i.e.

Alkali metals + chlorine gas → chlorides of alkali metals

e.g. $2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

$2Li_{(s)} + Cl_{2(g)} \rightarrow 2LiCl_{(s)}$

$2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$

- The reactivity of alkali metals with chlorine increases down the group; **This because of the increase in atomic radius which leads to increase in ease to lose the electron in the outermost energy level.**

Uses of alkali metals and their compounds

1. Sodium is used in manufacture of sodium cyanide for use in the extraction of gold.
2. Lithium is used in the manufacture of special high strength glasses and ceramics.
3. Lithium compounds are used in the manufacture of dry cells for use in mobile phones, laptops, stop watches and zero emission electric vehicles.
4. A molten mixture of sodium and potassium is used as a coolant in nuclear reactors.
5. Sodium vapour is used to produce the yellow glow in streetlights.
6. Molten sodium is used as a reducing agent in the extraction of titanium (IV) chloride;

   $TiCl_{4(g)} + 4Na_{(l)} \rightarrow Ti_{(s)} + 4NaCl_{(l)}$

7. Sodium chloride is used as a food additive.
8. A mixture of NaOH (caustic soda) and carbon disulphide is used in the manufacture of artificial silk called rayon.

b) Alkaline earth metals

- These are members of group II of the periodic table.
- They include;
  - Beryllium [Be]

  - Magnesium [Mg]
  - Calcium [Ca]
  - Strontium [Sr]
  - Barium [Ba]
  - Radium [Ra]

- The electron arrangement of the first three alkaline earth metals are as follows;
  - $_{4}Be=2.2$
  - $_{12}Mg=2.8.2$
  - $_{20}Ca=2.8.8.2$
- An atom of an alkaline earth metal has two electrons in the outermost energy level.

<u>Trends in properties of alkaline earth metals</u>

1. **<u>Gradation in size of atoms and ions;</u>**

- The table below summarizes the atomic and ionic radius size of alkaline earth metals.

Element	Symbol	Atomic no.	Atomic arrangement	Atomic radius (nm)	Ionic radius (nm)
Beryllium	Be	4	2.2	0.089	0.031
Magnesium	Mg	12	2.8.2	0.136	0.065
Calcium	Ca	20	2.8.8.2	0.174	0.099

- From the table, the following can be noted;

  a) Atomic radius

  - The atomic radius increases down the group as more energy levels are occupied.

  b) Ionic radius

  - Alkaline earth metals form ions by losing the two electrons in the outermost energy level in order to attain a stable electron arrangement.
  - The loss of two electrons in the outermost energy level accounts for the smaller ionic radius compared to the atomic radius of the corresponding atom.

2. Physical properties of alkaline earth metals

- The table below summarizes the physical properties of alkaline earth metals.

Element	**Atomic no.**	**Melting point(^{0}C)**	**Boiling point (^{0}C)**	**Atomic radius (nm)**	**1st I.E (Kj.mol^{-1})**	**2nd I.E (Kj.mol^{-1})**
Beryllium	4	1280	2450	0.089	900	1800
Magnesium	12	650	1110	0.136	736	1450
Calcium	20	850	1140	0.174	590	1150
Strontium	38	789	1330	0.210	550	1060
Barium	56	725	1140	0.220	503	970

- The following can be noted from the table;

  a) The alkaline earth metals have relatively high melting and boiling points than alkali metals; ***This is because as one moves from left to right, in the periodic table there is addition of protons in the nucleus and hence increase in the force of attraction for the outermost electrons.***

  b) The melting and boiling point of beryllium are very high compared to other alkaline earth metals; ***This is because the beryllium atom is very small and the force of attraction between the atoms are very strong.***

  c) Down the group the melting and boiling point decrease; ***This is because in metal atoms are held together by forces of attraction between positive nucleus and delocalized electrons. As the atomic radius increases, this attraction decreases because of the increasing distance from the positive nucleus to the delocalized***

electrons.

d) The second ionisation energy is always higher than the first ionization energy; ***Once an electron has been lost from the atom, the overall positive charge holds the remaining electron more firmly. This then means that removing a second electron from the ion require more energy than the first electron.***

3. Chemical properties of alkaline earth metals.

a) Alkaline earth metals burn in air to form white oxides i.e.

Alkaline earth metals + oxygen → metal oxide

e.g. – magnesium burns in air with **a brilliant white** light which has a blinding glare forming a white solid; $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

{grey} {white}

- Calcium burns with a **faint orange-red** light forming a white solid;

$$2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$$

b) Alkaline earth metals react with water to give corresponding hydroxides and hydrogen gas i.e. Alkaline earth metals + water → metal hydroxide + hydrogen gas.

e.g. – Magnesium reacts slowly with cold water to form small bubbles of hydrogen gas which stick on the surface of the metal;

$$Mg_{(s)} + 2H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$

- A steady stream of hydrogen is evolved when calcium reacts with cold water and a white suspension appears in the beaker due to formation of calcium hydroxide which is sparingly soluble in water; $Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$

NB: Reacting between alkaline earth metals and water are faster if warm water is used e.g. Magnesium burns in steam with a bright white light forming a white solid which is magnesium oxide; $Mg_{(s)} + H_2O_{(l)} \rightarrow MgO_{(s)} + H_{2(g)}$

[Diagram pg 40,klb bk 2]

c) Alkaline earth metals react with chlorine to form their corresponding chloride i.e.

Alkaline earth metal + chlorine gas → metal chloride.

e.g. -Magnesium burns with a brilliant white flame to form a fine white solid which is magnesium chloride; $Mg_{(s)} + Cl_{2(g)} \rightarrow MgCl_{2(s)}$

- Calcium may not react readily with chlorine; This is because a coat of calcium oxide is formed first when the metal is heated. However, calcium reacts with chlorine to form calcium chloride; $Ca_{(s)} + Cl_{2(g)} \rightarrow CaCl_{2(s)}$

d) Alkaline earth metals react with dilute acids e.g. hydrochloric acid and sulphuric (VI) acid to form the salts of the metal and hydrogen gas;

- $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
- $Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$
- $Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$
- $Ca_{(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + H_{2(g)}$

{White}

NB: When sulphuric (VI) acid is used, the insoluble calcium sulphate forms a coat on the surface of calcium metal preventing further reaction between the metal and the acid.

Uses of some Alkaline earth metals and their compounds

1. Magnesium is used in the manufacture of magnesium hydroxide which is used as an anti-acid medicine. This is because $Mg(OH)_2$ is a non-toxic base.
2. A low-density alloy of magnesium and aluminium is relatively inert. It is used in aeroplane construction.
3. Calcium is used in the manufacture of hydrated calcium sulphate (plaster of Paris) which is used in hospitals to set fractured bones.
4. Cement is made by heating a mixture of calcium carbonate (limestone), clay and sand.
5. Calcium carbonate is used to remove certain impurities during extraction of iron.
6. Mortar is made by mixing slaked lime (calcium hydroxide) with sand and water.
7. Calcium oxide (quicklime) is added to acidic soils to raise the pH for agricultural purposes.
8. Calcium nitrate is used as a nitrogen fertilizer.
9. Magnesium oxide is used in the lining of furnaces.

c) Halogens

- Halogens are non-metals found in group VII of the periodic table.
- They include;
    - Fluorine [F]
    - Chlorine [Cl]
    - Bromine [Br]
    - Iodine [I]
- The electron arrangement of the first two element are as follows;
    - $_{9}F=2.7$
    - $_{17}Cl=2.8.7$
- Each halogen has seven electrons in the outermost energy level.

Trends in properties of Halogens.

1. Physical properties

a) Fluorine and chlorine are gases at room temperature.

b) Fluorine is pale-yellow while chlorine is green-yellow; bromine is brown liquid while iodine is shiny dark-grey solid.

c) Chlorine and fluorine are soluble in water and Tetrachloromethane. The solubility in water of halogens decreases down the group but all halogens are soluble in Tetrachloromethane.

d) Halogens are non-conductors of electricity.

- The table below gives a summary of the physical properties of some halogens.

Halogens	**Formula**	**Atomic no.**	**Atomic radius (nm)**	**Ionic radius (nm)**	**Appearance**	**Melting point (^0C)**	**Boiling point (^0C)**
Fluorine	F_2	9	0.064	0.136	Pale-yellow gas	-238	-188
Chlorine	Cl_2	17	0.099	0.181	Greenish-	-101	-35

					yellow gas		
Bromine	Br_2	35	0.114	0.195	Brown liquid	-7	59
Iodine	I_2	53	0.133	0.216	Shiny dark-grey solid	144	184

- Further physical properties can be deduced from the table above;

e) Atomic and ionic radius

- These radius increase down the group; ***This is because of the increase in the number of energy levels down the group.***

f) Melting and boiling points- increases down the group; ***This is because, as one moves down the group, the force of attraction between molecules increase with increase in size of the molecule.***

g) Comparison between the atomic and ionic radius of the same element

- The atomic radius of a halogen atom is less than the radius of its ion e.g. the atomic radius of chlorine atom is less than the ionic radius of a chloride atom; ***This is because the chlorine atom has 17 protons in the nucleus attracting 17 electrons in the energy levels while a chloride ion has 17 protons in the nucleus attracting 18 electrons in the energy levels. The effective pull of electrons in the energy levels by the positive nucleus is thus reduced. The electrons therefore tend to move away from the nucleus hence making the size of the ion larger than the atom.***

2. Chemical properties of Halogens

i. Most halogens react with metals to form the corresponding metal-halogen compounds e.g. with Iron;

$$2Fe_{(s)}+3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$$

$$2Fe_{(s)}+3Br_{2(g)} \rightarrow 2FeBr_{3(s)}$$

$$Fe_{(s)}+I_{2(g)} \rightarrow FeI_{2(s)}$$

With zinc;

$$Zn_{(s)}+Cl_{2(g)} \rightarrow ZnCl_{2(s)}$$

$$Zn_{(s)}+Br_{2(g)} \rightarrow ZnBr_{2(s)}$$

$$Zn_{(s)}+I_{2(g)} \rightarrow ZnI_{2(s)}$$

- Other salts formed in the same method are;
  - ✓ $MgCl_2$
  - ✓ $AlCl_3$
  - ✓ NaCl

NB: The reactivity with metals decrease down the group; **The ability of an atom to gain an extra electron in its outermost energy level decrease as the size of the atoms increase down the group**.

ii. Chlorine dissolves in water to form chlorine water which is a mixture of hydrochloric acid and chlorous (I) acid i.e.

$$Cl_{2(g)} + H_2O_{(l)} \rightarrow HCl_{(aq)} + HOCl_{(aq)}$$

On exposing to sunlight, chlorine water decomposes to oxygen and hydrochloric acid;

$$2HOCl_{(aq)} \rightarrow 2HCl_{(aq)} + O_{2(g)}$$

{Green-yellow} {Colourless}

Some uses of halogen and their compounds

1. Fluorine is used in the preparation of a synthetic fibre known as polytetraflouroethene.
2. Some compounds of fluorine are added to water and some tooth pastes in small quantities to reduce tooth decay.
3. Fluorine is used to manufacture hydrogen fluoride used to engrave words or picture on glass.
4. Chlorine is used to make bleaches for use in paper pulp and textile industries.
5. Chlorine is added to water to kill bacteria in water treatment works.
6. Chlorine is used in the manufacture of a type of plastic known as polyvinylchloride (PVC).
7. Chlorine is used in the large-scale manufacture of hydrochloric acid.
8. Bromine is used in the manufacture of silver bromide which is used to make light sensitive photographic paper films.
9. A solution of iodine in alcohol (tincture of iodine) is used as a medicine to kill bacteria in cuts on broken skin.

d) Noble gases

- These are elements in group VIII of the periodic table.
- They include;
  - Helium [He]
  - Neon [Ne]
  - Argon [Ar]
  - Krypton [Kr]
  - Xenon [Xe]
  - Randon [Rn]
- The electron arrangement of the first three noble gases are shown in the table below

Elements	**Symbol**	**Atomic number**	**Electronic arrangement**
Helium	He	2	2
Neon	Ne	10	2.8
Argon	Ar	18	2.8.8

- It can be seen from the table that only Helium has 2 electrons in the outermost energy level and only one energy level.
- The rest have eight electrons in their outermost energy level and thus they have the octet.
- Under normal conditions noble gases neither gains nor loses electrons and thus they are therefore stable and not reactive.
- The table below is a summary of some of the physical properties of noble gases.

Elements	**Symbol**	**Atomic no.**	**Atomic radius (nm)**	**1st I.E (Kj.mol^{-1})**	**Melting point (^{0}C)**	**Boiling point (^{0}C)**

Helium	He	2	0.128	2372	-270	-269
Neon	Ne	10	0.160	2080	-249	-246
Argon	Ar	18	0.192	1520	-189	-186
Krypton	Kr	36	0.197	1350	-157	-152
Xenon	Xe	54	0.217	1170	-112	-108

- The following physical properties can be seen from the table;

  a) The 1st I.E is very high as compared to the other chemical families; **This is because the outermost energy level is completely filled with electrons and hence their stability is very strong.**

  b) The atomic radius increases down the group due to the increase in the number of energy levels; **This explains why the first ionization energy of the gases decrease down the group.**

  c) Noble gases have low melting and boiling points; This is because of the weak intermolecular forces of attraction between the atoms. As one move down the group, this force increase in strength and hence the increase in the melting and boiling points.

Uses of noble gases

1. Argon is used in light bulbs to provide an inert environment to prevent oxidation.
2. Argon is used as an insulator in arch-welding.
3. Neon gas is used in street and advertising lights.
4. Helium mixed with oxygen is used in deep sea diving and mountaineering. The mixture is also used in hospitals for patients with respiratory problem and those undergoing certain forms of surgery.
5. Helium can be used instead of hydrogen in balloons for meteorological research.
6. Helium is used in thermometer for the measurement of very low temperatures.
7. Liquid helium is used to keep certain metals alloys at temperature low enough for then become super conductors.

Properties and trends across a period

- Those elements across a period show a gradual change in properties.

- These changes are illustrated by considering period 3.

Trends in physical properties of elements in period 3

1. Sodium, magnesium and aluminium are good conductors of electricity, silicon is a semi-conductor while phosphorous, Sulphur, chlorine and argon are poor conductors; **This is because sodium, magnesium, and aluminium have delocalized electrons in their structure. Phosphorous, Sulphur, chlorine and argon are all made-up of molecules and therefore are non-conductors of electricity**.

2. The atomic radii of the elements gradually decrease across the period from left to right; **This is because the nuclear charge increase as one moves from left to right due to the increase in the number of protons. Although there are additions of electrons, they enter in the same energy level. As a result, the electrons in the outermost energy level are pulled closer to the nucleus, thereby decreasing the size of the atoms across the period from sodium to chlorine.**

    - The following table gives a summary of some physical properties of the elements in period 3.

Elements/ properties	**Na**	**Mg**	**Al**	**Si**	**P (white)**	**S (Monoclinic)**	**Cl**	**Ar**
Physical state and appearance	Silvery solid	Silvery solid	Silvery solid	Black solid	White solid	Yellow solid	Greenish-yellow gas	Colourless gas
Electron arrangement	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 or 5	2	1	0
Atomic radius (nm)	0.157	0.136	0.125	0.117	0.110	0.104	0.09	0.192
Melting point (^{0}C)	98	650	660	1410	44	119	-110	-189
Boiling point (^{0}C)	890	1110	2470	2360	280	445	-35	-186
Structure	Giant metallic	Giant metallic	Giant metallic	Giant atomic	Molecular	Molecular	Molecular	Exist as an atom
Bond type	Metallic	Metallic	Metallic	Covalent	Covalent	Covalent	Covalent	Van der Waal

- The following can be seen from the table above;

3. Sodium, magnesium and aluminium have giant metallic structures and therefore have strong metallic bonds. These bonds require a lot of energy to break hence they have high melting and boiling points.
4. Silicon has a giant atomic structure in which all the atoms are held together by strong covalent bonds. These need a lot of heat energy to break, hence the high melting and boiling points of silicon.
5. Phosphorous, Sulphur, and chlorine are molecular. The atoms of the molecules are held by strong covalent bonds while the molecules themselves are held together by van der Waal forces which require little energy to break and hence their melting and boiling points are low.
6. Chlorine and Argon exist as gases at room temperature. They have low melting and boiling points due to the presence of weak van der Waal forces.

Trends in chemical properties of the elements of period 3

1. Sodium reacts vigorously with oxygen to form a white solid;

$$4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$$

-Magnesium burns with a bright white light to give a white solid,

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

-Aluminium foil is usually coated with a layer of aluminium oxide. This oxide prevents further reaction with oxygen. When polished, Aluminium reacts slowly with oxygen to form a white solid;

$$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$

NB: Na_2O and MgO are soluble in water to form an alkaline solution, while Al_2O_3 is insoluble.

-Silicon powder can only burn in oxygen at high temperature (about 450^0C) to form solid Silicon (IV) oxide.

$$Si_{(s)} + O_{2(g)} \rightarrow SiO_{2(s)}$$

NB: SiO_2 is insoluble in water.

-Phosphorous readily burns in oxygen with a bright orange flame to form a white solid, phosphorous (V) oxide.

$P_{4(s)}+5O_{2(g)} \rightarrow 2P_2O_{5(s)}$

-Sulphur burns in oxygen with a blue flame to form a gas,SO_2

$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$

NB: P_2O_5 and SO_2 dissolves in water to form acidic solutions.

-Chlorine burns in oxygen under certain conditions to form acidic oxides while Argon is not reactive.

- The following can be deduced from the above discussions;

i. All the elements across period 3, with the exception of Argon, burn in oxygen to form oxides.

ii. The reactivity of the metals with oxygen decrease from left to right across the period; **This is because of the increase in nuclear charge from sodium to aluminium, which makes it easier to remove an electron from a sodium atom than from an aluminium atom.**

iii. Metallic elements burn in oxygen to form basic oxides. Soluble metallic oxides dissolve in water to form alkaline solutions.

iv. The reactivity of the non-metallic elements with oxygen increase from left to right across the third period; **This is because the ease of gaining electrons increases from left to right.**

v. The non-metallic elements burn in oxygen to form acidic oxides which dissolve in water to form acidic solution.

2. -Sodium reacts violently with cold water to form sodium hydroxide and hydrogen gas

$2Na_{(s)}+H_2O_{(l)} \rightarrow 2NaOH_{(aq)}+H_{2(g)}$

-Magnesium reacts very slowly with cold water to form magnesium hydroxide and hydrogen gas

$Mg_{(s)} +2H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)} + H_{2(g)}$

-Aluminium does not normally react with cold water or steam due to the presence of a coating of aluminium oxide, which prevents any reaction.

NB: In general, the reactivity of metals with water decreases from sodium to aluminium. Sodium is more reactive because it loses its valence electron more readily than magnesium and aluminium.

-Non-metals do not displace hydrogen from water. However, Chlorine is an exceptional since it dissolves in water to form chlorine water which is a mixture of hydrochloric acid and chloric (I) acid.

$Cl_{2(g)}+H_2O_{(l)} \rightarrow HCl_{(aq)} + HOCl_{(aq)}$

- The following table summarizes the reaction of elements of period 3 with air, water and dilute acids.

Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
Reaction with air or oxygen	Readily reacts with air. Burns brightly in oxygen to form Na_2O.	React slowly with air. Burns in oxygen with a bright flame to form MgO.	Forms a protective layer of Al_2O_3 when it burns in oxygen.	Silicon powder burns at a temperature above 950^0C to form SiO_2	White phosphorous Smoulders in air. Burns in air to form P_2O_5 and P_2O_3	Burns in air or oxygen to form SO_2	No reaction with air or oxygen under normal conditions.
Reaction with water	Vigorous reaction producing H_2 and NaOH	Slow reaction with cold water but reacts with steam rapidly.	No reaction	No reaction	No reaction	No reaction	Dissolves in water to form chlorine water.
Reaction with dilute acids	Violent reaction giving H_2 and salt.	Rapid evolution of H_2 and salt formed.	Reacts slowly to form H_2	No reaction	No reaction	No reaction	No reaction

3. It can be seen from the table;

i. Magnesium reacts with both dilute hydrochloric and sulphuric (VI) acid to from a salt and hydrogen gas

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(s)} + H_{2(g)}$$

$$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$$

ii. Aluminium does not readily react with dilute acid due to the presence of Al_2O_3 coating. However, on removing the oxide coating, the metal reacts.

$$2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2AlCl_{3(aq)}+H_{2(g)}$$

$$2Al_{(s)} + 3H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 3H_{2(g)}$$

iii. The reaction between sodium and dilute acids is explosive and therefore should never be tried.

iv. Silicon, phosphorous, Sulphur and chlorine do not react with dilute acids

STRUCTURE AND BONDING

Specific objectives

By the end of this topic, the learners should be able to:

a) State the significance of valence electrons in bonding.

b) Explain qualitatively the formation of ionic, covalent and metallic bonds.

c) Diagrammatically illustrate ionic, covalent, dative, hydrogen bonds and Van der Waal forces.

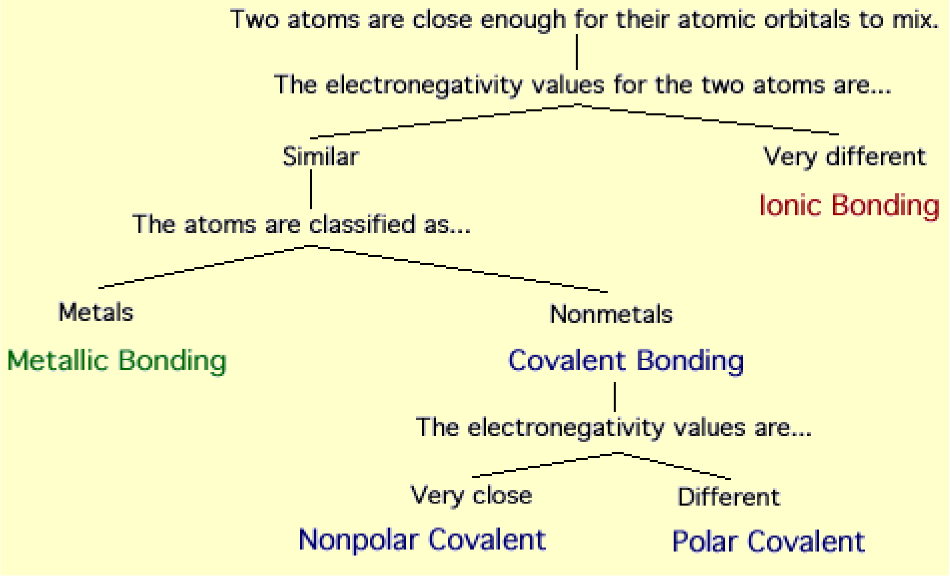
d) Predict the bond type and structure of a given substance from its physical properties.

e) Explain the changes in bond type across period 3.

f) Select materials for use based on bond types and structure.

g) Predict the properties of a given substance on the basis of the bonds present.

Bond types



- There are many types of chemical bonds between atom in a given structure, of most importance are;

(i) Ionic bond

(ii) Covalent bond

(iii) Dative bond

(iv) Hydrogen bond

(v) Metallic bond

(i) Ionic bond

- It involves total/complete transfer of electrons from one atom to another.
- It involves the reaction of metals with non-metals e.g.
  - The reaction between sodium and chlorine to form sodium chloride.

$$\underset{2.8.1}{Na} + \underset{2.8.7}{Cl} \rightarrow Na^+Cl^-$$

- An ionic bond is therefore formed when there is complete transfer of valence electron from one atom to another resulting in two ions with opposite charges.
  - Other examples of ionic compounds include KF, MgO and $MgCl_2$

[Diagrams of these examples]

Exercise

1. Draw the ionic bond formation of magnesium oxide [Atomic number of O=8; and Mg=12]

NB: This type of bonded compound have only one structure which is Giant ionic structure.

Giant Ionic structure

- Sodium chloride structure consists as many sodium ions and chloride ions that are arranged and packed in a regular pattern.
- Each sodium ion is surrounded by six chloride ions that are equidistant from it and the same is true for each chloride ion. [Diagram of NaCl]

- The pattern repeats itself many times in all directions resulting in the formation of a giant ionic structure.
- Other examples of ionic substances with giant ionic structure include; KNO_3, NaI, KBr and $Ca(NO_3)_2$

Physical properties of ionic compounds

- The table below summarizes some physical properties of ionic compounds.

Properties \ **Compounds**		**NaF**	**NaCl**	**NaBr**	**NaI**
Solubility		Soluble	Soluble	Soluble	Soluble
Mpt(^{0}C)		993	801	747	661
Bpt (^{0}C)		1695	1413	1390	1304
Electrical conductivity	**Solid**	Does not	Does not	Does not	Does not
	Melt or solution	Conducts	conducts	Conducts	Conducts

- The following physical properties can be seen from the table

  a) Ionic compounds dissolve in polar solvents like water.

  b) Ionic compounds conduct electricity only in molten or solution state**; this is because in solid state the ions are rigid/fixed while in molten or solution state they are free to move within the molten liquid or solution and therefore conducting electricity.**

(ii) Covalent bond

- When atoms combine by sharing electrons and their respective nuclei attract the shared electrons, this constitutes the covalent bond holding the two atoms together.
- It is mostly formed by atoms of the same element or a group of atoms of different elements combining to form a molecule.
- Substances consisting of molecules include water, most gases, sugar oils, fats, naphthalene, kerosene, wax and Sulphur and are called molecular substances.
- The covalent bond can be represented diagrammatically as follows; {in each case only Valency electrons are shown}

[Diagram of water molecule, Chlorine molecule]

Exercise

Using a dot (.) and cross diagram, draw the chemical bond for the following molecule;

(i)O_2 (ii)N_2(iii)NH_3(iv)CO_2

(iii) Co-ordinate/ Dative bond

- This a type of covalent bond in which the shared pair of electrons forming the bond is contributed by only one of the atoms forming the bond.
- Examples of molecules that have co-ordinate bonds are NH_4^+ ion and CO.

[Diagram showing both molecules]

(iv) Hydrogen bond

- Hydrogen bond is an attraction between slightly positive hydrogen on one molecule and a slightly negative atom on another molecule.
- In this bond, the pair of electrons that is shared comes from both atoms.
- One atom attracts (pulls) the pair of electrons towards itself more than the other, it becomes partially negative (δ-) leaving the other atom partially positive (δ+).
- When this occurs, the molecule is said to be polar.
- Example of polar molecules includes water.
- The attraction between the partially negative atom and the partially positive hydrogen constitute the hydrogen bond and is represented by a dotted line as shown below.

[Diagram of hydrogen bond in water]

NB: Substances which have covalent bonds have either molecular or giant covalent (Atomic) structures.

Molecular structure

- Generally, molecular substances are gases or liquids at room temperatures.
- Molecular substances include Sulphur, sugar, ice, dry ice, iodine, fats, paraffin wax and naphthalene.

- The atoms forming the molecules are held together by strong covalent bonds.
- The molecules in the solid substances are in turn held together in a regular pattern by weak Van der Waal forces in a pattern that repeats itself many times regularly resulting to a molecular structure as shown below;

  [Diagram of the structure of iodine molecule]

Physical properties of molecular substances

- The table below summarizes some physical properties of molecular structures.

SUBSTANCES **PROPERTIES**	**Sugar (Sucrose)**	**Naphthalene**	**Iodine**	**Rhombic Sulphur**
Solubility in water	Soluble	Insoluble	Insoluble	Insoluble
Molecular mass	183	128	186	256
Mpt(^{0}C)	200	80.2	114	113
Bpt (^{0}C)	-	218	183	444
Electrical conductivity (in solid or melt)	Poor	Poor	Poor	Poor

- The following properties can be deduced from the table;

  (i) Molecular substances are poor conductors of both heat and electricity; **this is because they have structures composed of molecules and therefore do not have electric carriers like ions or electrons.**

  (ii) Molecular substances have low melting and boiling points; **although the atoms forming a molecule are held by strong covalent bonds, the intermolecular forces are usually the weak Van der Waal forces or hydrogen bond.**

  (iii) Melting and boiling points of molecular substances increase with increase in relative molecular mass.

  (iv) Most molecular substances are insoluble in water but sugar is soluble due to the presence of hydrogen bonding in it.

Giant Atomic structure

- This structure consists of an indefinite number of atoms which are all covalently bonded together.
- The pattern repeats itself and extends in all directions.
- A good example of such substances include; Diamond, Graphite and Silicon (IV) oxide.

Diamond

- This is an allotrope of carbon
- In its structure, each carbon atom is bonded to other carbon atoms by strong covalent bonds.
- The carbon atoms in diamond form an octahedral structure as shown below;

[Diagram pg77, klb bk 2]

- The pattern repeats itself in all directions resulting to a giant atomic structure.
- In the process above all the valence electrons of each carbon atom are used in bonding and therefore no delocalized electrons are found in the structure.

Graphite

- Each carbon atom in graphite is bonded to three other carbon atoms.
- Because only three of the four valency electrons are used in bonding, the forth valency electron is delocalized in the structure.
- The structure of graphite consists of layers in which the carbon atoms are held by strong covalent bonds forming hexagonal rings as shown below;

[Diagram pg 78, klb bk 2]

- Each layer is held to the other by weak Van der Waal forces.

(v) The metallic bond

- In this bond there are many atoms surrounding any one atom and therefore the valency electrons are mutually attracted to many nuclei.

[Diagram pg 79, klb bk 2]

- The electrostatic forces of attraction between the positively charged nuclei and the negatively charged delocalized electrons hold the atoms together in the metallic structure.

- The mutual electrostatic forces constitute metallic bonds which are strong bonds.
- The following table summarizes the physical properties of metals which have metallic bonds;

Metals	**Valency**	**Mpt(^{0}C)**	**Bpt(^{0}C)**	**Atomic radii (nm)**	**Electrical conductivity**
Lithium	1	180	1330	0.133	Good
Sodium	1	98	890	0.155	Good
Potassium	1	64	774	0.203	Good
Magnesium	2	651	1110	0.136	Good
Aluminium	3	1083	2582	-	Good

- The following properties can be deduced;-

1. All metals are good conductors of both heat and electricity**; this is because there are delocalized electrons in the metallic structure.**
2. Thermal and electrical conductivity increases with the increases number of delocalized electrons from each atom in the structure; **this explains why aluminium is the best conductor.**
3. Metals have relatively high melting and boiling points; **this is because the metallic bond is a very strong bond.**

Types of bonds across a period

- The type of bond formed depends on the electron in the outermost energy level.

Oxide of elements in period 3

- The table below summarizes the bond type and properties of oxide of elements in period 3.

Oxide	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_2O_5**	**SO_2**	**Cl_2O_5**
Physical state	Solid	Solid	Solid	Solid	Solid	Gas	Gas
Mpt (^{0}C)	1193	3075	2045	1728	563	-76	-60
Bpt(^{0}C)	1278	3601	2980	2231	301	-10	-9
Structur	Giant	Giant	Giant ionic	Giant	Molecula	Molecula	Molecula

e	ionic	ionic		Atomic	r	r	r
Type of bond	Ionic	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
Nature of oxide	Basic (Alkaline)	Basic (Alkaline)	Amphoteric	Acidic	Acidic	Acidic	Acidic
Reaction with acids	React to form salt and water	React to form salt and water	React to form salt and water	No reaction	No reaction	No reaction	No reaction

- The following can be seen from the table;

  (i) All the oxides of elements in period 3 except those of Sulphur and chlorine are solid at room temperature.

  (ii) Both sodium oxide and magnesium oxide have a giant ionic structure, however, the melting point of Magnesium oxide is higher because the electrostatic forces of attraction between magnesium ions and oxide ions are stronger since it has a charge of +2 and is smaller than sodium ion.

  (iii) Silicon (IV) oxide has a giant atomic structure. Each silicon atom is attracted to two silicon atoms.

- In general therefore;

  (a) The bond type change from ionic to covalent across the period

  (b) The structure of the oxide changes from giant ionic to giant atomic and finally to molecular which explains the trend in properties of oxides.

Chlorides of elements in period 3

- Most elements of period 3 form stable chlorides.
- The trends in bonding type and properties of chlorides of elements of period 3 is as shown below;

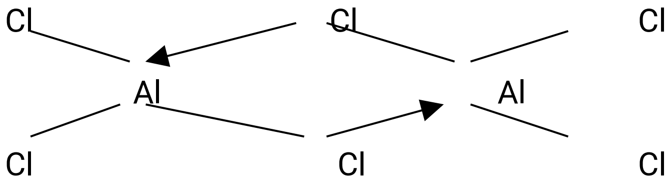
Chlorides	**$NaCl$**	**$MgCl_2$**	**$AlCl_3$**	**$SiCl_4$**	**PCl_5**	**SCl_2**

Physical state	Solid	Solid	Solid	Liquid	Liquid	Liquid
Mpt (^{0}C)	1074	987	451	203	435	195
Bpt(^{0}C)	1686	1685	-	331	-	332
Structure	Giant ionic	Giant ionic	Molecular (Dimer)	Molecular	Molecular	Molecular
Bond type	Ionic	Ionic	Ionic/ Covalent	Covalent	Covalent	Covalent
pH of solution	7	7	3	3	3	3
Effect on water	Soluble	Soluble	Hydolysed	Hydolysed	Hydolysed	Hydolysed
Temperature **Change during dissolving**	Slightly drop	Small change	High	High	Very high	High

- The properties from the table are as follows;

  (i) NaCl and $MgCl_2$ register small changes in temperature when dissolved in water to form neutral solutions. They are also ionic and therefore expected to dissociate fully into ions.

  (ii) Anhydrous $AlCl_3$ exist in molecular form and it differs from (Al_2Cl_6). The $AlCl_3$ molecular pair through co-ordinate (dative) bond as shown below



- $AlCl_3$ is hydrolysed by water to form an acidic solution and therefore behave as a covalent rather than an acidic chloride.

  (iii) All the other chloride hydrolyses in water i.e. breaking down some substance in water e.g.

$$SiCl_{4(l)} + 2H_2O_{(l)} \rightarrow SiO_{2(s)} + 4HCl_{(aq)}$$

- The following are characteristics of bonds summarized.

Property	Substances with		
	Covalent bonds	**Ionic bonds**	**Metallic bonds**
Electrical conductivity	Non-conductors, except graphite	Solids do not conduct. Aqueous solutions and molten state conduct	Conduct
Thermal conductivity	Non-conductors, except graphite	Do not conduct	conduct
Melting point (^0C)	Low for molecular High for giant structures	Usually high	Generally high
Boiling point (^0C)	Low if molecular High if giant structure	Usually high	Generally high
Solubility	Generally insoluble in water but soluble in organic solvents	Generally soluble in water	Some metals react with water

SALTS

Specific objectives

By the end of the lesson the learners should be able to:

(a) Select and use appropriate methods of preparing particular salt.

(b) Explain the terms saturated solution, crystallization, neutralization and precipitation.

(c) Write ionic equations for the preparation of salts.

(d) State types of salts.

(e) Identify soluble and insoluble salts.

(f) Describe and explain from experimental observations the action of heat on various salts.

(g) State use of some salts.

SALT.

- A salt is a compound formed when metal atoms or ammonium ions combine with non-metallic atoms or radicals except oxides and hydrides.
- There are four types of salts namely;

(1) Normal salts

(2) Acid salts

(3) Basic salts

(4) Double salts

Normal salt

These are salts which do not contain a hydrogen atom e.g. NaCl, K_2SO_4, Na_2CO_3 and Ca $(NO_3)_2$.

Acid salts

These are salts that contain a replaceable hydrogen atom e.g. $NaHCO_3$, $KHSO_4$, NaH_2PO_4 (sodium dihydrogen phosphate).

NB: The acidic property is due to the presence of the hydrogen.

Basic salts

These are salts that contain hydroxyl (OH^-) ions e.g. $Mg(OH)Cl${basic magnesium chloride},$Pb(OH)_2.2PbCO_3${basic lead(II) carbonate},$Zn(OH)Cl${basic zinc chloride} and $(CuCO_3).Cu(OH)_2${basic copper(II) carbonate.

NB: The presence of the hydroxyl ion makes the salts basic.

Double salts

These are salts in which there are two different anions or cations e.g. hydrated aluminium sulphate {$KAl(SO_4)_2.12H_2O$} ,hydrated diammonium iron(II) sulphate{$Fe(NH_4)_2(SO_4)_2.6H_2O$}.

NB: They are normally crystalline solids with specific shapes.

METHODS OF PREPARING SALTS

There are several general methods of preparing salts, the choice of which method to use largely depends on whether the salt is soluble or not in water.

Those salts that dissolve in water are termed as *soluble* salts while those that do not are called *insoluble* salts.

Solubility of salts in water

The solubility of salts can be summarized as in the table below;

Carbonates	All are insoluble except Na, K and NH_4^+ carbonates
Chlorides	All are soluble except Ag and Pb chlorides with lead chloride being soluble in hot water.
Sulphates	All are soluble except Ba and Pb sulphates while calcium sulphate is slightly soluble.
Nitrates	All are soluble.

NB: a) All compounds of NH_4^+, Na and K are soluble.

b) The carbonates of Al and Fe do not exist.

Solubility of bases in water.

The solubility of bases in water and the effect of the resultant solution can be summed up as per the table below.

CATIONS	**SOLUBILITY**		**EFFECTS OF RESULTING SOLUTION ON RED LITMUS PAPER**	
	Oxide	**Hydroxide**	**Oxide**	**Hydroxide**
K	Soluble	Soluble	Turns blue	Turns blue
Na	Soluble	Soluble	Turns blue	Turns blue
Ca	Slightly soluble	Slightly soluble	Turns blue	Turns blue
Mg	Slightly soluble	Slightly soluble	Turns blue	Turns blue
Zn	Insoluble	Insoluble	No effect	No effect
Al	Insoluble	Insoluble	No effect	No effect
Cu	Insoluble	Insoluble	No effect	No effect

NB: Soluble oxides and hydroxides turns red litmus paper blue while those that are insoluble have no effect on the litmus paper.

Terms used in this topic

- SATURATED SOLUTION: **is a solution that cannot dissolve any more of the solute at given temperature and contains undissolved solute.**
- CRYSTALLIZATION:
- NEUTRALIZATION:
- PRECIPITATION:

The methods of preparation of salts include:

(1) Action of an acid on a metal

(2) Action of an acid on an insoluble base(metal oxide or hydroxide)

(3) Action of an acid on an insoluble or soluble carbonate and hydrogencarbonate.

(4) Action of an acid on an alkali

(5) Direct combination of elements/direct synthesis.

(6) Preparation of insoluble salts by precipitation method.

1) Action of an acid on a metal

Exp: Preparation of a soluble salt

Aim

- To prepare $ZnSO_4$ as an example of a soluble salt.

Apparatus and chemicals

1) Beakers ($100cm^3$)
2) Filter funnel
3) Filter papers
4) Evaporating dish
5) Conical flasks
6) Bunsen burner
7) Dilute H_2SO_4 (2M)
8) Zinc granules

Procedure

1) Pour $50cm^3$ of dilute H_2SO_4 into a beaker.
2) Add zinc granules. What do you observe**? bubble of a gas produced**
3) If the reaction is too slow, add a little $CuSO_4$ solution and warm gently, but do not boil. What do you observe? **the reaction become fast**
4) When all zinc has reacted, add more until it has all reacted. Add more again until some remain unreacted. Give a reason why we must keep on adding zinc until no more can react. **To ensure that all the acid had reacted**
5) Filter off the excess zinc and solid impurities and collect the filtrate in a conical flask. What is the colour of the filtrate? **colourless**
6) Evaporate the filtrate to dryness. What do you observe? **White salt crystals are formed**
7) Record your observations in your books.

Conclusion

The soluble salt $ZnSO_4$ was effectively obtained in this experiment.

NB: Mg powder could be used in this experiment instead of Zn granules.

The equation for the above process is;

$Zn_{(s)} + H_2SO_{4\,(aq)} \longrightarrow ZnSO_{4\,(aq)} + H_{2\,(g)}$

Other examples include:

$Mg_{(s)} + H_2SO_{4\,(aq)} \longrightarrow MgSO_{4\,(aq)} + H_{2\,(g)}$

$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}$

$2Al_{(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3\,(aq)} + H_{2\,(g)}$

$2Fe_{(s)} + 6HCl_{(aq)} \longrightarrow 2FeCl_{3\,(aq)} + 3H_{2\,(g)}$

The general equation can therefore be written as follows:

Metal + dilute acid ⟶ soluble salt + hydrogen gas.

2) Reaction of a dilute acid with an insoluble metal oxide or hydroxide

An acid is reacted with an insoluble oxide or hydroxide to give soluble salt.

$CuO_{(s)} + H_2SO_{4\,(aq)} \longrightarrow CuSO_{4\,(aq)} + H_2O_{(l)}$

$Mg(OH)_{2(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_2O_{(l)}$

$PbO_{(s)} + 2HNO_{3(aq)} \longrightarrow Pb(NO_3)_{2(aq)} + H_2O_{(l)}$

$Zn(OH)_{2(s)} + H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + 2H_2O_{(l)}$

3) Action of acids on insoluble carbonates

In this method, an acid is reacted with an insoluble metallic carbonate e.g.

$CaCO_{3(s)} + 2HNO_{3(aq)} \longrightarrow Ca(NO_3)_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

$CuCO_{3(s)} + H_2SO_{4(aq)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

$CuCO_{3(s)} + 2HNO_{3(aq)} \longrightarrow Cu(NO_3)_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

$ZnCO_{3(s)} + H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

$(NH_4)_2CO_{3(aq)} + H_2SO_{4(aq)} \longrightarrow (NH_4)_2SO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

Therefore the general equation can be written as follows;

Carbonate + Acid ⟶ Salt + Carbon (IV) oxide + water.

4) Action of acids on alkalis

This reaction is known as the neutralization reaction.

The method is used for the preparation of very soluble salts which do not form crystals easily i.e. mainly salts of sodium, potassium and ammonium e.g.

$NH_4OH_{(aq)} + HCl_{(aq)} \longrightarrow NH_4Cl_{(aq)} + H_2O_{(l)}$

$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

$2KOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$

Therefore the general equation can be written as follows;

Alkali + Acid $\longrightarrow$ Salt + water

5) Preparation of salts by direct combination of elements (direct synthesis)

Used for the preparation of both soluble and insoluble salts.

Exp: How can iron (II) Sulphide be prepared?

Procedure:

1. To a spatulaful of iron filling in a crucible, add a spatulaful of sulphur and mix them well.
2. Heat the mixture strongly until the reaction is complete.
3. Allow the products to cool.

Observations:

- The colour of iron (grey) and that of sulphur (yellow) changed to black (iron (II) Sulphide).

The equation involved is given as;

$Fe_{(s)} + S_{(s)} \longrightarrow FeS_{(s)}$

Other salts that can be prepared by this method are; NaCl, $FeCl_3$

6) Preparation of insoluble salts by precipitation method.

Exp: How is lead(II) sulphate prepared?

Aim

To prepare an insoluble salt e.g. $PbSO_4$

Apparatus and reagents.

- Beaker
- Filter papers
- Stirring rod
- Test tube
- Distilled water
- Lead(II) nitrate
- Magnesium sulphate

Procedure.

1. To 10cm^3 of $Pb(NO_3)_2$in a beaker, add excess $MgSO_4$ solution {about 14cm^3} and mix.
2. Let the solid to settle, and then pour off the liquid.
3. Wash the solid with distilled water.
4. Filter and dry the solids between filter papers.

When the two soluble salts are mixed, they react to form lead (II) sulphate and Magnesium nitrate salts.

In this reaction, the metal ions of the salts exchange ions of their salts. This reaction is called double decomposition.

$$Pb(NO_3)_{(aq)} + MgSO_{4(aq)} \longrightarrow PbSO_{4(s)} + Mg(NO_3)_{(aq)}$$

$$Pb^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \longrightarrow PbSO_{4(aq)}$$

Stoichiometric equation; These are chemical equations showing all the reactants and the product.

Ionic equations; This is an equation obtained by writing only those ions that undergo change during the chemical reaction.

Spectator ions; These are ions that remain unchanged during a chemical reaction e.g. $Mg^{2+}_{(aq)}$ and $NO_3^-{}_{(aq)}$ in the above equation. They are usually omitted when writing ionic equations.

NB: In the precipitation reaction, two reactants must be soluble salts and one of the products is an insoluble salts.

Example; Write the Stoichiometric and ionic equation leading to the formation of lead (II) chloride salt.

Solution

Stoichiometric equation; $Pb(NO_3)_{2(aq)} + MgCl_{2(aq)} \rightarrow Mg(NO_3)_{2(aq)} + PbCl_{2(s)}$

$Pb^{2+}_{(aq)} + \cancel{2NO_3^-}_{(aq)} + \cancel{Mg^{2+}}_{(aq)} + 2Cl^-_{(aq)} \rightarrow \cancel{Mg^{2+}_{(aq)}} + \cancel{2NO_3^-}_{(aq)} + PbCl_{2(s)}$

Ionic equation; $Pb^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow PbCl_{2(s)}$

WHAT HAPPENS WHEN SALTS ARE EXPOSED TO THE ATMOSPHERE

1. Deliquescent salts; Those salts that absorb so much water from the atmosphere that they form solutions e.g. anhydrous calcium chloride, zinc chloride, sodium hydroxide, magnesium chloride, iron (III) chloride and potassium hydroxide. The process is called deliquescence.

2. Hygroscopic salts; When exposed to the atmosphere, they absorb water from the atmosphere, but do not form solutions e.g. anhydrous copper (II) sulphate, potassium nitrate and common table salt (pure sodium chloride is not hygroscopic, but common table salt is because it contain other salts such as magnesium chloride which makes it behave like a hygroscopic salt).The process is called hygroscopy.

3. Efflorescent salts; These salts lose some of their water of crystallization e.g. crystals of sodium carbonate decahydrate ($Na_2CO_3.10H_2O$),Iron(II) sulphate heptahydrate ($FeSO_4.7H_2O$) and sodium decahydrate ($Na_2SO_4.10H_2O$).This process is called efflorescence.

ACTION OF HEAT ON SALTS

CARBONATES.

1. Carbonate of potassium and sodium: Pure carbonate of potassium and sodium are not affected by heat. They are stable to heat but lose their water of crystallization.

   $Na_2CO_3.10H_2O_{(s)} \xrightarrow{heat} Na_2CO_{3(s)} + 10H_2O_{(l)}$

2. Other metal carbonates: They give carbon (IV) oxide and the corresponding metal oxide. Metal carbonate $\rightarrow$ Metal oxide + Carbon (IV) oxide. e.g.

   $CaCO_{3(s)} \xrightarrow{heat} CaO_{(s)} + CO_{2(g)}$

   (White) (white)

   $ZnCO_{3(s)} \xrightarrow{heat} ZnO_{(s)} + CO_{2(g)}$

   (Yellow-hot;

White-cold)

$PbCO_{3(s)} \xrightarrow{heat} PbO_{(s)} + CO_{2(g)}$

(Red/Orange-hot;

Yellow-cold)

$CuCO_{3(s)} \xrightarrow{heat} CuO_{(s)} + CO_{2(g)}$

(Black)

<u>NB:</u> a) Carbon (IV) oxide forms a white suspension (cloudy) with calcium hydroxide, lime water.

b) The lower the metal in the reactivity series, the more easily its carbonate is decomposed.

3. Decomposition of ammonium carbonate; Decomposes to form ammonia gas, carbon (IV) oxide and steam.

$$(NH_4)_2CO_{3(s)} \xrightarrow{heat} 2NH_{3(g)} + CO_{2(g)} + H_2O_{(g)}$$

The presence of ammonia gas is confirmed by the turning of a moist red litmus paper blue.

NITRATES.

1. Decomposition of nitrates of potassium and sodium; They decompose to form oxygen and a white residue. $2KNO_{3(s)} \xrightarrow{heat} 2KNO_{2(s)} + O_{2(g)}$

$$2NaNO_{3(s)} \xrightarrow{heat} 2NaNO_{2(s)} + O_{2(g)}$$

2. Other metal nitrate; Other nitrate decompose to give the oxide,nitrogen(IV) oxide and oxygen e.g. $2Ca(NO_3)_{(s)} \xrightarrow{heat} 2CaO_{(s)} + 4NO_{2(g)} + O_{2(g)}$

   Metal oxide $\xrightarrow{heat}$ metal +Nitrogen (IV) oxide + oxygen

3. Decomposition of nitrate of silver and mercury;They decompose to give nitrogen(IV) oxide, oxygen and the corresponding metal e.g. $2AgNO_{3(s)} \xrightarrow{heat} 2Ag_{(s)} + 2NO_{2(g)} + O_{2(g)}$

Metal nitrate heat Metal + Nitrogen(IV) oxide +oxygen gas

4. <u>Decomposition of ammonium nitrate</u>; Decomposes on heating to give steam and nitrogen(I) oxide i.e. $NH_4NO_{3(s)} \xrightarrow{heat} 2H_2O_{(g)} + N_2O_{(g)}$

SULPHATES

1. <u>Action of heat on sulphate of potassium, sodium, calcium and magnesium</u>; These sulphates are stable and are thus not affected by heat. However, if hydrated, they lose water of crystallization on heating which condenses high up in the boiling tube. e.g.

   $Na_2SO_4.10H_2O_{(s)} \xrightarrow{heat} Na_2SO_{4(s)} + 10H_2O_{(l)}$

2. <u>Action of heat on sulphates of zinc and copper</u>; Decompose on strong heating to form metal oxide, water and sulphur (VI) oxide.

   $ZnSO_4.7H_2O_{(s)} \xrightarrow{heat} ZnO_{(s)} + SO_{3(g)} + 7H_2O_{(l)}$

   (white) (Yellow-hot;

   white-cold)

   $CuSO_4.5H_2O_{(s)} \xrightarrow{heat} CuO_{(s)} + SO_{3(g)} + 5H_2O_{(l)}$

   (blue) (black)

3. <u>Action of heat on iron(II) sulphate crystals;</u> Decompose to give Iron(II) oxide,Sulphur(IV) oxide,Sulphur(VI) oxide and water.

   $2FeSO_4.7H_2O_{(s)} \xrightarrow{heat} Fe_2O_{3(s)} + SO_{2(g)} + 7H_2O_{(l)} + SO_{3(g)}$

   <u>NB:</u> All ammonium salts decompose on heating.

Uses of salts

1. Sodium nitrate and potassium are used as fertilizers.
2. Sodium chloride is used as food addiditive.
3. Sodium hydrogen carbonate is used in making baking powder.
4. Calcium sulphate(plaster of paris) is used in hospitals as plaster on people with flactures or dislocation.
5. Calcium chloride is used in extraction of sodium metal by electrolysis.

6. Potassium nitrate is used in making fireworks and gun powder.
7. Sodium carbonate is used in softening of water,making of glass and detergents.
8. Salts are used for de-ionising snow in cold countries during winter by lowering its freezing point (REM:impurities lower the melting points).
9. Some lead salts are used in the manufacture of lead paints.
10. Milk of magnesia (magnesiun hydroxide) is used to reduce excess acid in the stomach,as it neutralises the excess acid.
11. Lime (calcium hydroxide) is used to neutralize soil having excess acidity.

CARBON AND SOME OF ITS COMPOUNDS

Specific objectives

By the end of the topic, the learners should be able to:

(a) Define allotropy and allotrope.

(b) Explain the physical properties of the carbon allotropes in terms of bonding and how the properties are related to the uses of the allotrope.

(c) Describe some chemical properties of carbon.

(d) Describe laboratory preparation and properties of carbon (IV) oxide.

(e) State and explain the physical and chemical properties of carbon (IV) oxide.

(f) Describe laboratory preparation and some properties of carbon (II) oxide.

(g) Describe the chemical reactions of carbonate and hydrogen carbonate.

(h) Describe the manufacture of sodium carbonate.

(i) Explain the advantages and disadvantages of carbon (IV) oxide gas in the atmosphere.

(j) Explain the importance of carbon compounds in the natural environment and industry.

INTRODUCTION.

- Carbon is the first element in group (IV) and period two of the periodic table.
- It has an atomic number of six hence it electron arrangement is 2.4.
- It forms four covalent bonds.
- It occurs in pure state as diamond and graphite and impure state as amorphous carbon e.g. charcoal, soot, coke and coal.

Allotropes of carbon.

- Allotropy; This is the existence of an element in more than one form, in the same physical state.

- The different forms are called allotrope. Carbon has four allotropes, namely;
  - ✓ Diamond
  - ✓ Graphite
  - ✓ Fullerenes
  - ✓ Amorphous carbon

Amorphous carbon;

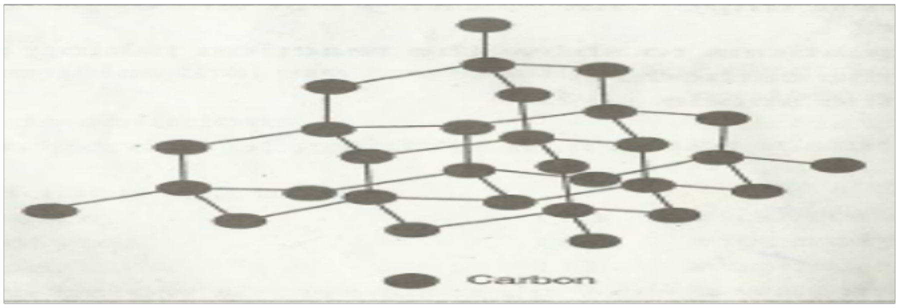
- Amorphous carbon does not conduct electricity or heat because of the irregular arrangement of atoms in the crystals in their structure.

Uses of amorphous carbon.

1. Charcoal is used in gas masks to absorb poisonous gases because it has the ability to absorb gases.
2. Charcoal is also used to absorb coloured impurities in sugar refining industries.
3. Soot is mixed with rubber as a dye in making of motor vehicle tyres.
4. Activated charcoal is used as a catalyst because it provides large surface area for reactant in some reaction.

Diamond;

- It is a colourless, transparent and shiny crystalline solid.
- It does not conduct both heat and electricity. This because carbon atoms in diamond uses all the four valence electron to form covalent bonds i.e.in the crystal, each carbon atom is bonded to four other carbon atoms.



- Diamond has high density of $3.51gcm^{-3}$.This is due to continuous close packing of carbon atoms.

- Diamond has a high melting point (3700^0C) and boiling point (4827^0C).This is due to the strong covalent bonds.
- It is the hardest substance in nature. This is due to the uniformity of covalent bonds between the atoms through-out the structure.
- It has a giant atomic structure; each atom is bonded to four other carbons arranged in a regular tetrahedron shape. It also extends in all directions forming a rigid mass of atoms.

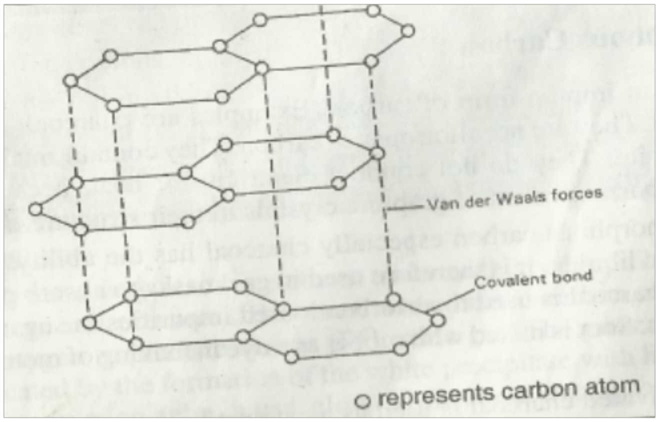
Uses of diamond

i. In manufacture of glass cutters and drilling bits because of its hardness.

ii. Used as a jewel because of its shinny appearance when polished.

Graphite;

- Graphite is composed of black, slippery and hexagonal crystals. This is because layers can slide over each other.



- Graphite has a low density of $2.3gcm^{-3}$.This is due to the spaces between the layer.
- Graphite conducts both electricity and heat. This is because, in graphite, a carbon atom uses only three electrons in the highest energy level to form covalent bonds with three carbon atom neighboring it. So one electron in the highest energy level of each atom is not used in bonding and hence leads to it being delocalized. The delocalized electrons are responsible for conduction of both electricity and heat.
- Graphite has a giant atomic structure (a planar structure of hexagonal rings).This structure contains weak van der Waal forces which hold the layers together. These forces between the layers allow the layers to slide over each other when pressed.

Uses of graphite

i. Graphite is mixed with clay to make pencil "leads" because the carbon layer can slide over each other.

ii. It is used as a lubricant in moving machine parts where a lot of heat is produced. This is because graphite has slippery nature, high melting point and boiling point.

iii. Graphite is used as electrode; this is because it has delocalized electrons within its structure which conducts electric current.

iv. Used in making moulds for casting metals; because of it high melting point.

Fullerenes;

- Fullerenes are synthetic allotrope of carbon.
- They are found in different shapes such as spherical, tubular or elliptical.
- The best known fullerene is the spherical C_{60} molecule. [Diagram pg 117, klb bk 2]
- It consist of hexagonally arranged carbon atoms similar to the panels of some soccer ball

Chemical properties of carbon

1. Carbon reacts with excess oxygen to produce carbon (IV) oxide;

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$$

2. Carbon removes combined oxygen from metals lower in the reactivity series. The carbon is oxidized to carbon (IV) oxide and the metal oxides are reduced to metals.

Metal oxide + Carbon → Carbon (IV) oxide + Metal

E.g. $2CuO_{(s)} + C_{(s)} \longrightarrow 2Cu_{(s)} + CO_{2(g)}$

3. At high temperatures carbon combines with many metals and non-metals.
With Aluminium; $3C_{(s)} + 4Al_{(s)} \longrightarrow Al_4C_{3(s)}$ (Aluminium tricarbide).

With Silicon; $Si_{(s)} + C_{(s)} \longrightarrow SiC_{(s)}$ (Silicon carbide).

With Sulphur; $C_{(s)} + 2S_{(l)} \longrightarrow CS_{2(l)}$ (Carbon disulphide).

4. Carbon reduces hot concentrated nitric (V) acid to brown Nitrogen (IV) oxide and water while itself is oxidized to carbon (IV) oxide.

$$C_{(s)} + 4HNO_{3(l)} \rightarrow CO_{2(g)} + 4NO_{2(g)} + 2H_2O_{(l)}$$

5. Hot concentrated sulphuric(VI) acid is reduced to sulphur(IV) oxide and water while carbon is oxidized to carbon(IV) oxide.

$$C_{(s)} + 2H_2SO_{4(l)} \rightarrow 2SO_{2(g)} + CO_{2(g)} + 2H_2O_{(l)}$$

NB: Both sulphur(IV) oxide and carbon(IV) oxide react with lime water to form white precipitate of calcium sulphite and calcium carbonate;

$$SO_{2(g)} + CaCO_{3(aq)} \rightarrow CaSO_{3(s)} + H_2O_{(l)}$$

$$CO_{2(g)} + CaCO_{3(aq)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$$

6. At a temperature of about 1200^0C, carbon reduces steam to hydrogen and is itself oxidized to carbon (II) oxide. A mixture of carbon (II) oxide and hydrogen is called water gas. $C_{(s)} + H_2O_{(g)} \rightarrow \underbrace{CO_{(g)} + H_{2(g)}}_{\text{Water gas}}$

SOME OF THE COMPOUNDS OF CARBON.

CARBON (IV) OXIDE, CO_2

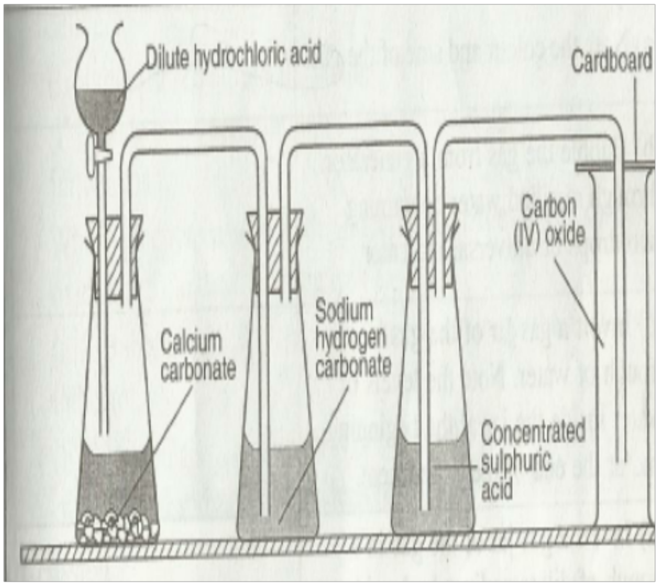
- Carbon (IV) oxide occurs in air to about 0.03%.It is also present in mineral springs, natural drinking water and volcanic regions.
- This is the most stable oxide of carbon. It forms strong covalent bonds with two oxygen; O-C-O, thus it is a molecular substance.

Laboratory preparation of carbon (IV) oxide

- Carbon (IV) oxide is prepared by reacting calcium carbonate with dilute hydrochloric acid. Effervescence occurs releasing carbon (IV) oxide gas.

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

- $NaHCO_3$ or water is used to remove traces of hydrochloric acid fumes.
- For dry carbon (IV) oxide gas, it is passed through concentrated sulphuric (VI) acid or anhydrous calcium chloride.

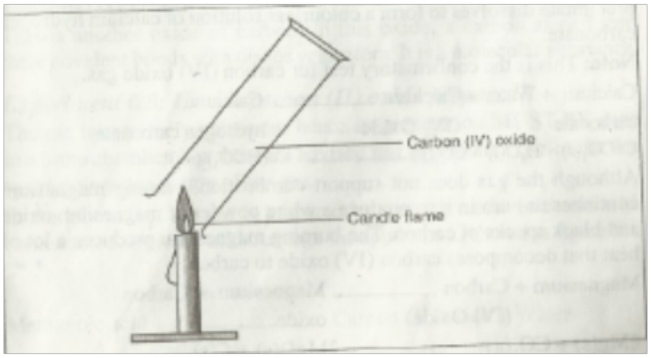


NB: i) Any other carbonate can be used instead of calcium carbonate.

ii) Acids which form insoluble salts of the metal are not suitable for the preparation of carbon (IV) oxide. This is because the insoluble salts formed coats these carbonates preventing further reaction and the evolution of carbon (IV) oxide e.g. Sulphuric (VI) acid should not be reacted with calcium carbonate, barium carbonate and lead carbonate; hydrochloric acid should not be reacted with silver carbonate and lead carbonate.

Physical properties

1) Colourless gas
2) Odourless
3) It is denser than air and hence collected by downwards delivery.
4) Neither burns nor supports combustions



5) The gas is acidic and turns universal indictor solution red.

Chemical properties

1) Carbon (IV) oxide dissolves slightly in water to form carbonic acid
   $CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$

2) Limited supply of Carbon (IV) oxide is acidic and readily reacts with sodium hydroxide solution to form sodium carbonate; $CO_{2(g)} + 2NaOH_{(aq)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$;Excess supply of carbon (IV) oxide leads to formation of sodium hydrogen carbonate; $CO_{2(g)} + NaOH_{(aq)} \rightarrow NaHCO_{3(aq)}$

3) A white precipitate is formed when carbon (IV) oxide is bubbled through limewater;
   $Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$ (This is the confirmatory test for the gas).

4) Though the gas does not support combustion, burning magnesium continues to burn in it to produce a white powder of magnesium oxide and black specks;

   $Mg_{(s)} + CO_{2(g)} \rightarrow MgO_{(s)} + C_{(s)}$; This because burning Magnesium has enough heat to decompose Carbon (IV) oxide into carbon and oxygen atom. The addition of dilute nitric (V) acid dissolves magnesium oxide leaving black specks of carbon.

Uses of carbon (IV) oxide

I. Used as a refrigerating agent for perishable goods (This is because solid carbon (IV) oxide sublimes on heating leaving no residue).

II. Used to extinguish fire (This is because carbon (IV) oxide is non-flammable and has a higher density than air and hence forms a 'blanket' on fire cutting of oxygen supply which is used for burning).

III. Used in the manufacture of sodium carbonate in the Solvay process.

IV. It is added to water to make aerated drinks thereby adding taste to the drink.

V. Used in making of baking powder.

CARBON (II) OXIDE, CO

In carbon (II) oxide, a carbon atom forms three covalent bonds (two perfect covalent and one dative) with on oxygen atom. Thus, it is a molecular substance.

Laboratory preparation of carbon (II) oxide

a) From reaction between methanoic acid and concentrated sulphuric (VI) acid.(diagram pg 140)

Carbon (II) oxide is prepared by melting HCOOH with Conc.H_2SO_4;

$$HCOOH_{(s)} \xrightarrow{\text{Conc. sulphuric(VI)acid}} CO_{(g)} + H_2O_{(l)}$$

The gaseous mixture should be passed through a concentrated solution of sodium hydroxide to remove carbon (IV) oxide.

The gas is then collected over water since it is insoluble in water.

NB: Ethanedioic acid may be used in place of methanoic acid;

$$H_2C_2O_{4(s)} \xrightarrow{\text{conc.Sulphuric(VI)acid}} CO_{(g)} + CO_{2(g)} + H_2O_{(l)}$$

b) By passing carbon (IV) oxide over heated charcoal.(diagram pg 140)

When carbon (IV) oxide is passed over heated charcoal as shown, Carbon (II) oxide is evolved; $CO_{2(g)} + C_{(s)} \xrightarrow{\text{heat}} 2CO_{(g)}$

Physical properties of carbon (II) oxide

1. Colourless gas.
2. Odourless
3. Combustible
4. Low boiling point of -111°C: Because of the weak forces of attraction between the molecule.
5. Slightly dissolves in water forming a neutral solution.
6. Extremely poisonous as it combines with haemoglobin in the blood to form a stable compound known a carboxyl-haemoglobin. This being stable, one suffocates and dies within a short time.

Chemical properties of carbon (II) oxide

1. Carbon(II) oxide burns in air to form carbon(IV) oxide;

$$2CO_{(g)} + O_{2(g)} \longrightarrow 2CO_{2(g)}$$

[Diagram pg 141,klb bk2]

2. Carbon (II) oxide is a strong reducing agent.It reduces metals oxides to the metals;

$$CuO_{(s)} + CO_{(g)} \longrightarrow Cu_{(s)} + CO_{2(g)}$$

$$Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

[Diagram pg 142,klb bk2]

NB: It is this property of carbon (II) oxide that makes the gas useful in the extraction of some metals from their oxides.

BURNING OF CHARCOAL (in a Jiko)

[Diagram pg 144,klb bk 2]

- Three reaction occurs;

1) Region A

   Oxidation of carbon to carbon (IV) oxide;

   $$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$$

2) Region B

   The carbon (IV) oxide produced in region A rises up region B where it is reduced;
   $CO_{2(g)} + C_{(s)} \longrightarrow 2CO_{(g)}$

3) Region C

   There is enough oxygen in this region therefore carbon (II) oxide produced in region B burns to form carbon (IV) oxide;

   $$2CO_{(g)} + O_{2(g)} \longrightarrow 2CO_{2(g)}$$

HYDROGEN CARBONATE AND CARBONATES, HCO_3^- and CO_3^{2-}

- Carbonates are salts obtained when metal ions or ammonium ions combine with carbonate ions.
- They occur naturally in rocks such as limestone, chalk, malachite and dolomite ($MgCO_3$).
- Hydrogen carbonates are salts derived from carbonic acid formed by partial displacement of hydrogen by a metal of ammonium radical.
- Sodium, potassium and ammonium carbonate exist as solid at room temperature, whereas

those of calcium and magnesium exist in solution form.

- Other metal hydrogen carbonates do not exist.

Properties of carbonates

1. All Carbonates give off carbon (IV) oxide when reacted with dilute acids e.g.

Exp : What happens when dilute acids are added to carbonate and hydrogen carbonate?

i) Set up the apparatus as shown below.[Diagram pg, kbl 145]

ii) Place a small amount of $CaCO_3$ in a test tube.

iii) Add dilute hydrochloric acid using a dropper.

iv) Repeat the experiment using $CuCO_3$, $ZnCO_3$, Na_2CO_3 and $NaHCO_3$ in place of $CaCO_3$.

v) Record your results in the table below.

Results

CO_3^{2-} and HCO_3^{2-}	HCl	HNO_3	H_2SO_4
$CaCO_3$	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**
$CuCO_3$	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**
$ZnCO_3$	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**
Na_2CO_3	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**
$NaHCO_3$	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**	✓ **A gas that turns lime water white precipitate.**

The equations involved are as follows;

$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

$ZnCO_{3(s)} + 2HNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

$Na_2CO_{3(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$ etc.

NB: Reaction between $CaCO_3$ and $PbCO_3$ with dilute H_2SO_4 does not go to completion. This is due to formation of insoluble $CaSO_4$ and $PbSO_4$ respectively which forms a coating on the carbonate and prevent further reaction.

2. All carbonates except those of potassium and sodium decompose when heated to produce a metal oxide and carbon (IV) oxide. The ease of decomposition depends on the position of the metal in the reactivity series i.e. those lower in the series are easily decomposed e.g.

$CuCO_{3(s)} \xrightarrow{Heat} CuO_{(s)} + CO_{2(g)}$

(Black)

$PbCO_{3(s)} \xrightarrow{Heat} PbO_{(s)} + CO_{2(g)}$

$ZnCO_{3(s)} \xrightarrow{Heat} ZnO_{(s)} + CO_{2(g)}$

(Yellow-hot; White-cold)

$CaCO_{3(s)} \xrightarrow{Heat} CaO_{(s)} + CO_{2(g)}$

- Ammonium carbonate decompose when heated to produce ammonia, carbon(IV) oxide and water;

$(NH_4)_2CO_{3(s)} \xrightarrow{Heat} 2NH_{3(g)} + CO_{2(g)} + H_2O_{(l)}$

- Hydrogen carbonates decompose on heating to produce carbon(IV) oxide, CO_2 and water e.g.

$2NaHCO_{3(s)} \xrightarrow{Heat} Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$

$2KHCO_{3(s)} \xrightarrow{Heat} K_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$

SODIUM CARBONATE AND SODIUM HYDROGEN CARBONATES.

- Sodium carbonate and sodium hydrogen carbonate are obtained in two ways;

  i) Extraction from Trona

  ii) Preparation by Solvay process.

i) Extraction from trona

- Trona is a double salt containing Na_2CO_3,$NaHCO_3$ and water of crystallization

  $\{Na_2CO_3.NaHCO_3.2H_2O\}$

- It is found in some lakes in the rift valley part of east Africa which have only in-lets but no out-lets e.g. Lake Magadi in Kenya.

- The flow chart below show how extraction of Trona is done from salty water

  [Diagram pg 150, klb bk 2]

- Lake Magadi contains large amounts of solid Trona and small amounts of NaCl.

- The water is pumped into shallow basins where evaporation takes place until the solution is about 14% NaCl.

- Then it is transferred to another basin for crystallization.

  NB: 1) During the day when the temperature is about 40^0C,Na_2CO_3 crystallizes out and is removed.

  2) During the night when the temperature is about 21^0C, NaCl crystallizes out and is removed. {Contains about 4% Na_2CO_3}

- The Trona is dug by buckets dredges and taken to the washery where it is washed in water to remove mud and small rocks.

- The wet Trona is then centrifuged to remove water and to purify it further.

- It is then heated to about 300^0C in a kiln to decompose it to Na_2CO_3.

  $$Na_2CO_3.NaHCO_3.2H_2O_{(s)} \xrightarrow{Heat} Na_2CO_3.NaHCO_{3(s)} + 2H_2O_{(g)}$$

  $$Na_2CO_3.NaHCO_{3(s)} \xrightarrow{Heat} 3Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$$

- The Na_2CO_3 obtained is grounded, sieved and bagged ready for storage and transportation to the consumers.

ii) Solvay process

- This process is used to produce sodium carbonate and sodium hydrogen carbonate in places where Trona is not available.

[Diagram pg]

Raw material

- ➢ Sodium chloride, NaCl
- ➢ Calcium carbonate, $CaCO_3$
- ➢ Coke , C
- ➢ Ammonia gas, NH_3

Step I

- Ammonia is dissolved in brine (Concentrated NaCl) in the Solvay tower to produce Ammoniacal brine.
- The formation of Ammoniacal brine is highly exothermic leading to the heat produced being controlled by regulating the flow of brine into the Solvay tower.

Step II

- The Ammoniacal brine is then pumped into the carbonator from the top where it trickles down the carbonator while carbon (IV) oxide from the kiln is pumped from the base.
- The carbonator has to be cooled regularly to reduce the amount of heat produced in it.
- The equation involved in the carbonator is as follows;

$$\underbrace{NH_{3(g)} + NaCl_{(aq)}}_{\boxed{\text{Ammoniacal brine from step I}}} + CO_{2(g)} + H_2O_{(l)} \rightarrow NaHCO_{3(s)} + NH_4Cl_{(aq)}$$

- Sodium hydrogen carbonate being less soluble than ammonium chloride at low temperatures, it forms crystals in the lower cooled parts of the carbonator.
- Sodium hydrogen carbonate is therefore separated out from ammonium chloride by

filtration.

- Sodium hydrogen carbonate is dried and decomposed by heat at 300^0C to produce sodium carbonate.

$2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$

Process in the Kiln

- The initial carbon (IV) oxide used in the process is mainly from the decomposed calcium carbonate and the burning of coke;

$$CaCO_{3(s)} \xrightarrow{Heat} CaO_{(s)} + CO_{2(g)}$$

$$C_{(s)} + O_{2(g)} \xrightarrow{Heat} CO_{2(g)}$$

- The calcium oxide produced in the kiln is slaked(combined) with water;

$$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(s)}$$

- The calcium hydroxide formed is then mixed with ammonium chloride from the carbonator to form ammonia gas, water and calcium chloride;

$$Ca(OH)_{2(s)} + 2NH_4Cl_{(aq)} \rightarrow CaCl_{2(s)} + H_2O_{(l)} + 2NH_{3(s)}$$

Recycled materials

- Carbon (IV) oxide gas produced as a result of heating sodium hydrogen carbonate.
- Ammonia gas produced when calcium hydroxide is reacted with ammonium chloride.
- Water produced in the heating of sodium hydrogen carbonate and the reaction of calcium hydroxide and ammonium chloride.

By- products

- The only by-product in this process is calcium chloride.

Efficiency of the process

- All by-products except calcium chloride are recycled to minimize cost.
- The process does not involve large consumption of energy because the coke burned in the kiln supply the necessary heat in the process.
- The only by-product in this process is calcium chloride which has a wide range of uses.

NB: Potassium carbonate cannot be prepared using Solvay process since potassium hydrogen carbonate is fairly soluble in water and therefore cannot precipitate out.

Uses of sodium carbonate [Na_2CO_3]

1. Water softening.
2. Glass making.
3. Paper making.
4. Making of sodium silicate used in making of detergents.

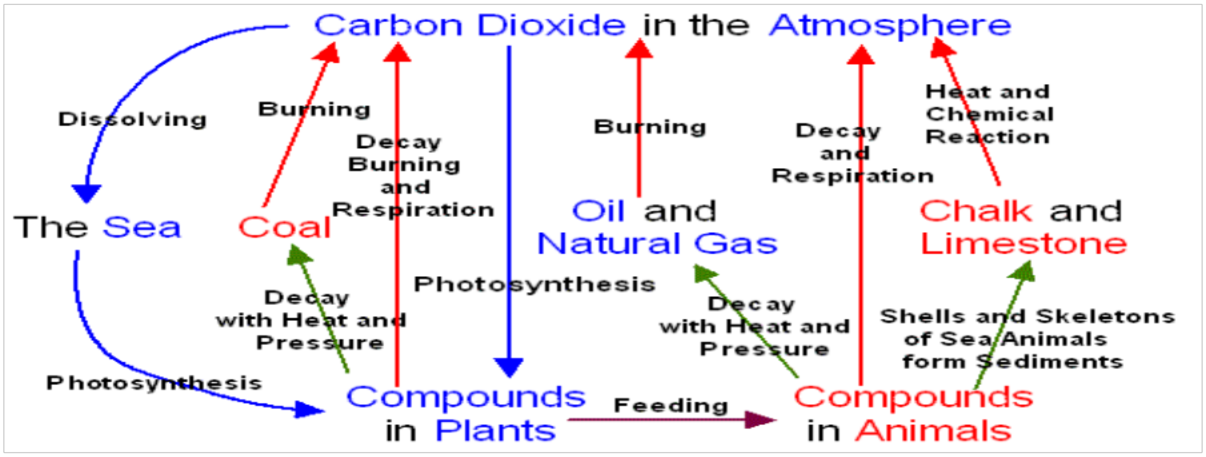
Uses of sodium hydrogen carbonate [$NaHCO_3$]

1. Making health drinks.
2. Making baking powder in food industry.

Effects of carbon (IV) oxide and carbon (II) oxide on the environment.

- High levels of carbon (IV) oxide in the atmosphere is the cause of global warming which results in the melting of ice caps leading to rise in level of the seas.
- Carbon (II) oxide is oxidized to carbon (IV) oxide in presence of abundant oxygen but in a poorly ventilated room, it may cause death through suffocation.

Carbon cycle



- The amount of carbon (IV) oxide in the atmosphere remains fairly constant at 0.03% owing to the delicate balance between the processes that absorb the gas and those that evolve it.

- The ways in which carbon compounds circulate in nature constitute the carbon cycle which is shown in the diagram above.
- Processes that increase carbon (IV) oxide in the atmosphere include;
    - Combustion of fuels
    - Respiration in both animals and plants
    - Fermentation
    - Decomposition of carbonates and hydrogen carbonates
- Those processes that reduce the amount of carbon (IV) oxide in the atmosphere include;-
    - Photosynthesis in plants
    - Dissolving of carbon (IV) oxide in seas and oceans