S2 CHEMISTRY NOTES

ATOMIC STRUCTURE

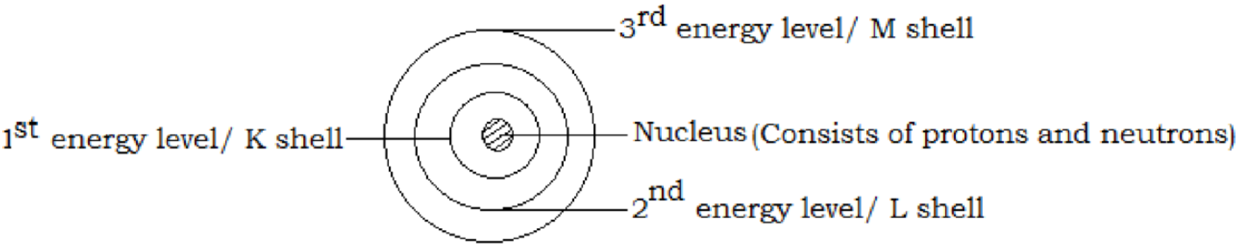
An atom is the smallest particle of an element that takes part in a chemical reaction. All atoms of the same element are identical and different from those of other elements.

Shape of an atom

An atom is spherical in shape and has a small region in the centre called the **nucleus**. The nucleus is surrounded by circular paths known as **electron shells** or **energy levels** that carry electrons. These shells are represented as circles.

The energy level (shell) closest to the nucleus is called the 1st energy level or the K shell. The 2nd closest shell to the nucleus is the 2nd energy level or the L shell. This is followed by the 3rd shell or the M shell. So, the energy levels are named 1st, 2nd, 3rd, 4th, e.t.c. or K, L, M, N e.t.c. from the shell closest to the nucleus outwardly.

General structure of an atom



An atom is composed of particles namely; protons, neutrons and electrons.

Protons (p)

Protons occur in the nucleus; Protons have a unit mass of one (1); each proton carries a charge of positive one (+1).

The total number of protons in the nucleus of an atom is the **atomic number**. The atomic number of each element is unique and it is used to identify the element. All atoms of the same element have the same atomic number.

Neutrons (n)

Neutrons occur in the nucleus of an atom; it has a unit mass of one i.e. the mass is approximately the same as that of a proton; a neutron carries no charge(it is nuetral). The sum of all protons and all neutrons in the nucleus of an atom is called the **atomic mass** or **mass number**.

Electrons (e)

Electrons occur in the energy levels or electron shells and are constantly rotating around the nucleus; an electron has a mass of 1/1840 of a proton, there fore its mass is said to be negligible; it carries a charge of negative one(-1).

N.B. The number of electrons in the energy levels (negative charges) is equal to the number of all protons in the nucleus (positive charge) hence an atom is neutral because the number of positive charges(protons) is equal to the number of negative charges(electrons).

Mass number and Atomic number

Atomic number (Z)

Every atom has its specific atomic number. This atomic number is the same for all atoms of the same element.

Atomic number is the total number of protons in the nucleus of an atom (of an element). An atom being neutral has equal number of protons and electrons.

Mass number/Atomic mass (A)

Mass number is the sum of protons and neutrons in the nucleus of an atom of an element. The protons and neutrons occupy the nucleus and are referred to as **nucleons**.

Mass number = number of protons + number of neutrons

Notation of an atom

An atom can be represented as

$^{A}_{Z}W$ Where W is the symbol of an element; A is the atomic mass or mass number and Z is the atomic number.

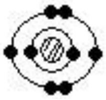
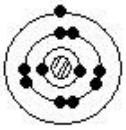
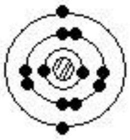
Electron arrangement in atoms

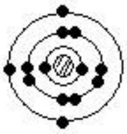
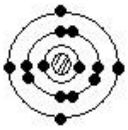
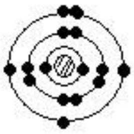
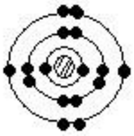
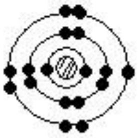
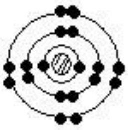
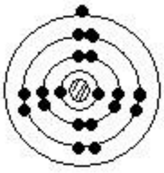
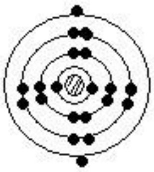
The arrangement of electrons in an atom is the **electronic configuration** and structures drawn showing the distribution of electrons in the energy levels is the electronic structure. In writing electronic configurations or structure, the following rules are considered.

1. The 1st electron shell (one that is nearest to the nucleus) has the lowest energy level and electrons fill it first.
2. The 1st electron shell or energy level holds a maximum of two electrons. It cannot hold more than two electrons.
3. The 2nd electron shell holds a maximum of eight electrons. It can hold less the maximum number but never beyond the maximum number.
4. In the outer most shell of any atom, the maximum number of electrons possible is 8.

Electron shells are numbered 1, 2, 3, etc, or L, M, N e.t.c.outwards from the nucleus. All electrons in a given shell have approximately equal energy. This energy increases in successive shell outwards from the nucleus.

Table showing the electron arrangement of the first 20 elements

Element	**Symbol**	**No. of electrons**	**No. of electrons in each shell**				**Electronic configuration**	**Electronic structure**
			1st	2nd	3rd	4th		
Hydrogen	H	1	1				1	
Helium	He	2	2				2	
Lithium	Li	3	2	1			2;1	
Beryllium	Be	4	2	2			2;2	
Boron	B	5	2	3			2;3	
Carbon	C	6	2	4			2;4	
Nitrogen	N	7	2	5			2;5	
								
Oxygen	O	8	2	6			2;6	
Fluorine	F	9	2	7			2;7	
Neon	Ne	10	2	8			2;8	
Sodium	Na	11	2	8	1		2;8;1	
Magnesium	Mg	12	2	8	2		2;8;2	

Aluminium	Al	13	2	8	3		2;8;3	
Silicon	Si	14	2	8	4		2;8;4	
Phosphorus	P	15	2	8	5		2;8;5	
Sulphur	S	16	2	8	6		2;8;6	
Chlorine	Cl	17	2	8	7		2;8;7	
Argon	Ar	18	2	8	8		2;8;8	
Potassium	K	19	2	8	8	1	2;8;8;1	
Calcium	Ca	20	2	8	8	2	2;8;8;8;2	

Isotopes and Isotopy

Isotopes are atoms of the same element with the same number of protons but different number of neutrons. Isotopes therefore have different mass numbers. For elements that show isotopy, the most abundant (common) isotope is taken to be the representative of all the element. The abundance is usually given as a percentage.

Isotopy is the existence of atoms of the same element with the same number of protons but different number of neutrons.

Examples of isotopes

Element	**Symbol**	**Atomic number**	**Isotopes**	**Abundance**
Hydrogen	H	1	${}^{1}_{1}H$ ${}^{2}_{1}H$ ${}^{3}_{1}H$	99.99% 0.01% Rare
Carbon	C	6	${}^{12}_{6}C$ ${}^{13}_{6}C$ ${}^{14}_{6}C$	98.9% 1.1% Trace
Chlorine	Cl	17	${}^{35}_{17}Cl$ ${}^{37}_{17}Cl$	75% 25%

Example
Copy and complete the tabe below.

Relative Atomic Mass (RAM)
This is the mass of one atom of an element compared to 1/12 (a twelfth) of the mass of one atom of Carbon 12 isotope.

$$R.A.M = \frac{mass\ of\ an\ atom\ of\ \ an\ element}{\frac{1}{12} of\ the\ mass\ of\ one\ atom\ of\ C{-}12\ isotope}$$

Relative Atomic Mass of an element that shows isotopy is dependent on the relative abundance of each isotope and the atomic masses of the isotopes. R.A.M is a ratio of the same quantity and therefore has no units.

Example
Neon has three isotopes,${}^{20}_{10}Ne$(percentage abundance 90.5%);${}^{21}_{10}Ne$ (percentage abundance of 0.3%) and ${}^{22}_{10}Ne$ (percentage abundance of 9.2%)

Solution

$\left(\frac{90.5\times20}{100}\right)+\left(\frac{0.3\times21}{100}\right)+\left(\frac{9.2\times22}{100}\right)$

$\left(\frac{(90.5\times20)+(0.3\times21)+(9.2\times22)}{100}\right)$

$\left(\frac{2018.7}{100}\right)$ =20.187.

THE PERIODIC TABLE

The periodic table is the arrangement of elements in order of increasing atomic masses. The periodic table is made up of rows and columns. The horizontal rows are called **periods** and are assigned numbers according to the number of electron shells/energy levels that are filled by electrons. All elements in the same period have the same number of electron shells /energy levels. The periods are written in roman numbers as 1, 2, 3, etc. The vertical columns are referred to as **groups** and the group number is written in Roman numeral on top of the group (column). The group number represents the number of electrons in the outer most shell/energy level. Elements in a particular group resemble each other chemically because they have the same number of electrons on the outer most shell/energy level.

Periodic table of the first 20 elements

Group

Period				
	4	K	Ca	

The elements in the periodic table are classified as metals, metalloid, non metals and noble/inert/rare gases.

i) **Metals**: these are on the left of the periodic table in group I, II and III. Metals react by loss of electrons.
ii) **Semi metals (metalloids)**: These are found in the middle of the periodic table in group IV.
iii) **Non metals**: they are on the right of the periodic table in groups V, VI and VII. Non metals react by gain of electrons.
iv) **Noble/rare of inert gases**: They are at the extreme right of the periodic table in group VIII (O). Non metals are not reactive because the are very stable.

Transition elements: these are found between group II and III and towards the bottom of the periodic table. Transition elements form

	I	II		III	IV	V	VI	VII	VIII
1	H								He
2	Li	Be		B	C	N	O	F	Ne
3	Na	Mg		Al	Si	P	S	Cl	Ar

colored compounds and have variable valencies. Examples of transition elements include copper, iron, chromium, cobalt etc.

Trend of reactivity of elements in the periodic table

1. Elements in the same group react in similar ways because they have the same number of electrons in their outer most energy level/ shell.
2. Reactivity increases as you go down the group of metals because of increase in atomic size (atomic radius). As the size of atoms increase, the outer most electrons become loosely held because they are no longer strongly attracted by the nucleus hence such electrons are easily lost. Because metals react by loss of electrons, the more easily a metal loses electrons, the more reactive that metal is.
3. As you go down the group of non metals, reactivity decreases due to increase in atomic size (atomic radius). As the atomic sizes increase, the forces of attraction of the nucleus for the electrons decrease hence electrons cannot be gained readily. So, the more easily a non metal gains electrons, the more reactive that non metal is.
4. Across the period of elements from left to right, reactivity decreases from group I to group IV and then increases up to group VII and falls sharply in group VIII. This is because, as you go across the period, the effective force of the nucleus increases hence the outer most electrons become more strongly attracted.

IONS

An ion is a charged particle formed from an atom or a group of chemically combined atoms by gaining or losing one or more electrons. Ions are either negatively charged (anions) or positively charged (cations).

Ion formation

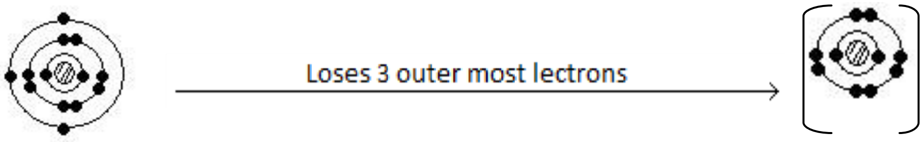
a) Positive ions(Cations)

These are formed when atoms or group of chemically combined atoms lose electrons. Atoms lose electrons so as to acquire a stable electronic configuration as the noble gases. The positive ions (cations) are formed by metal atoms, hydrogen atoms and ammonium.

i) Formation of Aluminium ion, Al^{3+}

Aluminium atom with 13 electrons (electronic configuration 2.8.3), has 3 electrons more the stable configuration of noble gases,(2.8). During reaction, Aluminium loses the 3 electrons in the outer most energy level (electron shell) to become stable and form aluminium ion, Al^{3+}, (electronic configuration 2.8)

Aluminium atom (Al), unstable Aluminium ion((Al^{3+}), stable
Electronic configuration 2.8.3 Electronic configuration 2.8
Charge=0 Charge=+3



Aluminum atom $\xrightarrow{\text{Loses 3 outer most electrons}}$ Aluminiumion

$\longrightarrow$

Al–3e $\longrightarrow$

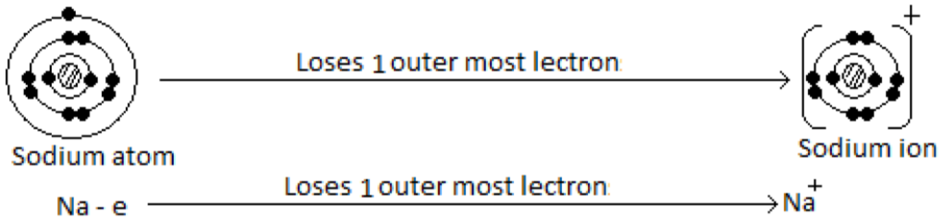
(Al + 3e)

ii) Formation of sodium ion, Na^+

Sodium atom with 11 electrons (electronic configuration 2.8.1), has 1 electrons more the stable configuration of a noble gases, (2.8). During reaction, sodium loses the 1 electron in the outer most energy level (electron shell) to become stable and form sodium ion,Na^+,(electronic configuration 2.8)

Sodium atom (Na), unstable Sodium ion ((Na^+), stable Electronic configuration 2.8.1 Electronic configuration 2.8

Charge=0 Charge=+1



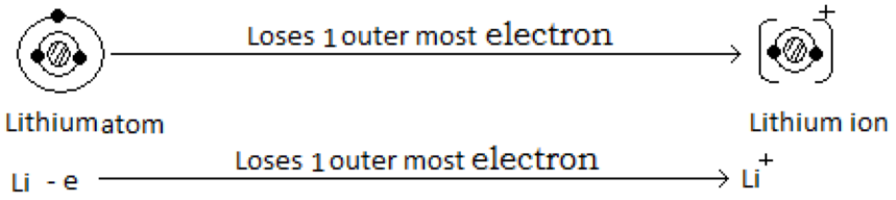
(Na $\longrightarrow$ + e)

iii) Formation of lithium ion,Li^+

Lithium atom with 3 electrons (electronic configuration 2.1), has 1 electrons more the stable configuration of a noble gases, (2). During reaction, lithium loses the 1 electron in the outer most energy level (electron shell) to become stable and form lithium ion,Li^+,(electronic configuration 2)

Lithium atom (Li), unstable Lithium ion((Li^+), stable Electronic configuration 2.1 Electronic configuration 2

Charge=0 Charge=+1



(Li $\longrightarrow$ + e)

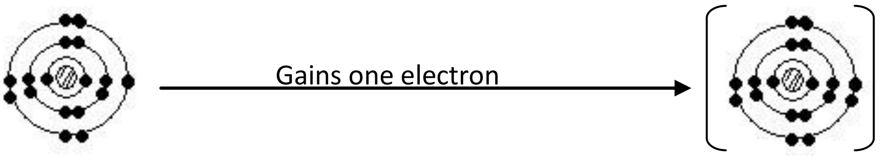
b) Formation of negative ions (Anions)

These are formed when atoms gain electrons. Atoms gain electrons so as to acquire a stable electronic configuration as the noble gases. The negative ions (anions) are formed by non metal atoms). Examples of anions are chloride ion and oxide ion.

i) Formation of a chloride ion, Cl^-

The ion is formed from chlorine atom. Chlorine atom (electronic configuration 2,8.7) lacks only one electron to attain the stable electronic configuration of a noble gas(2.8.8). The chlorine atom there fore gains the one electron during reaction and forms a chloride ion.

Chlorine atom (Cl), unstable	Chloride ion ((Cl^-), stable
Electronic configuration 2.8.7	Electronic configuration 2.8.8
Charge=0	Charge=-1



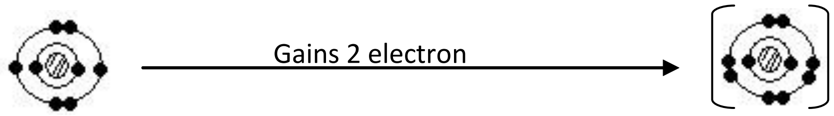
Chlorine atom Chloride ion –

$$Cl + e^- \xrightarrow{\text{Gains one outer electron}} Cl^-$$

ii) Formation of an oxide ion, O^{2-}

The oxide ion is formed from an oxygen atom. Oxygen atom (electronic configuration (2,6) lacks two electrons to attain the stable electronic configuration of a noble gas(2.8). The oxygen atom therefore gains the two electrons during reaction and forms an oxide ion, O^{2-}.

Oxygen atom (O), unstable	Oxide ion ((O^{2-}), stable
Electronic configuration 2.6	Electronic configuration 2.8
Charge=0	Charge=-2



Oxygen atom $\xrightarrow{\text{Gains two electrons}}$ Oxide ion

$$O + 2e^- \longrightarrow O^{2-}$$

VALENCY

Valency is the combining power of an element or radical. In other words it is the measure of the power of an element to combine with others. The power of an element to combine with others is determined by the structure of its atoms. This therefore implies that the valency of an element can be determined from the electron structure of its atoms.

Elements with atoms having 1-4 electrons in the outer most shell, their valencies are equal to the number of electrons in the outer most shell.i.e. Group I elements have a valency of 1; group II elements have a valency of 2; group III elements have a valency of 3 and group IV elements have a valency of 4.

Elements in group V-VII, their valencies are equal to the number of electrons gained to acquire a stable electronic configuration i.e. Elements in group V gain three electrons each to make them stable, they therefore have a valency of 3; Elements in group VI gain two electrons each to make them stable, they therefore have a valency of 2 and elements in group VII gain one electron each to make them stable therefore have a valency of 1.

NB. Therefore, valency can also be defined as the number of electrons lost or gained by an element in order to attain a stable electronic structure.

Examples of elements with their valencies

Valency 1	**Valency 2**	**Valency 3**	**Valency 4**
Hydrogen (H)	Beryllium (Be)	Nitrogen (N)	Carbon (C)
Lithium (Li)	Magnesium (Mg)	Phosphorus (p)	Manganese (Mn)
Sodium (Na)	Calcium (Ca)	Boron (B)	Lead (Pb)
Potassium (K)	Oxygen (O)	Aluminium (Al)	Silicon (Si)
Silver (Ag)	Sulphur (S)	Iron (Fe)	Tin (Sn)
Copper (Cu)	Barium (Ba)		
Chlorine (Cl)	Zinc (Zn)		
Bromine (Br)	Iron (Fe)		
Iodine (I)	Lead (Pb)		
	Copper (Cu)		
	Mercury (Hg)		

NB. For metals with variable valencies, Roman numerals are included in their name to indicate the valencies of the metal. E.g. Iron (II) has a valency of 2,Lead (IV) has a valency of 4 etc.

RADICALS

A radical is an atom or group of atoms that exist in several compounds but does not exist on its own.

Valencies of common radicals

Radical	**Formula**	**Ion**	**Valency**
Hydroxide	OH	OH^-	1
Chloride	Cl	Cl^-	1

Nitrate	NO_3	NO_3^-	1
Nitrite	NO_2	NO_2^-	1
Ammonium	NH_4	NH_4^+	1
Hydrogen carbonate	HCO_3	HCO_3^-	1
Chlorate	ClO_3	ClO_3^-	1
Permanganate	MnO_4	MnO_4^-	1
Hydrogen sulphate	HSO_4	HSO_4^-	1
Hydrogen sulphite	HSO_3	HSO_3^-	1
Oxide	O	O^{2-}	2
Carbonate	CO_3	CO_3^{2-}	2
Sulphate	SO_4	SO_4^{2-}	2
Sulphite	SO_3	SO_3^{2-}	2
Sulphide	S	S^{2-}	2
Phosphate	PO_4	PO_4^{3-}	3
Nitride	N	N^{3-}	3

CHEMICAL FORMULAE

A chemical formula is a representation showing the proportions of elements present in a chemical compound using symbols.

Important concepts about chemical formulae

1. In a chemical formula, the number written as a subscript after the symbol of an element indicates the number of atoms of that element chemically combined.

Consider examples of the chemical formulae below.

In water, H_2O-the two (2) represents the number of hydrogen atoms chemically combined. There fore, water consist of two (2) hydrogen atoms and one (1) oxygen atom chemically combined together.

In sulphuric acid (H_2SO_4)- there are 2 atoms of hydrogen, 1 atom of sulphur and 4 atoms of oxygen chemically combined together.

$CaCO_3$ is the chemical formula of calcium carbonate (chalk), which consists of 1 atom of calcium, 1 atom of carbon and 3 atoms of oxygen chemically combined.

2. For groups of atoms (radicals), a bracket is used showing that they are being considered under the same valency.

   Examples

   In calcium nitrate, $Ca(NO_3)_2$,the 2 indicates that there are 2 nitrate radicals (NO_3). Both nitrogen and oxygen in nitrate are being considered under the same valency.

   In aluminium sulphate, $Al_2(SO_4)_3$,-the 3 shows that there are 2 sulphate radical (SO_4), both the sulphur and oxygen in the sulphate are being considered under the same valency. There fore, $Al_2(SO_4)_3$ consists of 2 atoms of aluminium and 3 sulphate radicals.

3. A number put infront of the formula of a compound indicates the number of molecules of the compound. For example
   $2H_2SO_4$ this means two molecules of sulphuric acid
   $8HNO_3$ means eight molecules of nitric acid
   $3CO_2$ means three molecules of carbondioxide
   H_2O means one molecule of water

4. To be able to write a chemical formula, one has to know the symbol and valency of the atoms or radicals.

Steps taken in writing chemical formulae

1. Identify from the name of the compound, the elements and radicals present e.g. in sodium chloride, there is sodium and chloride radical.
2. Write the symbol of the element or the formula of the radical separately, beginning with the one of a metal or ammonium radial followed by a non metal or radical.

   Na Cl

3. Write the valencies as superscripts to the right of the symbols written separately.

   Na^1 Cl^1

   Compare the valencies and if the valencies are;

i) If the valencies are the same, counsel them out and write the symbols/formula close to each other.

   $Na^{\cancel{1}}$ $Cl^{\cancel{1}}$
   NaCl

ii) If the valencies are not the same, but have a common factor, reduce them to the simplest ratio then interchange the valencies and write them as subscripts to the right of the symbol or formula of the radical. e..g. in Lead (IV) oxide

   Pb^{42} $O^{\cancel{2}1}$

P
b
2

O
1

Pb_1O_2 the 1 is usually ignored and the formula is written as PbO_2.

iii) If the valencies are not the same and have no common factor, interchange them and write them as subscripts to the right of the symbol or formulae of radical. e.g. in aluminium oxide

Al^3 O^2

Al^3 O^2

Al_2O_3

N.B. If the formula of the radical consists of more than one symbol, the formula must be put in brackets before the subscript is written e.g. in ammonium sulphate

$NH_4{}^1$ $SO_4{}^2$

$(NH_4)_2SO_4$

Example

Write the formulae of the following compounds

a) Sodium chloride b) sodium sulpate c) Iron(III) oxide

Solution

a) Sodium Chloride
Na Cl
$Na^{\cancel{1}}$ $Cl^{\cancel{1}}$
NaCl

b) Sodium Sulphate
Na SO_4
Na^1 $(SO_4)^2$

Na_2SO_4

c) Iron(III) Oxide
Fe O
Fe^3 O^2
Fe^3 O^2

Fe_2O_3

Exercise

Write the formula of the following compounds

a) Potassium hydroxide b) Zinc carbonate c) magnesium sulphate d) ammonium phosphate e) copper(II) oxide f)potassium permanganate g) calcium hydroxide.

CHEMICAL EQUATIONS

This is the representation of a chemical change (chemical reaction) by means of symbols and formulae. So, when chemical equations are written, symbols and formulae are normally used. Word equations are not significant and not commonly used. A chemical reaction consists of reactants and products.

Consider the example below for illustration
Hydrogen reacts with oxygen to produce water vapor

Word equation: Hydrogen + Oxygen $\longrightarrow$ water vapor
Symbol equation: $\underbrace{2H_2(g) + O_2(g)}_{\text{Reactants}} \longrightarrow \underbrace{2H_2O(g)}_{\text{Product}}$

Components of a chemical equation

a) **Formulae of reactants and products**. These represent compounds reacting and compounds being formed during a chemical reaction respectively.
b) **State symbols**. A state symbol is one small letter or two small letters that represent the physical state of a compound. They are written after the formulae and enclosed in brackets. There are four state symbols used. These are aqueous solution (aq); liquid (l); solid state(s) and gaseous state (g).
c) **The plus (+) sign**. The plus sign on the left hand side of the equation means ‗react with‘ and the one at the right hand side of the equation means ‗and‘ .
d) **The arrow**. This means to produce and the arrow head points to the products.

Interpretation of equations

$2Mg(s) + O_2(g) \longrightarrow 2MgO(g)$

This means 2 molecules of solid magnesium reacts with 1 molecule of oxygen gas to produce 2 molecules of solid magnesium oxide.

$2H_2O_2(aq) \longrightarrow 2H_20(l) + O_2(g)$

2 molecules of aqueous hydrogen peroxide produce 2 molecules of liquid water and gaseous oxygen.
For an equation to be considered correct, it must have correct formulae of reactants and products; correct state symbols and must be balanced.

Balancing is the process of making each kind of atoms on both the reactant and product sides equal. A balanced equation there fore has equal numbers of each kind of atoms on both the reactant and product sides.
N.B. When balancing a chemical equation, the formula of the compound is not altered but rather it is the number of molecules that is adjusted.

Steps in writing chemical equations

1. Write the correct formulae of the reactants on the left hand side and the correct formulae of products on the right hand side of the

equation. The reactants and products are separated by an arrow pointing to the produncts.

e.g. consider the reaction calcium and hydrochloric acid producing calcium chloride and hydrogen gas.

$Ca + HCl \longrightarrow CaCl_2 + H_2$

2. Write the state symbol after each formula or symbol and enclose them in brackets.

   $Ca(s) + HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$

3. Balance the equation by making the number of each kind of atoms on both reactant and product side equal. Adjust the number of molecules of the reactants and products to balance the number of atoms.

   $Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$

Exercise

1. Write the equation of reaction between the following

a) Sodium + chlorine → Sodium chloride
b) Zinc + Sulphur → Zinc sulphide
c) Magnesium + Oxygen → Magnesium oxide
d) Hydrogen + Chlorine → Hydrogen chloride
e) Copper + chlorine → Copper chloride

2. Complete and balance the following equations.

a) $Pb(s) + \longrightarrow O_2(g)$
b) $H_2(g) + \longrightarrow O_2(g)$
c) $Na(s) + H_2O(l) \longrightarrow$
d) $Mg(s) + CO_2(g) \longrightarrow$
e) $Fe(s) + \longrightarrow HCl(aq)$

ATOMICITY

Atomicity is the number of atoms in a molecule or compound.

A molecule is the smallest particle of an element or compound that can exist on its own(i.e. in a free and separate state).

A molecule with one atom is referred to as a **monoatomic** molecule e.g. Helium(He); argon(Ar);sodium(Na) etc.

A molecule containing two atoms is referred to as a **diatomic** molecule e.g. Oxygen(O_2); hydrogen(H_2); Chlorine(Cl_2) etc.

A molecule containing three atoms is said to be **triatomic** e.g. Ozone (O_3) A molecule containing four atoms is said to be **tetraatomic** e.g. Phosphorus (P_4) A molecule containing many atoms is said to be **polyatomic** e.g. Sulphr (S_8).

Example

Find the atomicity of the following elements in the compound. a) $4CO_2$ b) $5H_2O$ c) $Zn(NO_3)_2$

Solutions

a) Carbon atoms =4, oxygen atoms ,4x2=8
Total number of atoms =4+8=12.

b) Hydrogen atoms, 5x2=10atoms, oxygen atoms =5 Total number of atoms =10+5=15 atoms.

c) Zinc atoms=1, nitrogen atoms =2, and oxygen atoms,3x2=6
Total number of atoms= 1+2+6=9 atoms

Exercise

Find the number of atoms of each element present in the following compounds. a) $MgCO_3$ b) $Pb(OH)_4$ c) $Ca_3(PO_4)_2$ d) $(NH_4)_2CO_3$ e) $2H_2SO_4$

Sample questions on the atomic structure and the periodic table

Atomic structure and periodic table

1. *With an aid of a labeled drawing, describe the structure of an atom showing the distribution of the different particles that make it up.*
2. *What are protons, neutrons and electrons? Give a brief detail of each.*
3. *Differentiate between mass number and atomic number, isotope and isotopy. Give three elements that show isotopy, write down and name the isotopes.*
4. *The figure below shows part of the periodic table. The letters used are not the correct symbols of the elements.*

I	*II*	*III*	*IV*	*V*	*VI*	*VII*	*VIII*
			P		*T*	*R*	
S							*Q*

Which of the elements are metals? Suppose element P reacts with element T, write the formula of the compound formed between P and T. Which element in the table is least reactive and explain why. Suggest a compound formed between any two elements shown, which would conduct electricity, give a reason for your answer.

5. *Silver is an element which exists naturally as a mixture of two isotopic forms. A and B represent atoms of these isotopes. They occur in equal numbers. A is* $^{107}_{47}Ag$ *and B is*$^{109}_{47}Ag$*. State the number of (a) protons, (b) electrons and (c) neutrons in atoms A and B. What is the relative atomic mass of the naturally occurring silver (Ans. 108(Atomic mass).*

6. *Elements X, Y, and Z have the following atomic number, 11, 15 and 19 respectively. Write the electronic configuration of the elements. State each element is a metal or non metal.*
7. *Metals lose electrons to become stable like the noble gases while non metals gain electrons, in the process they form ions. Write down the electronic configuration of the ions formed by magnesium, chlorine and calcium.*

ACIDS, BASES AND INDICATORS

ACIDS

An acid is a compound which when dissolved in water produces hydrogen ions (H^+) as the only positively charged ion. There are basically two types of acids; mineral and organic acids.

a) Mineral acids

These are called mineral acds because they ar derived from minerals, examples include, the common laboratory acids like sulphuric acid (H_2SO_4), hydrochloric acid (HCl), nitric acid (HNO_3) and phosphoric acid(H_3PO_4). **b) Organic acids**

These are acids derived from organic compounds and examples include: ethanoic acid (CH_3COOOH) found in vines, citric acid found in fruits and lactic acid found in milk.

Properties of acids

1. Acids have a sharp sour taste.
2. Acids change colors of indicators e.g. turn blue litmus paper red.
3. Acids ionize in water to produce hydrogen ions (H^+).

Sulphuric acid: $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}$
(aq) Hydrochloric acid: $HCl(aq) \longrightarrow H^+(aq) + Cl^-$
aq)

4. Most acids are corrosive (i.e. have burning effects) and poisonous
5. Dilute acids react with reactive metals to produce hydrogen gas and a salt.

Acid + Metal $\longrightarrow$ Salt + Hydrogen gas
For example
$H_2SO_4(aq) + 2Na(s) \longrightarrow Na_2SO_4(aq) + H_2(g)$

6. Acids react with carbonates and hydrogen carbonates to liberate carbondioxide gas. When an acid reacts with a carbonate,a salt, water and carbondioxide gas are produced.

Acid + Carbonate $\longrightarrow$ Salt + Water + Carbondioxide

Example
$HCl(l) + Na_2CO_3(s) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$
$H_2SO_4(aq) + 2NaHCO_3(s) \longrightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(l)$

7. Acids react with bases to to form salt and water

Acid + Base $\longrightarrow$ Salt + Water

Examples

$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

$H_2SO_4(aq) + CuO(s) \longrightarrow CuSO_4(aq) + H_2O(l)$

Preparation of acids

1. By the reaction between an acid anhydride (an oxide of a non metal) and water. Example

Sulphur dioxide + water $\longrightarrow$ Sulphurous acid $SO_2(g) + H_2O(l)$
$H_2SO_3(aq)$ Phosphorus(V)oxide + Water $\longrightarrow$ Phosphoric acid
$P_4O_{10}(s) + 6H_2O(l) \longrightarrow 4H_3PO_4(aq)$

2. By displacing a weaker acid (more volatile acid) from its salt by a stronger acid (less volatile acid).

Example

$2NaCl(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2HCl(aq)$

3. By precipitating an insoluble sulphide from a metallic salt by hydrogen sulphide.

For example

$Pb(CH_3COO)_2(aq) + H_2S(g) \longrightarrow PbS(s) + 2CH_3COOH(aq)$

Lead ethanoate Ethanoic acid

Basicity of an acid

This is the number of hydrogen ions that can be produced when one molecule of an acid ionizes in water. Acids can be categorized interms of basicity as: **a) Monobasic acid**

Is an acid that ionizes to produce one hydrogen ion (H^+). Their basicity is one.

Examples

Nitric acid: $HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$

Hydrochloric acd: $HCl(aq) \longrightarrow H^+(aq) + Cl^-$aq)

Ethanoic acid: $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

b) Dibasic acids

These acids ionize to produce two hydrogen ions ($2H^+$) from one molecule. They have a basicity of two. Examples include;

Sulphuric acid: $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$

Carbonic acid: $H_2CO_3(aq) \longrightarrow 2H^+(aq) + CO_3^{2-}(aq)$

c) Tribasic acids

These are acids that ionize to produce three hydrogen ions ($3H^+$). They have basicity of 3. For example:

Phosphoric acid: $H_3PO_4(aq) \longrightarrow 3H^+(aq) + PO_4^{3-}(aq)$

Strength of acids

Strength of an acid refers to the ease with which an acid ionize to produce hydrogen ions. According to strength, acids can be categorized as strong or weak.

1. **Strong acid:** Is an acid that dissociates completely into ions when dissolved in water (it ionizes completely in water). Examples include:

Sulphuric acid: $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$

Hydrochloric acid: $HCl(aq) \longrightarrow H^+(aq) + Cl^-$aq)

2. **Weak acid:** Is an acid that does not completely ionize in water. Some of the acid molecules do not ionize, examples include:

Carbonic acid: $H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$

Ethanoic acid: $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

Uses of acids

i) Acids are used in car batteries ii) Sulphuric acid is used in the manufacture of soap, detergents and paints. Plastics e.t.c.

iii) Ethanoic acid is used in the preservation of fruits and vegetables.

iv) Hydrochloric acid in the stomach of human being helps in the digestion of food. v) Acids are used in the manufacture of fertilizers e.g. sulphuric acid.

vi) Ascorbic acid in fruits are useful to our bodies. vii) Nitric acids are used in the manufacture of explosives.

Bases and Alkalis

A base is a substance which reacts with an acid to form salt and water only.

Bases are mainly oxides of metals like: copper(II)oxide, CuO; zinc oxide, ZnO;calcium oxide, CaO; magnesium oxide, MgO;or hydroxides of metals and ammonium groups like sodium hydroxide, NaOH; potassium hydroxide, KOH; calcium hydroxide,$Ca(OH)_2$ and ammonium hydroxide, NH_4OH.

Bases that are soluble in water are known as alkalis.

Alkalis

An alkali is a base that dissolves in water to produce hydroxide ions (OH^-) as the only negatively charged ion. A solution of a base in water is called an alkaline solution.

Examples of alkalis include: sodium hydroxide or caustic soda(NaOH), potassium hydroxide or caustic potash (KOH), calcium hydroxide or lime water ($Ca(OH)_2$) and ammonium hydroxide or aqueous ammonia (NH_4OH).

Properties of alkalis

1. Alkalis have a bitter taste
2. Alkalis are soapy and feel slippery 3. Alkalis turn red litmus paper blue
4. Alkalis react with acids to produce salt and water. Acid + Alkali ⟶ Salt + Water
   $NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$
5. Alkalis react with ammonium salts to produce ammonia gas. E.g.
   $NH_4Cl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l) + NH_3(g)$
6. Alkalis precipiatate insoluble metal hydroxides from solutions of their salts. E.g.
   $2NaOH(aq) + Pb(NO_3)_2(aq) \longrightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$

Strength of bases/ alkalis

This is the ease with which a base/ alkali dissociate or ionize. The bases/ alkalis can be categorized as strong or weak.

1. **Strong alkali:** Is one that ionizes completely in aqueous solution. Examples include:

Sodium hydroxide: $NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$

Potassium hydroxide: $KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$

2. **Weak alkali:** This is one that does not ionize completely in aqueous solution. Some molecules remain unionized in the solution. For example:

Ammonium hydroxide: $NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

Neutralization reaction

This is the reaction between an acid and a base to produce salt and water only. The resulting solution is neutral (i.e neither acidic nor alkaline). Examples of neutralization reactions are:

$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

$H_2SO_4(aq) + KOH(aq) \longrightarrow K_2SO_4(aq) + H_2O(l)$

Essentially neutralization takes place when a hydroxide ion (OH^-) reacts with a hydrogen ion (H^+) to form a water molecule (H_2O).

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

INDICATORS (ACID-BASE INDICATORS)

An acid-base indicator is a substance that shows different colors in acidic and alkaline solutions. Indicators are used to establish substances that are acidic, alkaline or neutral. The color change of the indicator depends on the strength of the acid or the base/ alkaline.

Some examples of indicators used in chemistry experiments include: Litmus paper; Methyl orange; Bromothymol blue; Bromothymol red; Phenolphthalein; Universal indicators e.t.c.

Color changes of some indicators in acidic, alkaline and neutral solutions are shown below

Plant extract as a simple acid-base indicators

Preparation of indicators from flowers

Procedure

Collect a handful of flowers with brightly colored petals such as hibiscus and morning glory.

NB Use only one type of flowers and do not mix them

- Remove the petals from the flowers, put them in a motor and crush them carefully
- Add a little ethanol to the crushed petals and stir with a pestle
- Carefully decant the mixture and keep the colored solution in a test tube. This acts as an indicator. Take note of the color of the indicator

Universal indicator

A universal indicator is a mixture of other simple indicators and under goes a range of color changes in solution of different P^H. A universal indicator shows the extent of alkalinity or acidity (i.e in addition to showing whether a solution is acidic or alkaline, it shows whether the acid is weak or strong.) The indicator may be available in solution or paper form.

Indicator	Color in acidic solution	Color in neutral solution	Color in alkaline solution
Litmus	Red	Purple	Blue
Methyl orange	Pink	Orange	Yellow
Phenolphthalein	Colorless	Colorless	Pink

P^H and P^H Scale

The PH of a solution is a number which shows the acidic or basic strength of the solution.

It can also be defined as the negative logarithm (to base ten) of hydrogen ion concentration. It is obtained by finding the negative logarithm of hydrogen ion concentration.

A PH scale measures the strength of an acid or a base/alkali. The PH scale runs from Zero (O) to fourteen (14).

The following are values on the PH scale for solutions of different PH.

- An acidic solution has a PH less than 7

- A neutral solutions has a PH of 7
- An alkaline solution has a PH greater than 7

Representation on a PH scale

PH 0 1 2 3	4 5 6	7	8 9 10	11 12 13 14
Strong acid	Weak acid	Neutral	Weak alkali	Strong alkali

← Increasing acidity   Increasing alkalinity →

Color changes on a universal indicator PH scale

PH 0 1 2 3	4 5	6	7	8 9	10 11	12 13 14
Red	Orange	Yellow	Green	Blue	Indigo	Violet

To find the PH of a solution using a solution of a universal indicator.

1. Add some drops of the universal indicator to the solution being tested.
2. Match the color change of the solution to the corresponding color on the universal indicator PH chart.
3. Record the color change and the PH value.

CHEMICAL FAMILIES

Elements in the same group are referred to as chemical families because of the similarities in their chemical and physical properties. These elements in the same group have similar chemical properties because they have the same number of electrons in their outer most energy levels. The common chemical families are; alkaline metals; alkaline earth metals; halogens and noble gases.

Alkaline metals

These are elements of group I in the periodic table. These metals are very reactive and are kept under oil in the laboratory where they have no contact with water and air. They include lithium (Li), sodium (Na) and potassium (K). Atoms of all these elements each have a single electron on the outer most shell, which is easily lost during chemical reactions leaving a single positively charged ion. i.e.

$\mathrm{Li} \longrightarrow \quad + e-$

$\mathrm{Na} \longrightarrow \quad + e-$

$\mathrm{K} \longrightarrow \quad + e-$

Physical properties

1. The metals are soft and can be cut with a knife. The softness decreases down the group.
2. Their melting and boiling points are low. The melting point decreases down the group.`
3. They have low densities and can float on water.
4. They have shinny surfaces when freshly cut. The shinny surfaces soon tarnish due to reaction with oxygen to form the oxides.
5. They are good conductors of electricity and heat.

6. They do not posses tensile strength to any appreciable extent.

Chemical properties

Reactivity of the alkaline metals increases down the group due to increase in atomic size.(i.e potassium is the most reactive followed by sodium and lithium is the least reactive).

1. Reaction with water

All the metals react with cold water to produce hydrogen gas and alkaline solution of metal hydroxides.

Potassium reacts very vigorously, darts on water and burns with a bright purple (lilac) flame producing fumes of hydrogen gas and an alkaline solution of potassium hydroxide.

$$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$$

Sodium melts into a silvery ball, darts on water reacting vigorously with it producing fumes of hydrogen gas with a hissing sound and an alkaline solution of sodium hydroxide. The reducing in size as the reaction proceeds until when it finally disappears.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Lithium reacts slowly with cold water liberating hydrogen gas.

$$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$$

2. Reaction with air

When the metals are heated in oxygen, they burn readily forming oxides of the metal. Potassium burns with a bright purple flame producing white solids of potassium oxide.

$$4K(s) + O_2(g) \longrightarrow 2K_2O(s)$$

Sodium burns in air with a bright yellow flame forming yellow solids of sodium peroxide in plenty of air. In limited in air, sodium forms white solids of sodium oxide.

$4Na(s) + 2O_2(s) \longrightarrow 2Na_2O_2(s)$ (In excess air)
$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$ (In limited air)

Lithium burns with a red flame forming white residue of lithium oxide.

$4Li(s) + O_2(g) \longrightarrow$
$2Li_2O(s)$

3. Reaction with chlorine

The alkaline metals react with chlorine to form metal chloride salts. Eg. Sodium continuous to burn in chlorine to form white fumes of sodium chloride.

$Na(s) + Cl_2(g) \longrightarrow NaCl(s)$
Potassium burns in chlorine to form fumes of potassium chloride.

$K(s)+Cl_2(g) \longrightarrow KCl(s)$

4. Reaction with acids

Alkaline metals react explosively with acids to liberate hydrogen gas and forming a corresponding salt. E.g. sodium reacts with hydrochloric acid to form sodium chloride and hydrogen gas.

$2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$

Potassium reacts with sulphuric acid to form potassium sulphate and hydrogen gas.

$2K(s) + H_2SO_4(a) \longrightarrow K_2SO_4(aq) + H_2(g)$

Alkaline earth metals

These are elements in group II of the periodic table. They include beryllium, magnesium and calcim. These metals have two electrons on their outer most shell, they lose the two electrons during reactions to form a dipositive ion (an ion with a charge of $^{+}2$).

$Mg \longrightarrow Mg^{2+} + 2e^-$

$Ca \longrightarrow Ca^{2+} + 2e^-$

$Be \longrightarrow Be^{2+} + 2e^-$

Physical properties

1. They re harder than group I metals.
2. They are good conductors of electricity and heat.
3. They are silvery grey in color when freshly cut. However the shiny surface slowly tarnishes on exposure to air forming the metal oxide.
4. Their melting and boiling points are higher than those of group I. This is because these metals release more electrons into the electron cloud forming a stronger metallic bond in the metal structure. This explains why they are also harder than group I metals.

Chemical reactions

Alkaline earth metals are more reactive than alkaline metals. This is because the valence electrons in alkaline earth metals are held more strongly due to increased effective nuclear charge making them not easily to be released. Reactivity of the metals increase down the group due to increase in atomic size. (calcium is the most reactive followed by magnesium then beryllium)

1. Reaction with air

The metals burn in air with their characteristic flame forming oxides of the metal. Magnesium burns in air with a brilliant white flame forming white ash of magnesium oxide.

$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

Calcium burns in air with a bright orange flame forming white ash of calcium oxide.

$Ca(s)+O_2(g) \longrightarrow 2CaO(s)$

Both metals also form some nitrides especially if the supply of oxygen is limited.

$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$ (Magnesium nitride)

$3Ca(s) + N_2(g) \longrightarrow Ca_3N_2(s)$ (Calcium nitride)

When the nitrides are dissolved in water and the solution warmed, ammonia gas is liberated and an alkaline solution is also formed.

$Mg_3N_2(s) + 6H_2O(l) \longrightarrow 3Mg(OH)_2(aq) + 2NH_3(g)$

$Ca_3N_2(s) + 6H_2O(l) \longrightarrow 3Ca(OH)_2(aq) + 2NH_3(g)$

2. Reaction with water

Beryllium does not react with water. Magnesium piece reacts very slowly with cold water producing small bubbles of hydrogen gas and an alkaline solution of magnesium hydroxide.

$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$

Magnesium reacts very rapidly with steam to produce white ash of magnesium oxide and hydrogen gas.

$2Mg(s) + 2H_2O(g) \longrightarrow 2MgO(s) + 2H_2(g)$

Calcium reacts steadily with water evolving hydrogen gas and calcium hydroxide. $Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$

3. Reaction with chlorine

The metals burn readily in chlorine to form white anhydrous chloride salts. Magnesium forms magnesium chloride and calcium forms calcium chloride.

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

$Ca(s) + Cl_2(g) \longrightarrow CaCl_2(s)$

4. Reaction with acids

Calcium reacts vigorously with acids forming a salt and hydrogen gas. Magnesium reacts steadily with acids forming a salt and hydrogen gas. E.g. magnesium reacts with sulphuric acid forming magnesium sulphate and hydrogen gas.

$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$

Halogens

These are group (VII) elements; they are chlorine, bromine, iodine and fluorine. Halogens exist as diatomic molecules. This is because they lack one electron to completely fill their outer most energy level hence their atoms share electrons in order to completely fill the outer most energy level. These elements react by gaining one electron each forming negatively charged ions.

$F + e^- \longrightarrow F^-$

$Cl + \longrightarrow Cl^-$

$Br + \longrightarrow Br^-$

$I +$ $\quad I^-$

Physical properties

1. They are colored (chlorine is greenish yellow gas; fluorine is pale yellow gas; bromine is a brown volatile liquid and iodine is a shinny black solid).

2. Down the group, physical state changes from gas to liquid to solid. (fluorine and chlorine are gases at room temperature; bromine is a liquid and iodine is in solid form, the shinny black iodine solids sublime into a purple vapor on slight heating).
3. Solubility in water decreases down the group. Fluorine, chlorine and bromine are soluble and iodine is slightly soluble in water.
4. They have low melting and boiling points because the intermolecular forces of attraction are weak. The melting and boiling point increase down the group as the weak vanderwaal's forces of attraction increase with increase in atomic size.

Chemical properties

The order of reactivity decreases down the group.(i.e. fluorine is the most reactive and iodine in the least reactive). This is because these elements react by gaining electrons and down the group, the atomic radius increases there fore, an in coming electron is less readily attracted by the nucleus. There fore, the smaller the atom, the more readily and strongly does the nucleus attract an electron to the outer most shell and the more reactive that atom becomes.

1. Reaction with water

All the halogens dissolve in water forming a mixture of acids. Chlorine dissolves in water forming a mixture of hypochlorous acid(chloric (I) acid)(HOCl) and hydrochoric acid(HCl).

$$Cl_2(g) + H_2O(l) \longrightarrow HOCl(aq) + HCl(aq)$$

This solution turns blue litmus paper red then bleaches it. The bleaching action is due to the presence of hypochlorous acid(HOCl) that readily gives up its oxygen to the dye.

A solution of bromine in water is weakly acidic and weakly bleaching due to formation of weak hydrobromic acid (HBr) and bromic(I)acid(HOBr) respectively.

$$Br_2(g) + H_2O(l) \longrightarrow HOBr(aq) + HBr(aq)$$

Chlorine is slightly soluble in water forming very weak acids. A solution of chlorine is too weak to bleach dyes, therefore, bleaching power of the elements decrease down the group.

2. Reaction with alkalis

Halogen react with alkalis to form a mixture of salts and water. E.g. chlorine is absorbed by a solution of sodium hydroxide forming a pale yellow solution of a mixture of sodium chloride(NaCl) and sodium hypochlorite(NaOCl)

$$Cl_2(g) + 2NaOH(aq) \longrightarrow NaCl(aq) + NaOCl(aq) + H_2O(l)$$

In hot concentrated solution of an alkali (e.g. NaOH), sodium hypochlorite is not formed but instead sodium hypochlorate is formed($NaOCl_3$).

$3Cl_2(g) + 6NaOH(aq) \longrightarrow 5NaCl(aq) + NaOCl_3(aq) + 3H_2O(l)$

3. Reaction with metals

Metals continue to burn in halogens forming salts. E.g. sodium continues to burn in a gas jar of chlorine to form dense white fumes which settle as white solids of sodium chloride. A similar reaction occurs with bromine and iodine.

$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$
$2Na(s) + Br_2(g) \longrightarrow 2NaBr(s)$

4. Displacement reaction of halogens

When chlorine is bubbled through a colorless solution of potassium bromide, the solution gradually change to red-brown as bromine is displaced from its solution by chlorine.

$2KBr(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + Br_2(g)$

If chlorine is bubbled through a colorless solution of potassium iodide,the solution turns dark brown as iodine is displaced from its solution.

$2KI(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + I_2(g)$

Chlorine being the most reactive displaces all the halogens from their solutions.

Noble Gases (Group (VIII) Elements)

These elements are referred to as inert elements because they are non reactive. The elements include; helium,neon,argon,krypton,xenon and radon. Characteristics of these elements include:

1. Their outer most energy levels are completely filled with electrons making them very stable elements and chemically non reactive. However in spite of their stable electronic configurations, a few compounds have been formed from xenon(e.g. Xenon(II)fluoride,XeF_2) and krypton(E.g. krypton(II)fluoride, KrF_2)
2. The noble gases exist as monatomic gases.
3. They have very low melting and boiling points due to negligible forces of attraction between the atoms.
4. Noble gases are colorless

Though nobles gases are non reactive, they are used for a variety of purposes including the following:

- Argon is filled in light bulbs to prevent the tungsten fillament from reacting with air.
- Argon provides an inert environment for gas-liquid chromatography, for risky welding jobs and for some chemical reactions.
- Neon is used extensively in advertising signs and in lasers.
- Helium is used to inflate air ships, weather balloons and aeroplane tyres.
- A mixture of oxygen and helium is used by divers instead of air.

Sample questions on Acids, Bases, Indicators and Chemical Families

Acids, bases and indicators

1. *Write short notes on: acids, bases, alkalis, indicators and basicity. Illustrate with examples and equations.*
2. *Outline three properties typical of acids and three properties typical of alkalis. (illustrate with equations). Differentiate between: a strong acid and a weak acid; an alkali and a base.*
3. *Describe how you can use a named material from the environment to make an indicator in the laboratory.*
4. *What is meant by neutralization reaction? Illustrate by use of equations.*

Chemical families

1. Describe with equations, the reactions of

a) Alkaline metals with air, water, acids and chlorine.
b) Halogen with water and alkalis.
c) Alkaline earth metals with water.

2. Discuss the displacement reaction of halogens.
3. Mention three examples of noble gases. Outline four uses of noble gases.

WATER AND HYDROGEN

WATER

Water is a chemical compound consisting of oxygen and hydrogen. Its chemical formula is H_2O. **Sources of water**

1. Rain water

The water originates from rain and it is relatively pure with no dissolved minerals. However, it contains suspended materials and dissolved gases such as carbon dioxide and sulphur dioxide. Rain water is good for washing as it forms lather easily with soap and does not contain dissolved salts. It is also good for drinking since it has a flat taste.

2. Wells and spring water

The water comes from underground and is obtained by drilling the ground inform of a bore hole or, the water just springs out of the ground. It is actually rain water that has drained into the ground and collected into a pool of underground water. It has less suspended materials but contain dissolved mineral salts making it good for drinking. It may not be good for washing because of the dissolved mineral salts that delay the formation of lather.

3. River and lake water

The water is obtained from rivers and lakes. It contains both suspended materials and dissolved minerals. As water flows in to rivers and lakes, it gathers floating materials and pollutants such as sewage and industrial wastes, so it is not good for washing and drinking.

4. Sea and ocean water

These provide large sources of water. The water contains a lot of suspended materials and dissolved substances. The mineral salts accumulate settle at the bottom of the water and the water may become saturated with the salt. This is why the water tastes salty.

5. Ice caps

The water is obtained from melting ice and it is relatively pure. The ice formed at the peaks of mountains melts and the water runs down forming spings.

Uses of water

1. Biological use

All biological processes in living organisms require water as a medium for the reactions. For example, photosynthesis, excretion, transport of materials and transmission of impulses.

2. Domestic use

Drinking, cooking, washing, bathing, mopping, construction, recreation e.t.c. **3. Farm use**

Irrigation, for dipping animals, for drinking by animals, for fish farming e.t.c.

4. Industrial use

Water is used as a solvent in many industries, for cleaning, for cooling purposes especially in power stations, used in steam engines, generating hydro electricity, transportation using boats, ships e.t.c, for recreation purposes and construction.

Water cycle

Water cycle is the process by which water circulates round the earth. The driving force for the cycle is the sun.

The water from the water bodies like seas, lakes, rivers e.t.c. evaporates into the atmosphere due to the sun heat. The vapor cools in the atmosphere and then condenses to form clouds which fall as rain. The rain then flows back into the water bodies.

N.B. Water vapor in the atmosphere may also result from transpiration processes in plants and respiration processes of all living organisms and burning of starchy food materials forming carbon dioxide and water which are expelled in to the atmosphere.

Water pollution

This is the discharge of unwanted materials/ substances into water bodies. The un wanted materials are generally called pollutants and are harmful to living organisms.

Examples include; acids, bases, insecticides, fertilizers, sewage, solid particles, oil e.t.c.

Sources of water pollutants

1. Atmospheric gases like carbondioxide, sulphur dioxide, nitrogen dioxide and hydrogen sulphide. These gases dissolve in the rain water making it acidic.
2. Dissolved maerials from rocks and soil.
3. Industrial discharge (effluent)
4. Fertilizers washed from farm lands by running water.

5. Insecticides and herbicides resulting from spraying from homes and farms.
6. Soaps, detergents and other chemicals from homes, salons e.t.c.

Water treatment

This is the removal of harmful substances from water making it suitable for use (domestic use. In Uganda, the treatment of water is done by Uganda National Water and Sewage Co operation (UNWSC). Processes involved in water treatment are:

1. Addition of chemicals

Chemicals are added to water to remove harmful substances and neutralize acids and bases. The chemicals as well remove hardness of water, precipitate metals and cause suspended solid particles to settle.

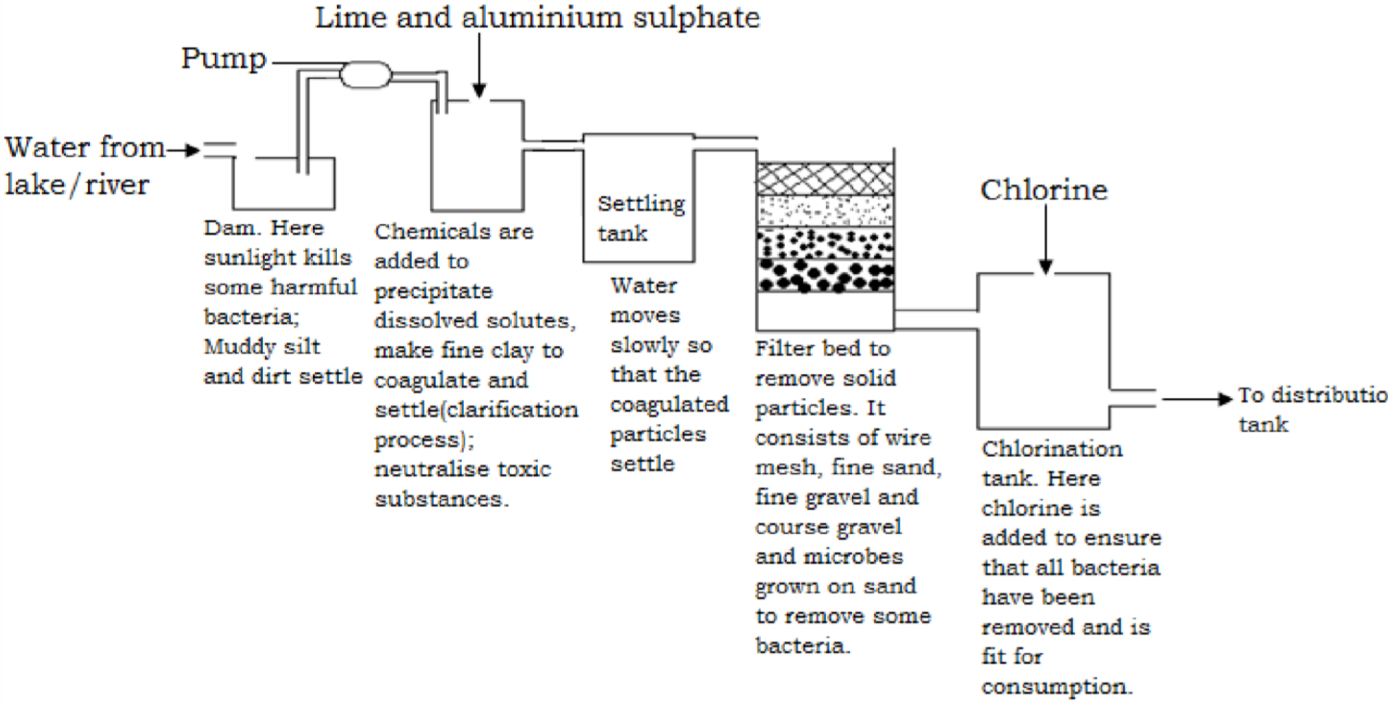
2. Filtration

Filtration is carried out to remove the solid particles. Water is made to pass through wire mesh screens to remove large and floating objects and then through sand and gravel to filter out smaller solid particles. The last part of the filter bed consists of sand with some microbes grown to remove some bacteria that might be harmful.

3. Chlorination and fluoridation

Chlorine or fluorine is added to kill the harmful organisms such as bacteria. Only a small concentration of chlorine is used because it is harmful. However, in swimming pools, high concentration of chlorine is used as the water may contain high bacteria concentration is not meant for drinking. In some countries, ozone is used instead of chlorine.

Summary of water purification process



Sewage

Sewage is running water containing wastes from toilets, bathrooms, sinks, factories and streets of towns.

Sewage treatment

The reason for treating sewage is to remove harmful pollutants before allowing the sewage to join water bodies. The following are the processes involved in sewage treatment:

1. Separation of solid and liquid wastes through sieving and sedimentation.
2. Addition of certain bacteria to the liquid wastes to convert harmful materials to harmless substances.
3. Addition of chlorine to kill the harmful bacteria
4. The liquid waste (effluent) from the treatment plant is the discharged into the water body. The solid waste from the treatment of sewage is the called the **sludge** and has a number of uses.

Uses of sludge

- It is used as fertilizer because it is rich in nitrogen and phosphorus.
- It is used as a raw material to produce bio gas.
- It is used for road surfacing since it forms hard solids.
- It is used in land reclamation to fill quarries.

Properties of water

Physical properties

1. Pure water is a clear, colorless and tasteless liquid.
2. It boils at 100˚C and melts at 0˚C.
3. It has a density of 1g/cm^3 at 4˚C.
4. Pure water is neutral to litmus i.e. has no effect on litmus.
5. Water expands on freezing.
6. It is a universal solvent.

Test for water

Water is chemically tested for using anhydrous copper (II) sulphate of cobalt chloride paper.

When water is added to anhydrous copper(II)sulphate, it changes color from white to blue. Or when water is added to cobalt chloride paper, it changes color from blue to pink.

Chemical properties

1. Reaction with metals

Water reacts with some metals and the vigour depends on the position of the metal in the electrochemical series.

Potassium
Sodium reacts with cold water

a) Potassium

When a piece of potassium is placed into a trough containing water, it reacts very vigorously, catches fire and burns with a purple flame; it darts on the water surface producing a hissing sound and white fumes of hydrogen gas. It explodes with a pop sound as it disappears. The resultant solution turns red litmus paper to blue.

$$K(s) + H_2O(l) \longrightarrow KOH(aq) + H_2(g)$$

b) Sodium

Sodium reacts vigorously with water. When a piece of sodium is placed in cold water, it melts into a silvery ball and darts on the surface of water as it produces a hissing sound with evolution of a colorless hydrogen gas. The sodium gradually becomes smaller and smaller as it reacts with water to produce an alkaline solution of sodium hydroxide.

N.B. Sodium burns with a yellow flame if its movement on water is restricted.

$$Na(s) + H_2O(l) \longrightarrow NaOH(l) + H_2(g)$$

c) Calcium

Calcium reacts steadily with water. When a small piece of calcium is dropped is water, it sinks to the bottom producing streams of bubbles of a colorless gas (hydrogen gas) and an alkaline solution that turns red litmus paper blue. The water turns milky as the calcium hydroxide is slightly soluble in water.

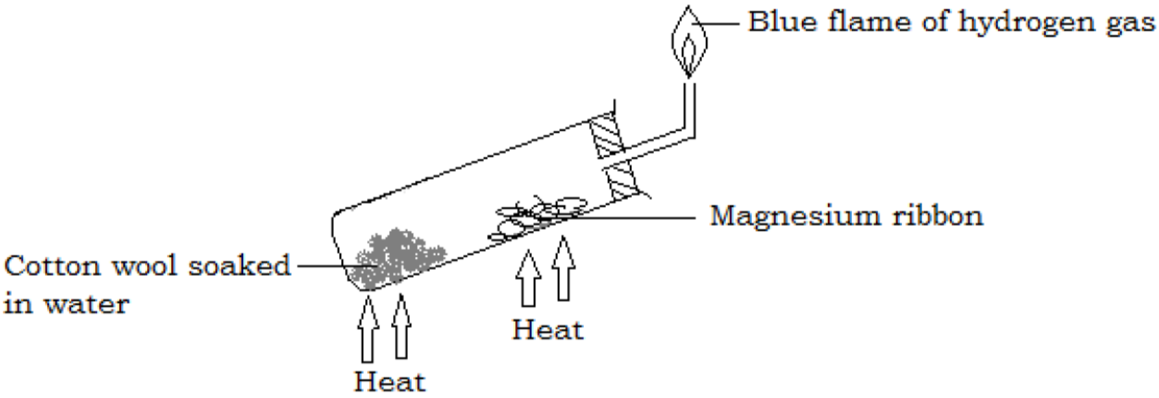
$$Ca(a) + H_2O(l) \longrightarrow 2Ca(OH)_2(aq) + H_2(g)$$

d) Magnesium

Magnesium sinks to the bottom and reacts very slowly with cold water to produce very few bubbles of hydrogen gas on the surface of the metal and an alkaline solution of magnesium hydroxide.

$$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

However, hot magnesium reacts rapidly with steam, burning with a bright white flame leaving white ash of magnesium oxide. The hydrogen gas produced burns with a blue flame in air.



$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$

Note

1. The magnesium metal used is first cleaned to remove the oxide coating due to corrosion.
2. The cotton wool soaked in water is heated to release steam that reacts with the heated magnesium.
3. The boiling tube is first heated to drive out all air (oxygen).

e) Iron and zinc

Zinc and iron do not react with cold water but react with steam when heated to red hot producing their respective oxides and hydrogen gas. Zinc reacts faster than iron.

$Zn(s) + H_2O(g) \longrightarrow ZnO(s) + H_2(g)$

Yellow when hot
White when cold

$3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$

Triiron tetra oxide
(Blue-black)

Aluminium does not react with water or steam due to the formation of an insoluble oxide layer which prevents contact between water and the metal inhibiting any further reaction.

All the other metals below iron do not react with either cold water or even steam.

2. **Reaction with oxides**

Water reacts with metallic oxides (basic oxides) to form alkaline solutions. E.g.

$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$

$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$

With non metallic oxides (acidic oxides), water forms acidic solutions. E.g.

$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$

$\longrightarrow$

$SO_3(g) + H_2O(l) \qquad\qquad\qquad\qquad H_2SO_4(aq)$

$CO_2(g) + H_2O(l) \quad = \quad H_2CO_3(aq)$

HARDNESS OF WATER

Water can be classified as hard or soft.

Soft water is one which forms lather readily with soap. Examples of soft water include:

rain water, distilled water, deionised water.

Hard water

This is water that does not form lather readily with soap. Examples of hard water include: tap water, sea water, borehole, well water, spring water.

Causes of hardness of water

Hardness of water is caused by the presence of the calcium ion (Ca^+) and magnesium ion (Mg^+). These ions reach water when the soluble salts of calcium and magnesium are dissolved by rain water and carried to water bodies. The compounds that cause hardness of water include; calcium hydrogen carbonate ($Ca(HCO_3)_2$), magnesium hydrogen carbonate($Mg(HCO_3)_2$),calcium chloride ($CaCl_2$), magnesium chloride ($MgCl_2$), Calcium sulphate ($CaSO_4$), Magnesium sulphate ($MgSO_4$).

Effects of magnesium and calcium ions on soap

Soap is a sodium salt of a long chain carboxylic acid. It is known as sodium stearate (Na-st). Soap reacts with magnesium or calcium ions to form dirty white solids called **scum**. So, when hard water is used for washing, a lot of soap is used because part of the soap initially reacts with calcium and magnesium ions to form scum.

$$CaSO_4(aq) + 2Na\text{-}st(aq) \longrightarrow Na_2SO_4(aq) + Ca\text{-}st_2(s)$$

Types of hard water

There are two types of hard water i.e. permanent and temporary hard water.

Temporary hard water

This is a type of hard water whose hardness can be removed by boiling. I.e. the water can be made soft by boiling.

Causes: it is caused by the presence of dissolved calcium hydrogen carbonate or magnesium hydrogen carbonate.

Permanent hardness

Is a type of hard water whose hardness cannot be removed by boiling. I.e. the hard water cannot become soft on boiling.

Causes: it is caused by calcium sulphate, calcium chloride, magnesium sulphate and magnesium chloride.

Distinguishing between permanent hard water, temporary hard water and soft water Procedure

- Place equal volumes of each unboiled water samples in three different conical flasks.
- Run soap solution from the burette into each sample while shaking the conical flask until when lather forms.
- Note the volume of soap solution required to form lather with each of the unboiled water samples.

Results

The soft water required the least volume of soap solution to form lather compared to the samples of hard water.

Repeat the above procedure with boiled water sample.

Results

Soft water required the same volume of soap solution to form lather as the unboiled soft water.

Temporary hard water after boiling required less volume of soap solution to form lather than the unboiled one. This is because; the water was made soft by boiling and thus formed lather readily.

Permanent hard water even after boiling required the same volume of soap solution as the unboiled to form lather. This is because even after boiling, the water still remained hard and took time to form lather.

Methods of softening water

Removal of only temporary hardness of water

1. Boiling (Physical method)

Boiling decomposes calcium and magnesium hydrogen carbonates to their respective carbonates, water and carbondioxide gas. The calcium and magnesium ions are removed from the water as insoluble calcium and magnesium carbonates. The calcium and magnesium carbonates may be deposited as solids at the bottom of the boiler and are referred to as boiler scales of kettle fur.

$$Ca(HCO_3)_2(aq) \xrightarrow{\text{(boiling)}} CaCO_3(s) + CO_2(g) + H_2O(l)$$

Disadvantages of boiler scale/fur

- The fur produced is an insulator hence a lot of heat is required for boiling and this wastes a lot of fuel or electricity.
- The fur produced may also block pipes carrying boiled water and possibly causing it to burst.

By chemical means the fur can be removed from he kettle by adding mineral acids like hydrochloric acid and sulphuric acid.

$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

2. Addition of calcium hydroxide (lime water) (chemical method)

Calcium hydroxide reacts with the soluble magnesium or calcium hydrogen carbonates to form insoluble magnesium and calcium carbonates.

$Ca(HCO_3)_2(aq) + Ca(OH)_2(aq) \longrightarrow 2CaCO_3(s) + 2H_2O(l)$

The disadvantage of this method is that when excess lime water is used, the water becomes hard again.

It has an advantage in that it is cheaper than boiling as calcium hydroxide is readily available.

3. Addition of aqueous ammonia (ammonium hydroxide solution) (chemical method)

During the process, the soluble hydrogen carbonates are converted into insoluble carbonates. This enables magnesium and calcium ions to be removed from the water.

$2NH_4OH(aq) + Mg(HCO_3)_2(aq) \longrightarrow MgCO_3(s) + (NH_4)_2CO_3(aq) + H_2O(aq)$

Methods for softening both temporary and permanent hardness

4. Distillation (physical method)

During distillation, the soft water is collected as the distillate and the dissolved calcium and magnesium ions remain in the distillation flask as residue.

The advantage of this method is that the water obtained is actually pure. However, it can only be used in small scale because it is expensive.

5. Addition of sodium carbonate solution (washing soda) (chemical method)

Sodium carbonate converts magnesium or calcium ions into insoluble carbonates.

$Na_2CO_3(aq) + Ca(HCO_3)_2(aq) \longrightarrow 2NaHCO_3(aq)$
$+ CaCO_3(s)$ The method is cheap and easy to carryout on a large scale.

6. Ion exchange method (permutit) (chemical method)

The calcium and magnesium ions in hard water can be exchanged with sodium ions using a suitable ion exchange material. This makes the water to become soft. The common ion exchange materials are zeolites and permutit. Zeolites are natural occurring forms of sodium aluminium silicate and permutit is the artificial form. Zeolite/permutit can be represented by the formula Na_2Y.

$CaCl_2(aq) + Na_2Y(aq) \longrightarrow CaY(s) + 2NaCl(aq)$
$MgSO_4(aq) + Na_2Y(aq) \longrightarrow MgY(s) + 2Na_2SO_4(aq)$
$Ca(HCO_3)_2(aq) + Na_2Y(aq) \longrightarrow CaY(s) + NaHCO_3(aq)$

The method works on the principle of exchange of ions. During the process, calcium and magnesium ions are removed from water and they appear as precipitates of aluminium silicate.

The advantage here is that, the permutit can be recovered and used several times.

Advantages of hard water

1. Calcium compounds in water are important in our diet for strong teeth and bones.
2. Magnesium in hard water is important for photosynthesis since it is used for the formation of chlorophyll.
3. Calcium in hard water is responsible for the formation of strong shells of many animals e.g. snails and egg shells.
4. Hard water tastes better and it is used in the brewing industry.
5. Hard water reduces lead poisoning and it is the only type of water transported by lead pipes.

Disadvantages of hard water

1. It wastes a lot f soap, as more of the soap reacts the calcium and magnesium ions to form scum. There for lather does not form readily.
2. When hard water ids used for washing, it leaves dirty marks on cloth due to formation of scum.
3. Hard water forms kettle fur and boiler scales which are insulators, there fore a lot of fuel or electricity is used during boiling of water.
4. The boiler scales may block the pipes carrying boiling water hence causing the pipe to burst.

Exercise

1. Explain why soap does not form lather readily at first with hard water and then eventually forms the lather?
2. Explain what is meant by hardness of water. What are the causes of temporary and permanent hardness of water?. How do these compounds that cause hardness of water reach the water?
3. A white precipitate is formed when

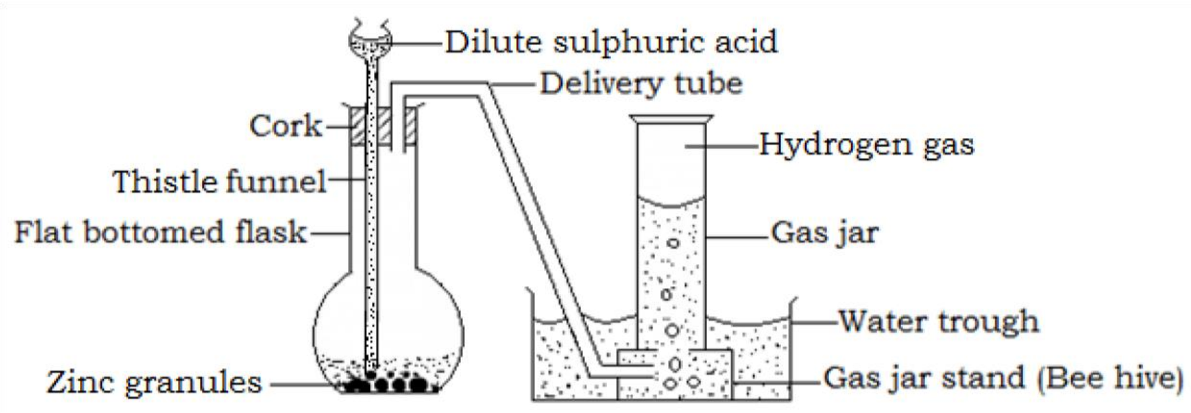
i) Carbondioxide is blown into lime water ii) Temporary hard water is boiled iii) Washing soda is added to water containing dissolved magnesium chloride.

HYDROGEN

Hydrogen is the smallest element and the lightest gas. Hydrogen usually does not occur is Free State but in combined states as water, acid, hydrocarbons and other organic compounds.

Laboratory preparation of hydrogen

Hydrogen is prepared in the laboratory by the action of either dilute hydrochloric acid or dilute sulphuric acid on zinc granules or zinc metal. (zinc granule are used instead of pure zinc because some impurities in the zinc granules act as catalysts). **Set up**



Place some zinc granules in the flask and add to it a little copper (II) sulphate solution. The copper (II) sulphate acts as a catalyst in the preparation of hydrogen gas. Arrange the apparatus as shown above and add dilute sulphuric acid/ hydrochloric acid to the zinc granules through the funnel.

Effervescence occurs as hydrogen gas is produced. The gas is then collected over water. However, if the gas is required dry, it is passed through a wash bottle containing concentrated sulphuric acid and collected by upward delivery method or the gas is passed through a U-tube containing fused calcium chloride to dry the gas. **Equation**

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Test for hydrogen

When a burning splint is brought into a gas jar of hydrogen, the splint will be extinguished with a ―pop‖ sound.

Properties of hydrogen

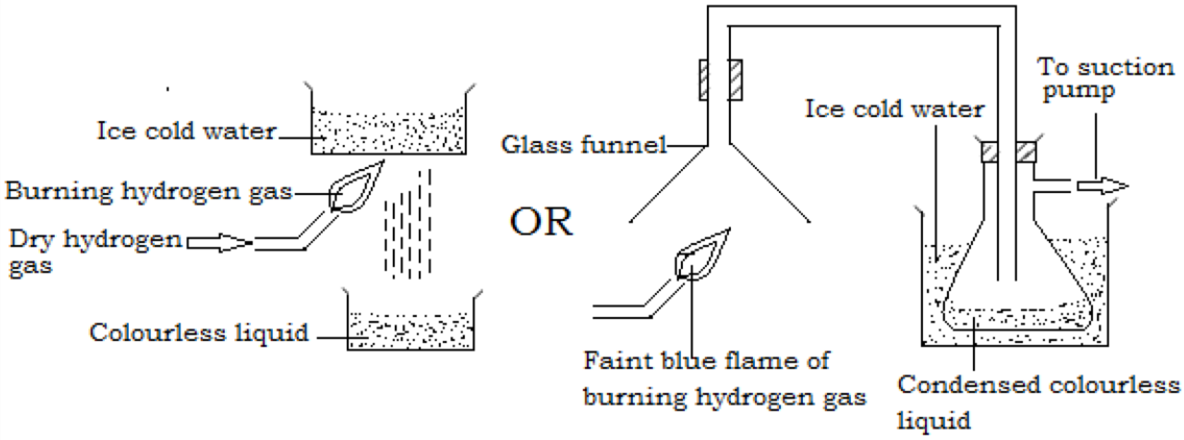
a) Physical properties

- It is the lightest gas known (lighter than air)
- It is colorless, odourless and tasteless
- It is slightly soluble in water
- It is a neutral gas (has no effects on indicators) **b) Chemical properties**

1. Combustion (burning) of hydrogen

Hydrogen burns in air with a faint blue flame to produce water vapor. The gaseous product can be condensed in a cool environment to form a colorless liquid the turns white anhydrous copper (II) sulphate to blue indicating that it is water.

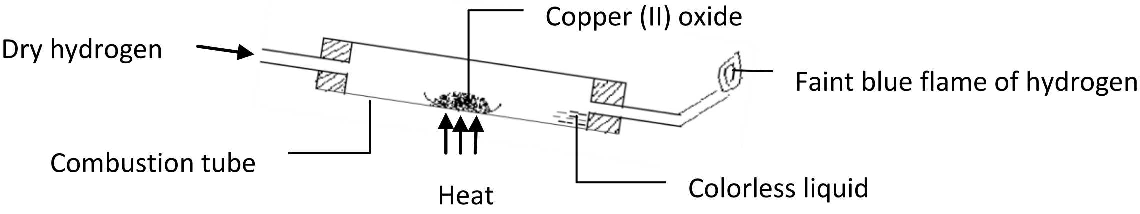
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$



2. Reduction action of hydrogen

Hydrogen is a reducing agent. It removes oxygen from the oxides of some metals (less reactive metals) like lead and copper forming the metal and the hydrogen gas itself is oxidized to water.

Reduction of copper(II)oxide



When dry hydrogen gas is passed over heated copper (II) oxide in a combustion tube, a red glow spreads through the copper (II) oxide and the oxide turns from black to brown as the oxide is reduced to copper metal (brown in color). The hydrogen itself is oxidized to water which collects as a colorless liquid.

$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$

Hydrogen also reduces lead (II) oxide and triirontetraoxide to lead and iron respectively.

Precautions

1. Dry hydrogen gas must be used.
2. The combustion tube must be in a slanting position so as to prevent the water formed from running back to the hot part of the tube.
3. Hydrogen should be passed through the combustion tube to expel out air before heating the oxide.
4. Hydrogen should be passed through the tube for some times after heating .This prevents re oxidation of the copper.
5. The excess hydrogen should be burnt to prevent explosion with air.

Reduction is the addition of hydrogen to a substance or the removal of oxygen from a substance.

Oxidation is the removal of hydrogen from a substance or the addition of oxygen to a substance.

3. Reaction with chlorine

A mixture of chlorine and hydrogen appears not to react at room temperature, but when exposed to sun light or heated, the mixture explodes forming misty fumes of hydrogen chloride gas.

$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

4. Reaction with reactive metals

Hydrogen reacts with highly reactive metals to form hydrides. For example, sodium reacts with hydrogen to form sodium hydride.

$Na(s) + \frac{1}{2}H_2(g) \longrightarrow NaH(s)$

Uses of hydrogen

1. Hydrogen is a very light gas and therefore used to fill balloons.
2. Hydrogen with oxygen form oxy-hydrogen flame which is very hot and used for welding and cutting metals.
3. Hydrogen is used in the manufacture of ammonia by Haber process.
4. Hydrogenation of vegetable oil makes it hard and used for making margarine and cooling fats like blue band and kimbo.
5. It is used in the manufacture of hydrogen bombs.
6. Hydrogen is also used as fuel for rockets.

Sample questions on water and hydrogen

1. *Briefly describe the water cycle. Mention the major sources of water. How is waater beneficial to you and to the community from where you come?*
2. *What is water pollution? What are the main sources of water pollution? How do the pollutants mainly reach the water bodies?. Describe briefly the effects of water pollution to living organisms in water bodies.*
3. *In an attempt to make water available for domestic use, water is treated. Explain in detail the steps taken in water purification processes.*
4. *Describe the reactions of sodium, potassium, magnesium and calcium with water stating clearly the conditions under which the reactions take place.*
5. *What is meant by soft and hard water? Give examples of each. Explain the types of hard water that you know giving the causes of each; outline the physical and chemical methods of softening each type of hardness. (Where appropriate use equations to illustrate). Mention the advantages and disadvantages of hard water.*
6. *Describe the aid of a labeled drawing how hydrogen gas is prepared in the laboratory. How do you confirm that the gas produced in your description is hydrogen gas?*

7. *How do you show that water is an oxide of hydrogen? Describe the reduction reaction of hydrogen using of copper(II) oxide (What are the precautions for the reaction). Outline at least five uses of hydrogen gas.*

SALTS

A salt is a compound formed when either all of part of the ionisable hydrogen of the acid is replaced by a metallic ion or ammonium ion. Or

A salt is an ionic compound consisting of a positive metallic or ammonium ion and a negative ion derived from an acid.

Salts get their names from the acids they are derived from. Examples are in the table below.

Acid	Salt	Example
Hydrochloric acid (HCl)	Chlorides	Sodium chloride ($NaCl$)
Sulphuric acid (H_2SO_4)	Sulphates	Ammonium sulphate $(NH_4)_2S0_4$
Carbonic acid (H_2CO_3)	Carbonates	Calcium hydrogen carbonate, $Ca(HCO_3)_2$
Nitric acid (HNO_3)	Nitrates	Barium nitrate,$Ba(NO_3)_2$
Phosphoric acid (H_3PO_4)	Phosphates	Sodium phosphate,Na_3PO_4
Ethanoic acid (CH_3COOH)	Ethanoates	Sodium ethanoate,CH_3COONa
Sulphurous acid (H_2SO_3)	Sulphites	Sodium sulphite. Na_2SO_3

Types of salts a) Normal salt

This is a salt produced when all the ionisable hydrogen of the acid is replaced by a metallic or ammonium ion. These salts do not contain ionisable hydrogen. Examples include; sodium chloride, $NaCl$; ammonium nitrate, NH_4NO_3; Magnesium sulphate, $MgSO_4$; lead (II) bromide, $MgBr_2$ and sodium phosphate, Na_3PO_4.

All normal salts have PH of 7 except salts formed from

i) Strong bases and weak acids e.g. sodium carbonate (Na_2CO_3) and potassium ehanoate (CH_3COOK). These salts in solution have PH value more than 7.

ii) Strong acids and weak bases e.g. ammonium chloride (NH_4Cl). The salts have a PH value less than 7 in solution.

b) Acid salts

An acid salt is a salt formed when only part of the ionisable hydrogen of the acid is replaced by ammonium or metallic ion. These salts contain ionisable hydrogen and examples include: sodium hydrogensulphate, $NaHSO_4$; calcium hydrogencarbonate, $Ca(HCO_3)_2$; and potassium hydrogen carbonate, $KHCO_3$.

These acid salts behave like salts because they contain metallic ion and a negative ion derived from an acid; they behave like acids because the negative ions are capable of further ionization to yield hydrogen ion(H^+) (i.e $HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$).

c) Basic salts

Basic salt are formed when insufficient acid is present to neutralize the available base. E.g. basic zinc chloride (Zn(OH)Cl) and basic magnesium chloride (Mg(OH)Cl).

All monobasic acids form normal salts while dibasic and tribasic acids form both normal and acid salts.

Naming of salts

Salts are named by adding the name of the radical or ion of the acid after the name of the metal or ammonium. Examples are:

Name of metal/ammonium radical	**Name of acid radical (negative ion)**	**Name of salt**
Sodium	Chloride	Sodium chloride, NaCl
Aluminium	Sulphate	Aluminium sulphate, $Al_2(SO_4)_3$
Ammonium	Nitrate	Ammonium nitrate NH_4NO_3
Potassium	Carbonate	Potassium carbonate, K_2CO_3

SOLUBILITY OF SALTS

Solubility is the amount of solute in grams required to saturate 100g of solvent (water) at a particular temperature.

A salt is described as soluble if it can dissolve in a given solvent and insoluble if it cannot dissolve in the solvent. Salts have varying degree of solubility in water as described below:

All ammonium, sodium, and potassium salts are soluble in water. All nitrate salts are soluble in water.

All chloride salts are soluble in water except silver chloride, lead (II) chloride (sparingly soluble) and mercury (I) chloride.

All sulphate salts are soluble in water except lead (II) sulphate and barium sulphate. Calcium sulphate is sparingly soluble in water.

All carbonate salts are insoluble in water except sodium, potassium and ammonium carbonates.

Determining the solubility of a salt e.g. sodium chloride Procedure

- Take about $50cm^3$ of distilled water in a beaker.
- Add sodium chloride crystals to the water a little at a time while stirring continuously until when no more salt dissolves. The solution formed is saturated.

- Weigh a clean evaporating dish and pour into it a little of the clear salt solution.
- Weigh the evaporating dish with the salt solution and evaporate the solution to dryness carefully through a water bath.
- Allow the evaporating dish to cool and reweigh the dish with the dry salt.

Results

Mass of empty dish= a g
Mass of dish + saturated solution = b g
Mass of dish + dry salt = c g

Calculations

Mass of saturated solution = (b-a) g
Mass of dry salt = (c-a) g
Mass of water (solvent) = (b-a)-(c-a)
$= b-\cancel{a}-c+\cancel{a}$
= (b-c) g

(b-c) g of water dissolves (c-a) g
of NaCl 1 g of solvent dissolves
$\frac{(c-a)}{(b-c)}$ g of NaCl

100g of solvent dissolves $\frac{(c-a)}{(b-c)}$ x 100 g of NaCl

There fore, the solubility of NaCl at room temperature is $\frac{(c-a)}{(b-c)}$ x 100 g/100 g of water.

Examples

1. 30g of sodium chloride crystals were dissolved in 75 g of water at 80°C. Calculate the solubility of sodium chloride at this temperature.

Solution

75 g of water dissolves 30 g of NaCl

1 g of solvent dissolves $\frac{30}{75}$ g of NaCl
100g of solvent dissolves ($\frac{30}{75}$ x 100) g of NaCl
=40g

There fore, the solubility of NaCl at 80°C temperature is 40g/100 g of water.

2. 12.0g of potassium chlorate was carefully evaporated to dryness, 2.4 g of potassium chlorate crystals were left on the evaporating dish. Calculate the solubility of potassium chlorate in grams per 100g of water at room temperature.

Solution

Mass of solvent = mass of solution–mass of solute

=(12.0-2.4)g = 9.6g of solvent (water)

9.6 g of water dissolves 2.4g of potassium chlorate
1 g of solvent dissolves $\frac{2.4}{9.6}$ g of potassium chlorate
100g of solvent dissolves ($\frac{2.4}{9.6}$ x 100) g of potassium chlorate
=25g
There fore, the solubility of potassium chlorate at room temperature is 25g/100 g of water.

3. 75g of a saturated solution contains 30g of salt. Calculate,
   i) The solubility of the salt
   ii) The percentage of salt in the saturated solution

Solution

i) Mass of solvent = masss of solution- mass of solute
=(75-30)g = 45g of solvent

45 g of solvent is saturated by 30g of salt

1 g of solvent is saturated by $\frac{30}{45}$ g of salt
100g of solvent saturated by ($\frac{30}{45}$ x 100) g of salt
=66.67g
There fore, the solubility of the salt at room temperature is 66.67g/100 g of water.

ii) Percentage of salt = $\frac{mass\ of\ salt}{mass\ of\ solution}$ x 100
= $\frac{30}{75}$ x 100
=40%

Exercise

1. In an experiment to determine the solubility of potassium nitrate at 20˚C, the following results were obtained.
   Mass of evaporating dish + saturated solution = 100.7g
   Mass of evaporating dish = 65.3g
   Mass of dish + dry salt = 73.8 g
   Use the data above to calculate the solubility of potassium nitrate at 20˚C. Clearly show your working.

Factors that affect the rate of solubility of salts

1. **Amount of solvent**

Solubility of most salts increase with increase in the amount of solvent used.

2. **Nature of solvent/solute**

Solubility of a salt may increase or decrease depending on the nature of solvent or solute.

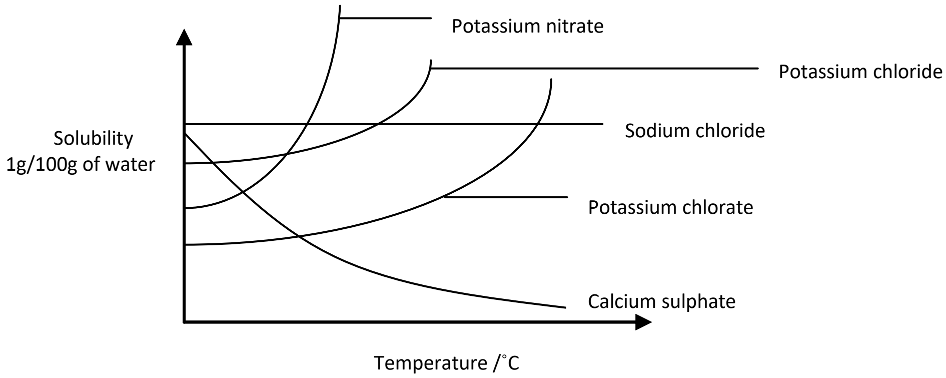
3. **Temperature**

Solubility of most salts increase with increase in temperature. For example, potassium chlorate and potassium nitrate. Solubility of a few salts like calcium chloride and calcium sulphate decrease with increase in temperature. The solubility of sodium hydroxide and gases as well also decrease with increase in temperature.

Solubility curve

A solubility curve is a graph that shows how the solubility of a salt varies with temperature. The graph is obtained by plotting solubility (on the vertical axis) against temperature (on the horizontal axis).

Solubility curve of some common salts



The solubility of potassium chloride, potassium nitrate and potassium chlorate increase with increase in temperature. The solubility of potassium nitrate increases most rapidly, followed by potassium chlorate then potassium chloride.

The solubility of sodium chloride increases very slightly with increase in temperature. The solubility of calcium sulphate decreases with increase in temperature.

Uses of solubility curves

1. It can be used to find the solubility of a salt at a given temperature.
2. It gives the temperature at which a given amount of salt saturates 100g of solvent.
3. It can be used to explain the trend of solubility of salts.
4. A solubility curve can be used to calculate the mass of salt obtained by cooling a solution from a higher temperature to a lower temperature.

Mass of salt= (solubility at a higher temperature – solubility at a lower temperature)

For example, if a salt P with solubility of 180g/100g of water at 90˚C was cooled to a temperature of 30˚C where its solubility is 25g/100g of water. Calculate the mass of salt formed formed after cooling the solution.

Solution

Mass of salt= (solubility at a higher temperature – solubility at a lower temperature)

= (180-45)g

= 155g

Application of solubility

1. Solubility is used to separate soluble salts from a mixture by fractional crystallization.
2. It is used in the extraction of salts from large water bodies like lakes and seas.

Exercise

1. a) Describe an experiment that you would carryout to determine the solubility of potassium nitrate at 15˚C.

   b) Determine the solubility in water of substance S at room temperature from the following data.

Mass of evaporating basin	25g
Mass of evaporating basin + Saturated solution of S	55g
Mass of evaporating basin + Solid S	30g

2. a) Define the term solubility?

   b) The table below shows the solubility (ies) of salt P in water at different temperatures.

Temperature/˚C	10	20	30	40	50	60
Solubility, g/100g of water	18	20	24	30	38	50

i) plot a graph of solubility of P against temperature ii) use your graph to determine

a) solubilities of P at 25˚C and 45˚C

b) the mass of crystals deposited when a solution of P is cooled from 50˚C to 25˚C

iii) Calculate the mass of P that would dissolve in 45g of water at 25˚C

PREPARATION OF SALTS

The method of salt preparation depends on whether the salt is soluble in water or not. Soluble salts are prepared by **crystallization** and **neutralization**. Insoluble salts are prepared by **precipitation** or **double decomposition.** Other salts are prepared by direct synthesis.

Preparation of soluble salts

Soluble salts are prepared using dilute acids and metals, metals oxides, metal hydroxides and metal carbonates.

General procedure

1. Place some dilute acid in a beaker
2. Warm the acid and add the metal, metal oxide, metal hydroxide and metal carbonate bit by bit until in excess to ensure that the acid is completely used up.
3. Filter the excess metal, metal oxide, metal hydroxide or metal carbonate and collect the filtrate.

4. Saturate the filtrate by evaporating and allow the solution to cool as it cools to form the salt crystals.
5. Filter the crystals and wash them with water.
6. Dry the crystals in an oven, or under sun shine or between filter papers.

1. Preparation of salts from metals and dilute acids

Salts prepared by this method are soluble salts of iron, magnesium, aluminium and zinc. (I.e. metals higher than lead and lower than calcium in the reactivity series.)

N.B. Nitrates cannot be prepared using this method because dilute nitric acid being an oxidizing agent, does not react with metal to liberate hydrogen gas.

Example

Laboratory preparation of zinc sulphate crystals from zinc metal/powder

- Put dilute sulphuric acid in a beaker and heat it gently until when it‘s hot.
- Add zinc powder to the hot acid bit by bit while stirring until when the zinc powder is in excess.
- Filter off the excess zinc powder to obtain zinc sulphate solution as the filtrate.
- Saturate the filtrate by evaporating.
- Allow it to cool and form crystals of the salt.
- Filter the crystals and wash them with distilled water.
- Dry the crystals either in qaan oven of under the sun or between filter papers.

Equation

$H_2SO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + H_2(g)$

Other salts formed in similar ways are:

Zinc chloride

Hydrochloric acid + zinc $\longrightarrow$ Zinc chloride + Hydrogen gas

$HCl(aq) + Zn(s) \longrightarrow ZnCl_2(aq)$

Iron (II) sulphate

Sulphuric acid + Iron $\longrightarrow$ Iron (II) sulphate + Hydrogen gas $H_2SO_4(aq) + Fe(s) \longrightarrow FeSO_4(aq) + H_2(g)$

2. Preparation of salts from metal oxides and dilute acids

Example

Preparation of copper (II) sulphate from copper (II) oxide in the laboratory - Put dilute sulphuric acid in a beaker and heat it gently until when it‘s hot.

- Add copper (II) oxide to the hot acid bit by bit while stirring until when the copper (II) oxide is in excess.

- Filter off the excess copper (II) oxide to obtain copper sulphate solution as the filtrate.
- Saturate the filtrate by evaporating.
- Allow it to cool and form crystals of the salt.
- Filter the crystals and wash them with distilled water.
- Dry the crystals either in an oven or under sunshine or between filter papers. **Equation**

$H_2SO_4(aq) + CuO(s) \longrightarrow CuSO_4(aq) + H_2O(g)$

Other examples of salts formed from metal oxides are:
Magnesium chloride
Hydrochloric acid + Magnesium oxide ⟶ Magnesium chloride + Water
$2HCl(aq) + MgO(s) \longrightarrow MgCl_2(aq) + H_2O(l)$

Copper (II) nitrate
Copper (II) oxide + Nitric acid ⟶ Copper (II) nitrate + Water
$CuO(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + H_2O(g)$

Aluminium sulphate
$H_2SO_4(aq) + Al_2O_3(s) \longrightarrow Al_2(SO_4)_3(aq) + H_2O(g)$

3. Preparation of salts from insoluble metal carbonates Example

Preparation of lead (II) nitrate from lead (II) carbonate - Pour dilute nitric acid in a beaker and warm it gently.
- Add lead (II) carbonate a little at a time. Effervescence occurs as carbondioixde is evolved.
- Continue adding the carbonate until when it is in excess and no more effervescence occurs.
- Filter off the excess carbonate to get a colourless filtrate.
- Evaporates the filtrate by heating gently to obtain a saturated solution.
- Cool the saturated solution to obtain white crystals of lead (II) nitrate salts.
- Wash the crystals with cold distilled water and dry either on sun shine, in an oven or between filter papers.

Equation

Lead (II) nitrate
Lead (II) carbonate + Nitric acid ⟶ Lead (II) nitrate + Water + carbondioxide
$HNO_3(aq) + PbCO_3(s) \longrightarrow Pb(NO_3)_2(aq) + H_2O(g) + CO_2(g)$

Other salts prepared in similar ways are:
Barium chloride
Hydrochloric acid + Barium carbonate ⟶ Barium chloride + Water + carbondioxide
$2HCl(aq) + BaCO_3(s) \longrightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$

Copper (II) sulphate

Sulphuric acid + copper (II) carbonate → Copper (II) sulphate + water + carbondioxide $H_2SO_4(aq) + CuCO_3(s) \longrightarrow CuSO_4(aq) + H_2O(g) + CO_2(g)$

4. Preparation of salts from metal hydroxides

Example

Preparation of lead (II) nitrate starting from lead (II) hydroxide - Pour dilute nitric acid in a beaker and warm it gently.

- Add lead (II) hydroxide a little at a time while until when it is in excess. - Filter off the excess hydroxide to get a colorless filtrate.
- Evaporates the filtrate by heating gently to obtain a saturated solution.
- Cool the saturated solution to obtain white crystals of lead (II) nitrate salts.
- Wash the crystals with cold distilled water and dry them either on sun shine, in an oven or between filter papers.
  -*

Equation

$Pb(OH)_2(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + 2H_2O(l)$

Other salts prepared in similar ways are:

Iron (II) nitrate

Iron (II) hydroxide + Nitric acid → Iron (II) nitrate + Water

$Fe(OH)_2(s) + HNO_3(aq) \longrightarrow Fe(NO_3)_2(aq) + H_2O(g)$

Copper (II) chloride

Hydrochloric acid + Copper (II) hydroxide → Copper chloride + Hydrogen gas

$2HCl(aq) + Cu(OH)_2(s) \longrightarrow CuCl_2(aq) + 2H_2O(l)$

Zinc sulphate

Zinc hydroxide + sulphuric acid → zinc sulphate + water

$Zn(OH)_2(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + 2H_2O(l)$

Laboratory preparations of salts whose carbonates, oxides and hydroxides sare insoluble

These salts include potassium, sodium and ammonium salts. The salts can be prepared by titration method (neutralization).

Neutralization is a reaction between an acid and a base to produce a salt and water only.

General procedure

- Put a known volume of hydroxide of a metal in a conical flask.
- Add 2 or 3 drops of an indicator.
- Run a suitable acid from the burette until when the color of the mixture just changes. Note and record the volume of acid used.
- Measure accurately the same volume of hydroxide as before and titrate with exactly the same volume of acid as recorded above.

NB. An indicator is not used in the second titration, since the volume of acid required to neutralize the fixed volume of base was already got.

- Stir and heat the solution to make it saturated.
- Allow the hot saturated solution to cool as it forms salt crystals.
- The crystals are filtered off, washed with cold distilled water and dried in an oven, under sun shine or between filter papers.

Preparation of sodium chloride crystals in the laboratory

Procedure

- Put a known volume of sodium hydroxide in a conical flask.
- Add 2 or 3 drops of an indicator.
- Titrate the sodium hydroxide with hydrochloric acid from the burette until when the end point is reached (when the indicator changes color). Note and record the volume of acid used.
- Measure accurately the same volume of sodium hydroxide as before and titrate with exactly the same volume of hydrochloric acid as recorded above without using an indicator.
- Stir and heat the solution to make it saturated.
- Allow the hot saturated solution to cool as it forms salt crystals.
- The crystals are filtered off, washed with cold distilled water and dried in an oven, under sun shine or between filter papers.

Equation

$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

Other salts prepared in similar ways are:

Ammonium sulphate ($(NH_4)_2SO_4$)

Ammonium hydroxide + sulphuric acid ⟶ ammonium sulphate

$2NH_4OH(aq) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq) + H_2O(l)$

Potassium nitrate, KNO_3

Potassium hydroxide + nitric acid ⟶ potassium nitrate + water

$KOH(aq) + HNO_3(aq) \longrightarrow KNO_3(aq) + H_2O(l)$

Sodium sulphate

Sodium hydroxide + sulphuric acid ⟶ sodium sulphate + water

$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

Preparation of insoluble salts

Insoluble salts are prepared by **double decomposition** or **precipitation** method. In this method, two soluble salts are mixed to form two new salts by exchange of radicals. One of the new salts formed is a soluble salt and one is an insoluble salt that appear as precipitates. The precipitate is filtered off and washed then dried.

Precipitation is the formation of solids when solutions are mixed.

A precipitate is the solid formed when two or more solutions are mixed.

Example

Preparation of barium sulphate (by reacting barium nitrate and sodium sulphate) Procedure

- Put a solution of barium nitrate in a beaker and add a solution of sodium sulphate to it. A white precipitate of barium sulphate immediately appears, - Filter off the precipitate and wash with distilled water.
- Dry the precipitate (salt formed) under sun shine, in an oven or between filter papers.

Equation

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + NaNO_3(aq)$

This method can be used to prepare salts such as lead sulphate, aluminium chloride, silver chloride, silver carbonate and barium sulphate.

Preparation of lead (II) sulphate (by reacting lead (II) nitrate and sulphuric acid)

Procedure

- Add dilute nitric acid to lead (II) nitrate solution in a beaker and stir the mixture. White precipitates of lead (II) sulphate is formed.
- Filter off the precipitates and wash with distilled water to remove traces of the acid.
- Dry the precipitates in a steam oven or leave it to dry in air.

NB. If any of the compounds to be used in the preparation of the salt is insoluble in water, it must first be made to dissolve in a mineral acid. For example, in the preparation of lead (II) sulphate using lead (II) oxide, the lead (II) is first divvolved in nitric acid to form lead (II) nitrate.

Equation

$PbO(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l)$

The lead (II) nitrate formed reacts with sulphuric acid to form lead (II) sulphate. $Pb(NO_3)_2(aq) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + HNO_3(aq)$

Preparation of salts by direct synthesis

Salts consisting of two elements (binary salts) can be prepared by direct synthesis/ direct combination.

Example:

In the preparation of sodium chloride from sodium and chlorine, burning sodium is lowered in a gas jar of chlorine. Sodium continues to burn in chlorine forming white fumes which settle into white solids (sodium chloride).

Equation

$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

Other salts prepared by direct synthesis include:

Magnesium chloride

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

Potassium chloride

$2K(s)+Cl_2(g) \longrightarrow 2KCl(s)$

$\longrightarrow$

$\longrightarrow$

Iron (III) chloride

$2Fe(s) + 3Cl_2(g) \qquad 2FeCl_3(s)$

Zinc bromide

$Zn(s)+Br_2(g) \qquad ZnBr_2(s)$

Effects of heat on salts

1. Carbonates

Potassium and sodium carbonates are very stable and are not decomposed by heat. But if the salts are hydrated, they lose their water of crystallization. In such a process, salts lose their crystalline nature and become amorphous.

$Na_2CO_3.10H_2O(s) \longrightarrow Na_2CO_3(s) + 10H_2O(g)$

(Crystalline solid) (Powdery solid)

All the other metallic carbonates decompose upon heating to give the oxide of the metal and carbondioxide gas.

Example

When white zinc carbonate is heated, it produces a colorless gas that turns lime water milky leaving a yellow residue when hot which turns white on cooling.

$ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$

(white solids) (yellow solids Colorless gas that
turns lime when hot but water milky
turns white on cooling)

When lead (II) carbonate is heated, a brown residue (when hot) which becomes yellow on cooling and a colorless gas that turns lime water milky are produced.

$PbCO_3(s) \longrightarrow PbO(s) + CO_2(g)$

(white solids) (Brown solids Colorless gas that turns lime
when hot but water milky
turns cooling) yellow on

When copper (II) carbonate is heated, black solids of copper (II) oxide is formed and a colorless gas that turns lime water milky evolved.

$CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$

(white solids) (black solids) Colorless gas that
turns lime water milky

Ammonium carbonate decomposes to give ammonia gas, carbondioxide and water vapor.

$(NH_4)_2CO_3(s) \longrightarrow 2NH_3(g) + H_2O(g) + CO_2(g)$

Hydrogen carbonate of metals decompose to form carbonate of metals, carbondioxide gas and water vapor.

Example

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

2. Sulphates

Sulphates of sodium and potassium donot decompose on heating. When hydrated sulphates of potassium or sodium is heated, it loses its water of crystallization and becomes amorphous.

$$K_2SO_4.10H_20(s) \longrightarrow K_2SO_4(s) + 10H_2O(g)$$

Sulphates of heavy metals decompose to give metal oxides and white fumes of sulphur trioxide gas. When heated more strongly, the sulphur trioxide gas decomposes to give sulphur dioxide and oxygen gas.

Examples

When hydrated copper (II) sulphate crystals are heated, they lose their water of crystallization and changes from blue crystals to white powder. The water condenses as a colorless liquid on the cooler parts of the test tube.

$$CuSO_4.5H_2O(s) \longrightarrow CuSO_4(s) + 5H_2O(g)$$

(blue crystals) (white powder)

On further heating,the white powder gives off white fumes of a gas (SO_3) and a black residue (CuO) is left.

$$2CuSO_4(s) \longrightarrow 2CuO(s) + SO_3(g)$$

(white powder)
(black solids)

When green solid of iron (II) sulphate is heated, it loses it water of crystallization and changes from green to dirty-yellow anhydrous solids.

$$FeSO_4.7H_2O(s) \longrightarrow FeSO_4(s) + 7H_2O(g)$$

(Green solids) (Dirty yellow solids)

When heated more strongly, it gives off sulphur dioxide (a colorless gas that turns potassium dichromate solution from yellow to green), white fumes of sulphur trioxide and brown residue of iron (III) oxide is left.

$$2FeSO_4(s) \longrightarrow Fe_2O_3(s) + SO_3(g) + SO_2(g)$$

(dirty yellow) (brown solids)

3. Nitrates

All nitrates decompose upon heating.

Sodium and potassium nitrates melt into colorless liquids then decompose upon heating to give their corresponding nitrites that form yellow solids on cooling and oxygen gas.

$$2NaNO_3(s) \longrightarrow 2NaNO_2(s) + O_2(g)$$

(white crystals) (yellow powder)

All nitrates from calcium down to copper decompose to give their corresponding oxides, brown fumes of nitrogen dioxide gas and oxygen gas.

Examples

$$2Zn(NO_3)_2(s) \longrightarrow 2ZnO(s) + 4NO_2(g) + O_2(g)$$

(White crystals) (Yellow-hot; white-cold) (Brown fumes)

$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$
(White crystals) (Brown-hot; yellow-cold) (Brown fumes)

N.B. Lead(II)nitrate decomposes with a cracking sound.

$2Cu(NO_3)_2(s) \longrightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$
(Green crystals) (Black powder) (Brown fumes)

Mercury (II)nitrate and silver nitrate decompose to give their corresponding metals, brown fumes of nitrogen dioxide gas and a colorless gas that relights a glowing spling (oxygen gas).

$2AgNO_3(s) \longrightarrow Ag(s) + 2NO_2(g) + O_2(g)$
(White crystals) (Grey solids) (Brown fumes)
$Hg(NO_3)_2(s) \longrightarrow Hg(l) + 2NO_2(g) + O_2(g)$

Ammonium nitrate sublimes upon heating to give dinitrogen oxide and water vapor.

$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)$

4. Chlorides

Metallic chlorides are not affected by heat because hey are very stable. However, if they are hydrated, the lose their water of crystallization.

For example
$MgCl_2.7H_2O(s) \longrightarrow MgCl_2(s) + 7H_2O(g)$

Ammonium chloride sublimes on slight heating and on further heating, it decomposes to give ammonia and hydrogen chloride gases.

Equation
$NH_4Cl(s) \rightleftharpoons NH_4Cl(g) \longrightarrow NH_3(g)$
$+ HCl(g)$

Effects of heat on hydroxides

Hydroxides of sodium and potassium are not decomposed by heat. If they are in solid forms, they absorb moisture and melt to form solutions.

However, the hydroxides of other metals decompose to give the corresponding oxides and water vapor.

Example
$Zn(OH)_2(s) \longrightarrow ZnO(s) + H_2O(g)$
(white crystals) (yellow-hot; white-cold)
$Pb(OH)_2(s) \longrightarrow PbO(s) + H_2O(g)$
(white crystals) (brown-hot; brown-cold)

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Sample questions on salts

1. *Briefly explain what is meant by the following and in each case give an example (formula and name): salt, acid salt, normal salt and basic salt.*
2. *What is solubility of a salt? Describe with the aid of a labeled drawing how you can determine the solubility of sodium chloride at 40 ˚C*

3. *(a) State one method for preparing: (i) lead (II) nitrate other than from lead (II) carbonate (ii)lead (II) sulphate. Describe how a pure dry sample of lead (II) nitrate can be prepared in the laboratory, starting from lead (II) carbonates.*
4. *Describe what would be observed and write equations when each of the following salts are heated: (a) lead(II) carbonate (b) lead(II) nitrate (c) sodium carbonate crystals (d) zinc carbonate (e) ammonium carbonate (f) hydrated copper(II) sulphate (g) sodium nitrate (h) zinc nitrate crystals (i) silver nitrate and ammonium nitrate.*
5. *Describe in detail how you would prepare pure dry samples of (a) zinc sulphate from zinc carbonate (b) lead(II) sulphate from lead(II) nitrate (c) sodium chloride by action of an acid and a base.*
6. *Write down the formula of the precipitate which is produced when pairs of aqueous solutions are mixed as follows:* (i) $Ba(NO_3)_2$ and Na_2SO_4; (ii) H_2S and $Pb(NO_3)_2$; (iii) NH_4Br and $AgNO_3$; (iv) $CaCl_2$ and K_2CO_3.

BONDING AND STRUCTURE

BONDING

Bonding is the chemical combination of atoms or elements to form compounds. The force of attraction holding atoms or elements together in a molecule/crystal is referred to as a **chemical bond**. Chemical bonding /combination occurs mainly in four forms as:

1. **Ionic/electrovalent bonding**-this involves transfer of electrons from a metal atom to a non metal atom. It occurs between metals and non metals.
2. **Covalent bonding**-this involves sharing of electrons between two or more non metal atoms/elements. The atoms/elements involved contribute to the bonding electrons.
3. **Dative/co-ordinate bonding**-this involves sharing of the bonding electrons which are donated by one molecule or atom involved.
4. **Metallic bonding**-this occurs between atoms of metal elements.

Electrovalent/ ionic bonding

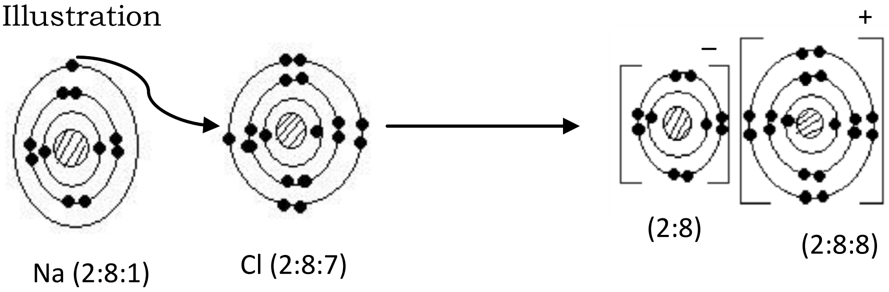
This involves transfer of electrons from a metal to a non metal. The number of electrons lost by the metal atom or gained by non metal atom is equivalent to its valency. The loss of electrons from a metal atom leads to formation of a positively charged ion (cation) and the gain of electrons by a non metal atom leads to the formation of a negatively charged ion (anion). The positively charged ion and the negatively charged ion are attracted to each other and the force of attraction holding them together is known as electrovalent/ionic bond. The compounds formed are referred to as **ionic/ electrovalent compounds**. Metal atoms lose electrons so as to gain stable electronic configuration of noble gases and non metal atoms also gain electrons to become stable.

Formation of some ionic compounds

Sodium chloride

When sodium burns in chlorine to form sodium chloride, the sodium atom loses an electron and forms a positively charged sodium ion(Na^+) while the chlorine atom gains the electron and forms a negatively charged chloride ion (Cl^-).

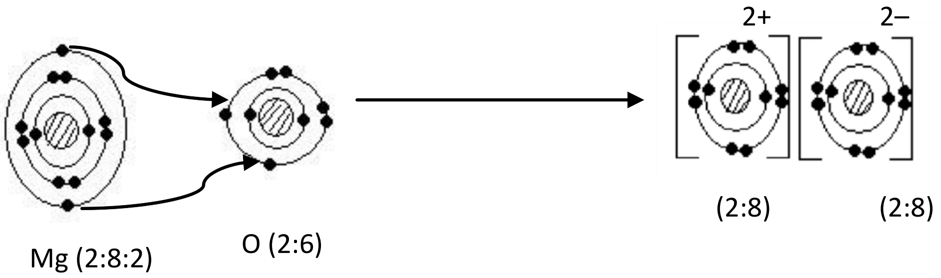
Illustration



The positively charged sodium ion and negatively charged chlorine ion are attracted to each other by a strong force called ionic bond.

Magnesium oxide

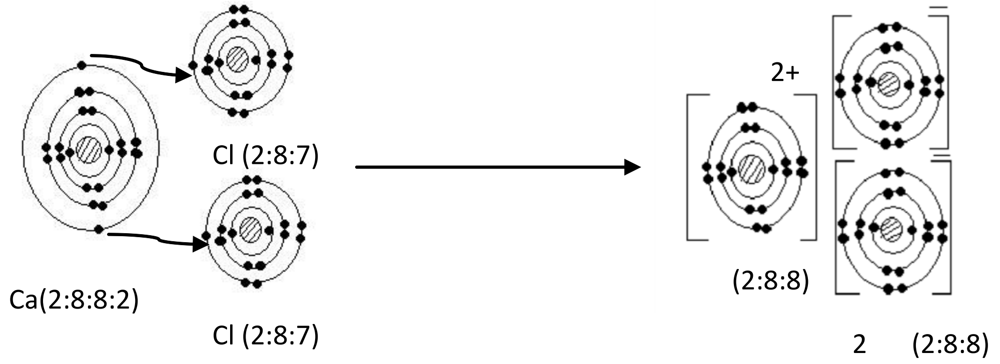
Magnesium atom loses two electrons to the oxygen atom and forms a positively charged magnesium ion (Mg^{2+}) while the oxygen atom forms a negatively charged oxide ion (O^{2-}). These two oppositely charged ions are then attracted to each other. Illustration



Calcium chloride

During the formation of calcium chloride, calcium atom loses two electrons which are gained by two chlorine atoms. This leads to the formation of Ca^{2+} and two Cl^-which are then attracted to each other.

Illustration



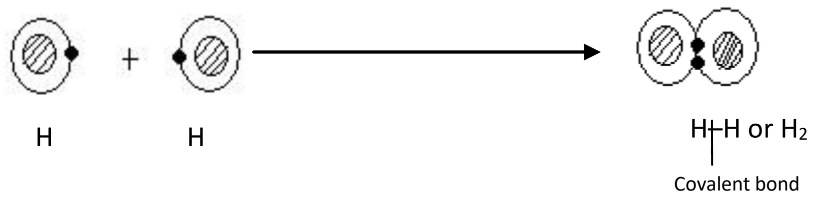
Covalent bonding

This type of bonding occurs between non metal elements/ atoms to form compounds. Covalent bonding involves mutual sharing of electrons between elements in which each of the atoms involved gives equal number of electrons to be shared. On sharing the electrons, the atoms attain stable electronic structure of noble gases. The compound formed are called covalent compounds. Covalent bonds may by classified as single, double, triple or quadruple depending on the number of electrons shared.

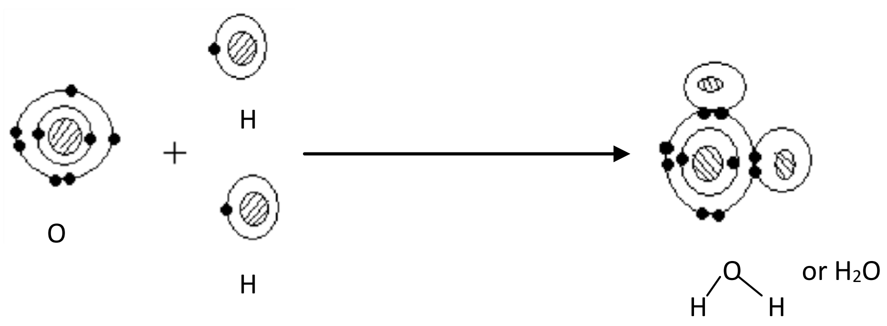
a) **Single covalent bond:** this is formed when one pair of electron is shared. Each of the atoms involved contribute an electron for the bond formation.

Example

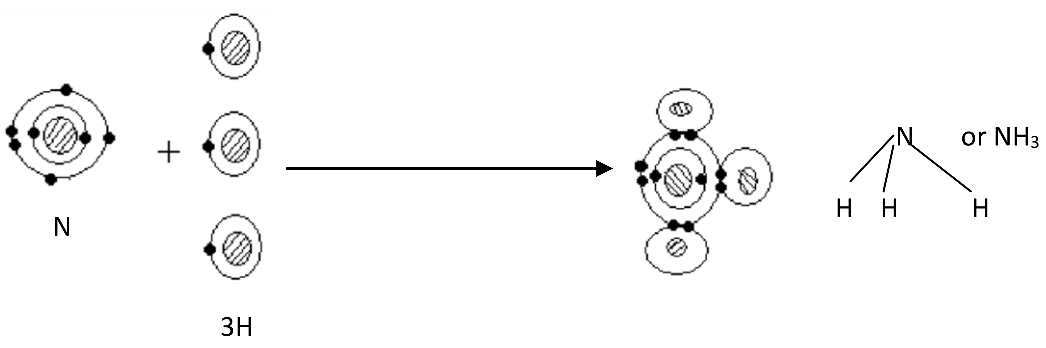
i) Hydrogen molecule (H_2)



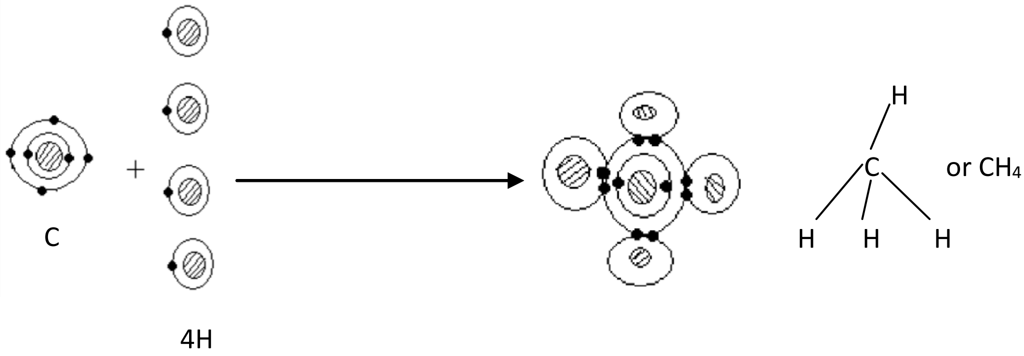
ii) Water molecule (H_2O)



iii) Ammonia (NH_3)



iv) Methane (CH_4)

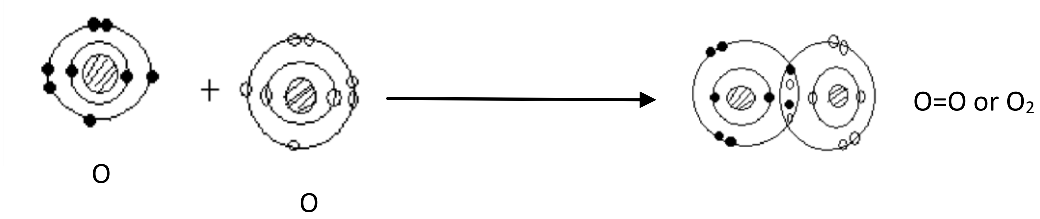


N.B. The pair of electron on the outer most energy level that does not take part in bonding is the **lone pair** or **non bonded pair.**

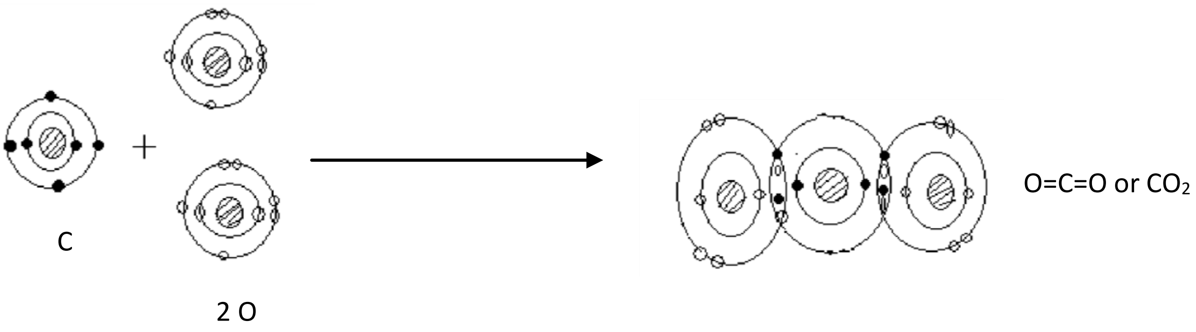
b) Double covalent bond

This is formed when two pairs of electrons are shared between two bonded atoms. Example

i) Oxygen (O_2)



ii) Carbon dioxide (CO_2)

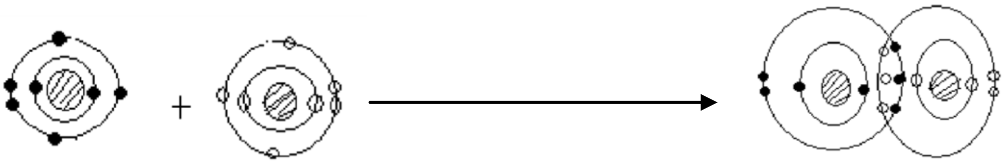


c) Triple covalent bond

This involves sharing of three pairs of electrons between atoms.

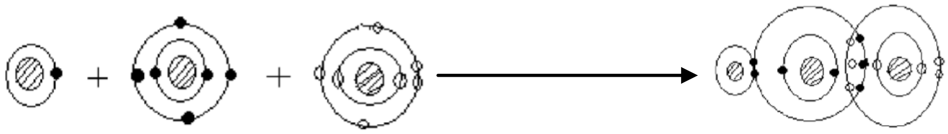
Example

i) Nitrogen molecule (N_2)



$N\equiv N$ or N_2

ii) Hydrogen cyanide (HCN)

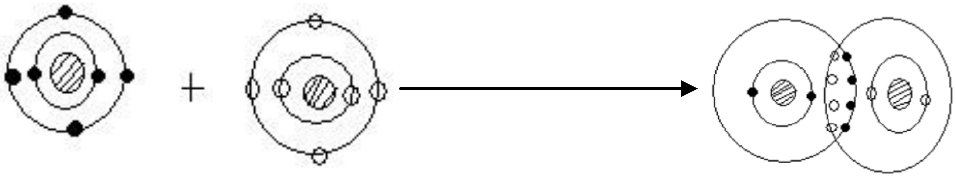


$H-C\equiv N$

d) Quadruple covalent bond

This is formed by sharing four pairs of electrons. It occurs between carbon atoms.

Illustration



$\bar{C}\equiv C$

Dative/ co-ordinate bond

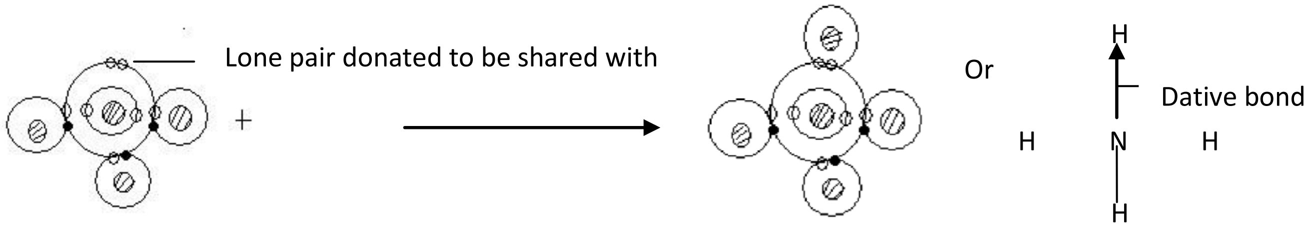
This involves sharing of electrons but the shared pair of electrons is donated by one atom/ molecule. Here, one molecule/ atom donates the pair of electrons to be shared with an ion or another atom. Normally it is atoms/ molecules with lone pair of electrons that form this bond-by donating the lone pair of electrons to be shared with another atom/ion. The bond may be represented by an arrow originating from the donor atom (atom that donates electrons to be shared) to the atom accepting the electrons.

Example

i) Formation of ammonium ion (NH_4^+)

During the formation of ammonium ion, ammonia reacts with an acid. The nitrogen in ammonia has a lone pair of electron which it donates to be shared with the hydrogen ion (proton) from the acid.

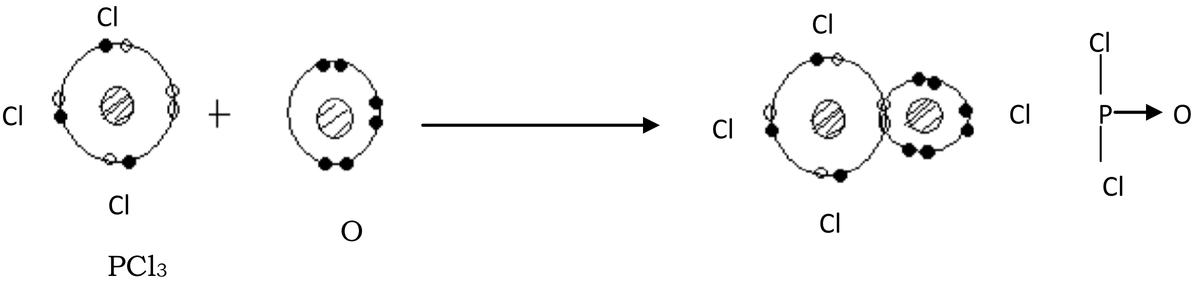
NH_3 + $\longrightarrow$
Illustration



ii) Formation of phosphorus (V) trichloride oxide.

This is formed between phosphorus (III) chloride which donates the lone pair of electron to be shared and oxygen atom.

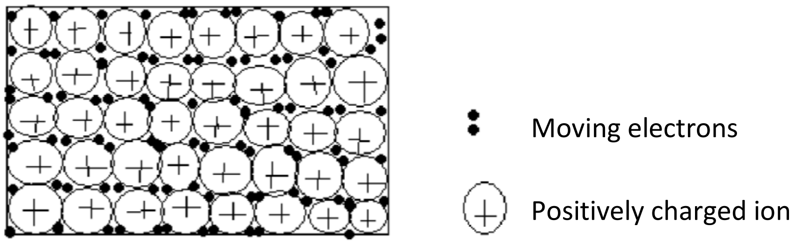
Illustration using outer most shell



Metallic bonding

This is the type of bonding in metals due to the attraction between metal ions and the valency electrons within the structure of the metal (metal lattice). In the structure of a metal, the loosely held valency electrons (electrons in the outer mos energy level) are released in to a general pool and the atoms become positively charged. These electrons move freely around the ions formed and are termed as **mobile/delocalized electrons**. The ions formed and the electrons attract each other forming metallic bond.

The strength of metallic bond increases with increase in the number of electrons released in to the electron cloud. Thus the bond is very strong in metals like iron and aluminum that release up to three electrons each to the electron pool and weak in metals like sodium and potassium that only release one electron each to the electron pool. Illustration



Exercise

1. Draw diagrams to show the arrangement of electrons in the following compounds

a) KCl b) Na_2O c) $MgCl_2$ d) CS_2

2. Draw diagrams to show the arrangement of electrons in the energy levels of atoms in the following molecules. Use only the outer most electrons to illustrate. a) Ethane (C_2H_4) b) Ethyne (C_2H_2)

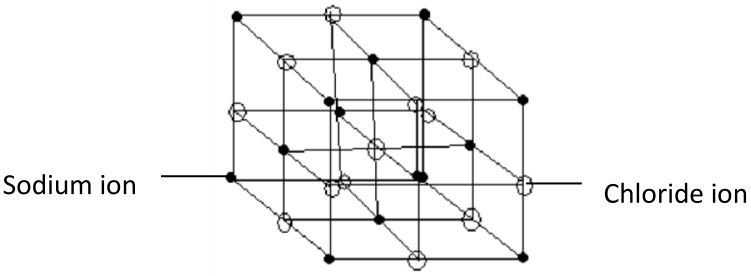
STRUCTURE OF COMPOUNDS/SUBSTANCES

Different types of binds give rise to different types of structures. This is illustrated in the table below.

Bonding	**Structure**
Ionic/ electrovalent	Giant ionic structure
Covalent	Simple molecular structure Giant molecular structure
Metallic	Giant metallic

Giant ionic structure

This is a giant structure consisting of a regular arrangement of a large number of oppositely charged ions resulting into a three dimensional crystal. Each ion is surrounded by a number of oppositely charged ions and they are held together by strong electrostatic forces of attraction. The number of ions surrounding a given oppositely charged ion is known as the **co-ordination number**. The examples of structures include: sodium chloride and magnesium chloride. **Structure of sodium chloride**



Each sodium ion is surrounded by six chloride ions and each chloride ion is also surrounded by six sodium ions. The co-ordination number is there fore 6:6.

Properties of ionic compounds

1. They are solids with a regular shape. This is because of the strong electrostatic forces of attraction keeping the ions closely packed.
2. They have high melting points. This is due to the strong electrostatic force af attraction between oppositely charged ions.
3. They do not conduct electricity in solid states but do conduct in their molten states or solution form. This is because in solid states, the ions are localized and the electrons are not mobile but in molten form, the ions and electrons are free to move (electrons become delocalized) and thus conduct electricity.
4. They have high density as the ions are closely packed.
5. Ionic compounds are soluble in water and other polar solvents but they a re insoluble in organic solvents like benzene. Ionic compounds dissolve

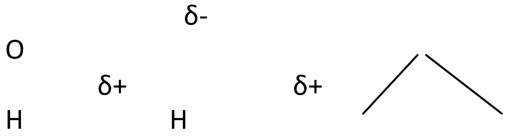
in polar solvents like water as the ions are attracted by the polar molecule.

A polar compound is a covalent compound in which charge separation exist between atoms. This is due to the differences in electro negativity between the atoms.

Electro negativity is the tendency of an atom to attract bonding electron towards itself. Electro positivity is the tendency of an atom to push away bonding electrons from itself.

For example, in a water molecule oxygen is more electro negative and attracts the bonding electrons towards itself giving it a partial negative charge. Hydrogen attains a partial positive charge.

Structure of a water molecule



NB Most metals are malleable and ductile. This is because, since the bonding agent in a metal is a moving electron cloud, the ions of the metal; usually slide relative to one another under stress (without shattering the lattice and produce a new position of stability).

Simple molecular structures

The structure consists of simple independent molecules joined together by weak intermolecular forces of attraction. E.g. Vander Waal‘s force of attraction. The atoms in the molecule are held together by strong covalent bond. Compounds with this structure exist as gases, liquids or solids with low melting points. Examples include iodine, carbon dioxide , ammonia and water.

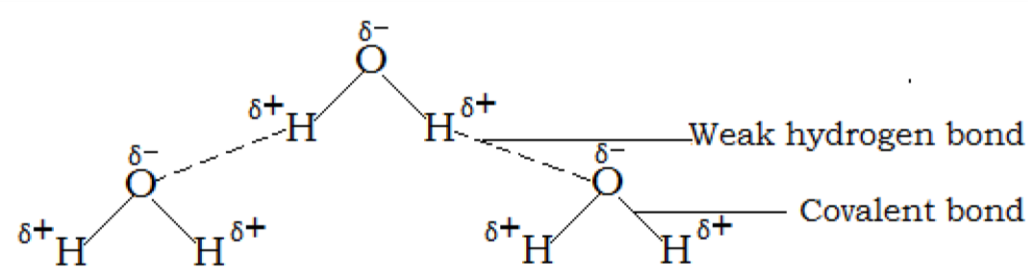
Structure of iodine

I–I-----------I–I----------- --I–I

Strong covalent bond holding iodine atoms together

Weak Vander Waal‘s force holding iodine molecules together

Illustration of weak hydrogen bond in water



Properties of simple molecular structures

1. They have low melting and boiling points because the molecules are held by weak intermolecular forces of attraction.
2. They are non conductors of electricity because they do not have mobile electrons/ ions.
3. Most of them are gases and liquids. Very few are solids.
4. They have low densities as the molecules are not closely packed.
5. They dissolve in organic solvents.

Giant atomic/molecular structure

Compounds with such structures consist of molecules /atoms linked together by strong covalent bonds resulting into a giant three dimensional or three dimensional structures like in graphite and diamond respectively (see details under Carbon and its compounds)

Properties of giant molecular structures

1. They have high melting and boiling pints because of great energy needed to break the bonds.
2. They do not conduct electricity except graphite which has delocalized electrons.
3. They are insoluble in water.

Giant metallic structure

Giant metallic structure consists of very many metal ions surrounded by a sea /cloud of electrons. These electrons are released by metal atoms.

Properties of giant metallic structures

1. They are solids with high melting and boiling points except mercury which is a liquid.
2. They conduct electricity in both liquid and solid states due to the presence of mobile electrons.

Sample questions on bonding and structure

Bonding and structure

1. *Describe briefly what is meant by the following types of bonding: electrovalent, covalent, coordinate and metallic. Use examples to illustrate how the above bonds are formed between any two named atoms or molecules.*

2. *Explain what is meant by metallic bond. Explain why a typical metal is (a) a good conductor of electricity (b) a good conductor of heat (c) in some conditions at leas malleable and ductile.*
3. *The compounds named below are all covalent compounds. With the aid of the table of electron structures, give a diagram for a molecule of each of these compounds showing the outer most electron shells only: (i) tetra chloromethane* CCl_4, *phosphorus(III) chloride* PCl_3 *(iii) silane* SiH_4 *(iv) tri chloromethane* $CHCl_3$ *(v) phosphine* PH_3 *and (vi) dichloromethane* CH_2Cl_2
4. *Explain briefly why ammonia and oxygen molecules participate readily in coordinate bonding. Give an example for the formation of an ion by ammonia by this means. Show by means of electronic diagram the formation of a coordinate linkage between phosphorus tri chloride and oxygen.*
5. *By means of electronic diagrams, show the structure of (a) calcium atom (b) chlorine atom (c) calcium chloride. State the differences between electrovalent and covalent compounds.*
6. *Describe how structures of the following compounds account for their property in term of electrical conduction (a) copper (b) graphite (c) diamond*
7. *With the aid of well labeled drawings explain the structure of the following: Graphite, diamond, sodium chloride and chlorine molecule.*

CARBON AND ITS COMPOUNDS

Carbon is an element in period II and group IV in the periodic table. It has atomic number of six.

Occurrence of carbon

Carbon occurs in both free states and combined states. In combined states, it occurs as coal, mineral oils, carbonates (e.g. lime stone, marble and sea shells) and all living things (animals and plants). As an element it occurs in both natural (as diamond and graphite) and synthetic forms (as coke, charcoal and carbon fiber).

Allotropes of carbon

Allotropy is the existence of an element in two or more forms in the same physical state.

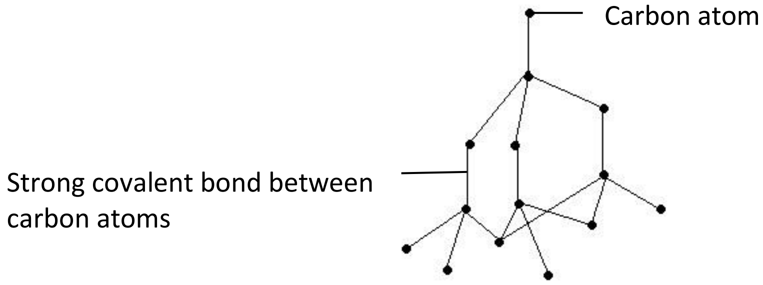
Allotropes are various forms in which elements exist without change in physical state. Different allotropes of the same element always have different crystalline structures and physical properties but the same chemical properties.

Allotropes of carbon are: diamond and graphite. Another form of carbon is amorphous carbon.

Diamond

In diamond, the structure consists of infinite number of carbon atoms. Each of the carbon atoms is joined to four other carbon atoms by covalent bonds resulting into a tetrahedral arrangement. This gives a diamond crystal a giant three dimensional structure. Diamond has no mobile electrons so cannot conduct electricity.

Structure of diamond



Physical properties of diamond

- Diamond is the hardest natural substance known. This is because the carbon atoms are closely parked and are joined by strong covalent bonds.
- Diamond has a very high melting pint because of the strong covalent bond between the carbon atoms.
- It has a very high density (3.5g/cm^3) because of the closely packed carbon atoms.
- Diamond is transparent, sparkling and glitters.
- Diamond does not conduct electricity because it has no mobile electrons.

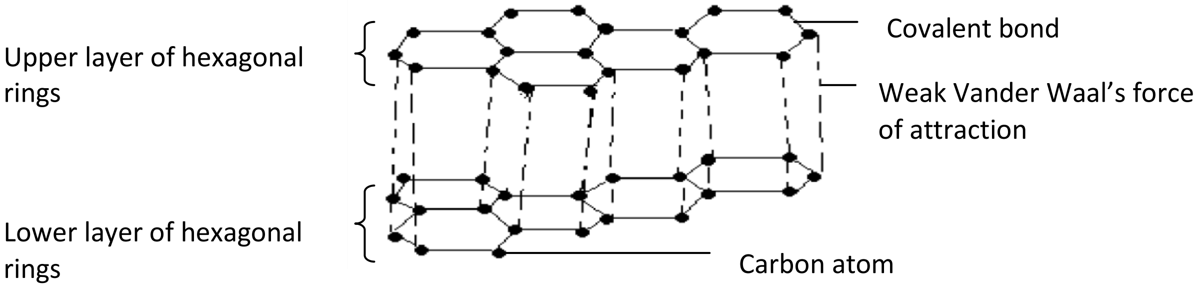
Uses of diamond

1. Diamond is very hard and used as drilling devices, rock borers and glass cutters.
2. Diamond is used jewellery because its sparkling appearance.
3. Diamond is bright and used to make laser beams.

Graphite

Graphite consists of infinite number of carbon atoms each covalently bonded to three other carbon atoms forming hexagonal rings that are arranged in layers. Each layer is a giant two dimensional structure. The different layers are held by weak Vander Waal‘s forces of attraction, making them to slide over each other thus they are slippery and soft. Some electrons in a layer are mobile making them to conduct electricity.

Structure of graphite



Graphite consists of infinite number of carbon atoms each covalently bonded to three other carbon atoms forming hexagonal rings that are arranged in layers. Each layer is a giant two dimensional structure. The different layers are held by weak Vander Waal‘s forces of attraction, making them to slide over each other thus they are slippery and soft. Some electrons in a layer are mobile making them to conduct electricity.

Physical properties of graphite

- Graphite conducts electricity because they have free mobile electrons.
- The melting point of graphite is high because of the strong covalent bond between the carbon atoms.
- Graphite is soft and slippery because its layers are held by weak Vander Waals force of attraction.
- Graphite is opaque and dark in color and shinny.

Graphite is less resistant to chemical attack than diamond because of the open spaces between the layers. The density of graphite is 2.3g/cm^3

Uses of graphite

1. Graphite is soft and can mark there fore used to make pencil ―leads‖. Graphite is mixed with clay to make pencil leads.
2. Graphite is a good conductor of electricity and thus used as electrodes.
3. It is soft and greasy, therefore used as lubricants especially in small bearings like those in dynamos.
4. Graphite is used to make brushes for electric motors.
5. Graphite (black lead) is used as a protective coating on iron to prevent rusting.

When 1g of both diamond and graphite burns in oxygen, they form 3.67g of carbon dioxide only. This shows that they are allotropes.

Amorphous carbon

Amorphous carbon is black and has the lowest density. It is non crystalline and fairly conducts electricity. Amorphous carbon exist in several forms like wood charcoal, animal charcoal, sugar charcoal, lamp black and coke. **i) Wood charcoal**

It is formed when wood is heated in limited supply of air (destructive distillation of wood). It is a black porous solid and a very good absorbent of gases.

ii) Animal charcoal

This is made by heating animal refuse and bones in limited supply of air. Animal charcoal has a property of absorbing coloring matter and is used to remove brown color from sugar during its manufacture.

iii) Sugar charcoal

This is a very pure form of carbon made by removing elements of water from sugar. E.g. when cane sugar is dehydrated by concentrated sulphuric acid, sugar charcoal is formed.

iv) Lamp black

This is formed by burning oils (e.g. turpentine, petroleum, kerosene) in limited supply of air. It is used for making ink for printing and shoe polish.

v) Coke

This is an impure form of carbon made by heating coal in the absence of air (destructive distillation of coal). It is used as a reducing agent in the extraction of metals like iron and zinc from their ores. Coke is also used as fuel.

Chemical properties of carbon

All allotropes of carbon have similar chemical properties but different physical properties.

1. Reaction of carbon with oxygen

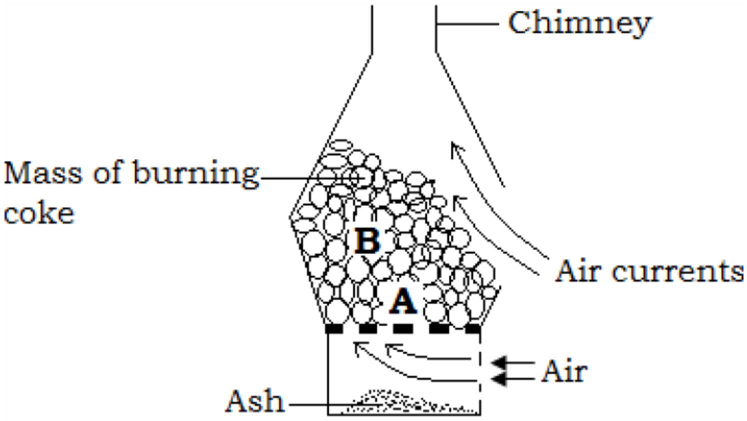
Carbon burns is excess oxygen to for carbon dioxide gas. In the process great heat is generated.

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

When carbon burns in limited amount of oxygen, carbon monoxide forms.

$$C(s) + \frac{1}{2}O_2(g) \longrightarrow CO(g)$$

Reaction of carbon in a deep, brightly glowing coke/coal fire.



In region A, plenty of air is available. Carbon burns to carbon dioxide.

$$C(s)+O_2(g) \longrightarrow CO_2(g)$$

In region B, ascending carbon dioxide is reduced by red hot coke to carbon monoxide.

$CO_2(g) + C(s) \longrightarrow 2CO(g)$

At the surface, the hot carbon monoxide burns in air to form carbon dioxide with a flickering blue flame.

$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$

2. Reaction of metal oxides with carbon

Carbon readily removes oxygen from (reduces) oxides of metals lower than it in the reactivity series. Thus it acts as a reducing agent. Examples of oxides reduced by carbon are: zinc oxide, ZnO; lead (II) oxide, PbO; copper (II) oxide, CuO; iron (II) oxide, FeO and iron (III) oxide, Fe_2O_3.

$2ZnO(s) + C(s) \longrightarrow 2Zn(s) + CO_2(g)$
$2PbO(s) + C(s) \longrightarrow 2\ Pb(s) + CO_2(g)$
$2CuO(s) + C(s) \longrightarrow 2Cu(s) + CO_2(g)$

3. Reaction with acids

Carbon reduces nitric acid and sulphuric acid to nitrogen dioxide and sulphur dioxide gases respectively. The carbon itself is oxidized to carbon dioxide gas.

$C(s) + 4HNO_3(aq) \longrightarrow 4NO_2(g) + CO_2(g) + 2H_2O(l)$
$C(s) + 2\ H_2SO_4(aq) \longrightarrow 2SO_2(g) + CO_2(g) + H_2O(l)$

4. Reaction with steam

If steam is blown through red hot charcoal, a mixture of carbon monoxide and hydrogen gases are formed. This mixture is called **water gas.**

$H_2O(g) + C(s) \longrightarrow CO(g) + H_2(g)$

OXIDES OF CARBON

CARBON MONOXIDE (CO)

Occurrence

Carbon monoxide is a poisonous gas and is commonly present in coal gas and other gaseous fuels. It is also produced in car exhaust fumes due to partial combustion of carbon.

Laboratory preparation

This can be done by the action of concentrated sulphuric acid on any of the following compounds.

1. Methanoic acid (HCOOH)

Concentrated sulphuric acid is a very strong dehydrating agent and removes water from methanoic acid forming carbon monoxide and water.

$HCOOH(l) \xrightarrow{\text{Conc. } H^2SO^4} CO(g)$
$+ H_2O(l)$ No heating is required.

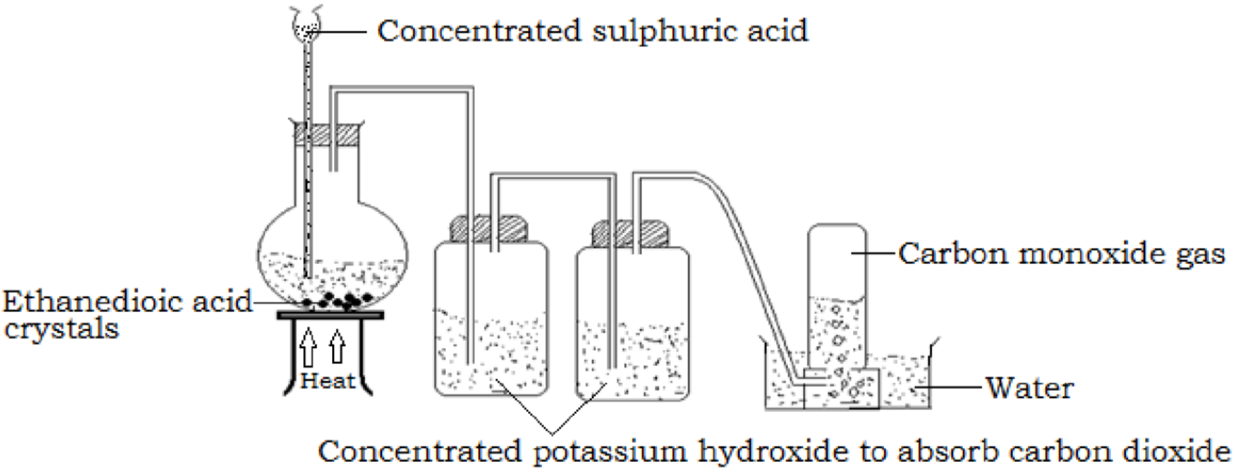
2. Sodium methanoate (HCOONa)

Here, concentrated sulphuric acid dehydrates sodium methanoate to form water, sodium hydrogen sulphate and carbon monoxide.

$$HCOONa(s) + H_2SO_4(aq) \longrightarrow NaHSO_4(aq) + CO(g)\ H_2O(l)$$

3. Oxalic acid (ethanedioic acid-$H_2C_2O_4$)

Set up of apparatus



Equation

$$H_2C_2O_4(s) \xrightarrow{\text{Conc. } H^2SO^4} CO_2(s) + CO(g) + H_2O(l)$$

In this preparation, the products formed are passed through concentrated potassium hydroxide solution to remove traces of carbon dioxide.

$$2KOH(aq) + CO_2(g) \longrightarrow K_2CO_3(aq) + H_2O(l)$$

If the carbon monoxide is required dry, it is passed through a bottle containing conc. Sulhuric acid and collected by use of a syringe or upward displacement of air.

Physical properties of carbon monoxide - It is colorless, odourless and tasteless.

- It is insoluble in water.
- It is denser than air.
- It is neutral to litmus paper

Chemical properties

a) Reaction with air

Carbon dioxide burns in air with a pale blue flame to produce carbon dioxide.

$$CO(g) + ½O_2(g) \longrightarrow CO_2(g)$$

b) Reducing action of carbon monoxide

Carbon dioxide is a powerful reducing agent and it reduces the oxides of metals (iron (illustrated in extraction of iron), lead and copper) to respective metals upon heating.

$$CuO(s) + CO(g) \longrightarrow Cu(g) + CO_2(g)$$

(Black solids) (Brown solids)

$$PbO(s) + CO(g) \longrightarrow Pb(s) + CO_2(g)$$

(Yellow solids-cold) (Grey solids)

Poisonous nature of carbon monoxide

Blood contains haemoglobin which combines with oxygen to form oxyhaemoglobin. The oxyhaemoglobin transports and supplies the oxygen to all parts of the body. However, when carbon monoxide is inhaled, it combines with haemoglobin to form carboxyhaemoglobin. This stops the blood from absorbing oxygen and may lead to death.

Haemoglobin has a higher affinity for carbon monoxide than oxygen. Carbon monoxide is even more dangerous because it has no colour and smell.

Uses of carbon monoxide

- It is used extensively in the extraction of metals as a reducing agent e.g. extraction of iron in a blast furnace.
- It is used as fuel in form of producer gas and water gas.
- It is used in the manufacture of methanol used in anti freezer mixture in cold countries to prevent ice from forming in car radiators.
- It is used in the manufacture of synthetic petrol.

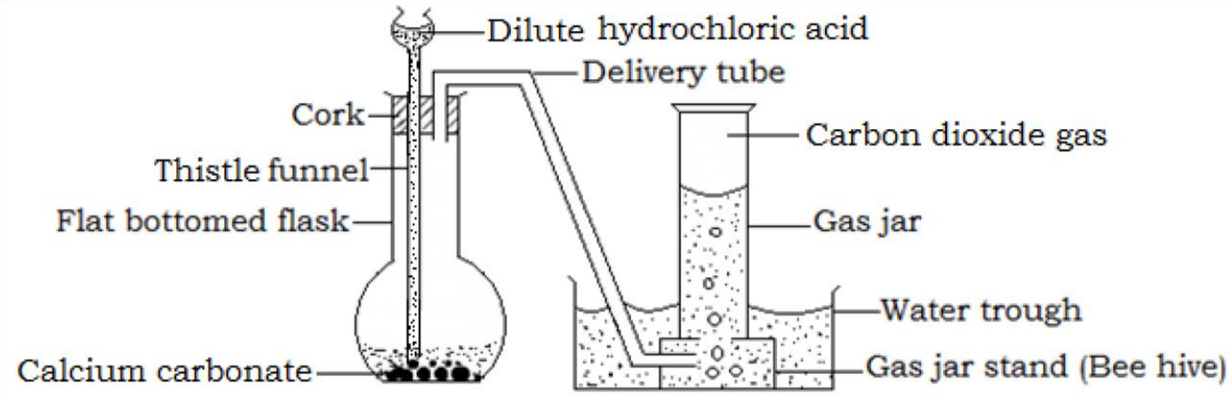
CARBON DIOXIDE (CO_2)

Carbon dioxide occurs in air and occupies 0.03% by volume. Carbon dioxide forms from rocks as a result of volcanic eruption; occurs in mines as ―choke damp‖ and it is always present in natural drinking water because it is slightly soluble in it.

Laboratory preparation

It is prepared by the action of dilute hydrochloric acid on marble chips (a form of calcium carbonate)

Set up of apparatus



Arrange the apparatus as shown above with the calcium carbonate in the flask. Run dilute hydrochloric acid in to the flask through a funnel.

Observation

Effervescence occurs as a colorless gas is given off.

Equation

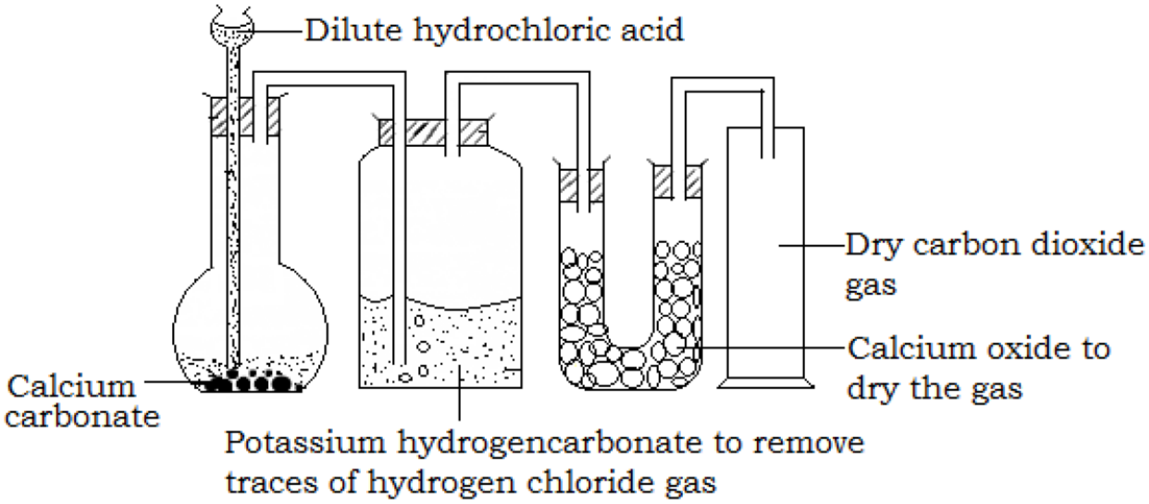
$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

Ionic equation

$CO_3^{2-}(aq) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$

The gas may be collected over water as it is only slightly soluble in water. If the gas is required dry, it is collected by down ward delivery method.

If a pure dry sample of carbon dioxide is require, the gas is first passed through a wash bottle containing concentrated solution of potassium hydrogen carbonate to remove traces of hydrogen chloride gas (water placed in a wash bottle can also be used). It is then passed through a U tube containing fused calcium chloride to dry the gas. The gas is then collected by down ward delivery method since it is denser than air.



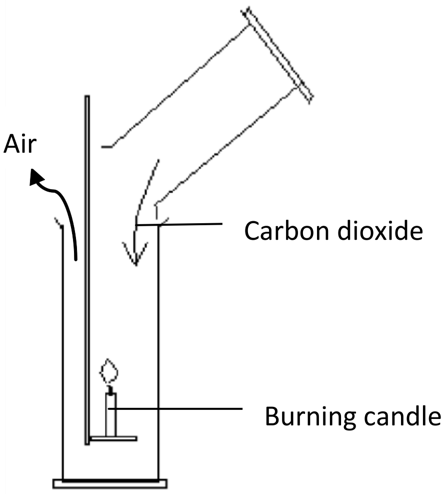
Sulphuric acid is not used in the preparation of carbon dioxide because the salt formed (calcium sulphate) forms a coating on the surface of calcium carbonate and this prevents further reaction between the acid and the calcium carbonate as the calcium sulphate coating is insoluble. The reaction there fore stops prematurely.

Other reactions that produce carbon dioxide include: fermentation; combustion of hydrocarbons; respiration and action of heat on carbonates and hydrogen carbonates.

Properties of carbon dioxide

Physical properties

- It is colorless, odourless and tasteless.
- It does not burn and does not support burning. Because of its density, it is also used as a fire extinguisher as it displaces oxygen that supports burning. **Illustration**



- It is denser than air that is why it is collected by down ward delivery method. Under high pressure, carbon dioxide is quite soluble in frizzy drinks (effervescent drinks)
- It is slightly soluble in water forming a weakly acidic solution of carbonic acid.

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

- It turns moist blue litmus paper pink indicating that it is weakly acidic.

Chemical properties

1. Carbon dioxide does not support burning. However when a piece of burning magnesium is lowered into a gas jar of carbon dioxide, it continues to burn forming black specks of carbon mixed with white solids of magnesium oxide.

   $$2Mg(s) + CO_2(g) \longrightarrow 2MgO(s) + C(s)$$

   In this case, magnesium reduces carbon dioxide to carbon ant it self is oxidized to magnesium oxide.

2. Carbon dioxide turns lime water (calcium hydroxide) milky. This is due to the formation of insoluble calcium carbonate (white precipitates)

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow \underset{\text{(White precipitate)}}{CaCO_3(g)} + H_2O(l)$$

However, when excess carbon dioxide is bubbled, through the solution (lime water), the white precipitates dissolve making the solution to appear clear. This is due to the formation of calcium hydrogen carbonate which is a soluble compound.

$$CaCO_3(s) + H_2O(l) + CO_2(g) \longrightarrow Ca(HCO_3)_2(aq)$$

3. When carbon dioxide is bubbled through a solution of fairly concentrated sodium hydroxide, the solution remains clear due to the formation of a soluble sodium carbonate.

$$2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

This reaction is sometimes used to remove carbon dioxide from a mixture of gases. However, when excess carbon dioxide is bubbled through the above solution, white precipitates appear due to formation of insoluble sodium hydrogen carbonate.

$$Na_2CO_3(aq)+H_2O(l)+CO_2(g) \longrightarrow \underset{\text{(White precipitate)}}{2NaHCO_3(s)}$$

4. When carbon dioxide gas is bubbled through a solution of potassium hydroxide, there is no observable change as the products formed are all soluble.

$$2KOH(aq) + CO_2(g) \longrightarrow \underset{\text{(Soluble)}}{K_2CO_3(aq)} + H_2O(l)$$

$$K_2CO_3(aq) + H_2O(l) + CO_2(g) \longrightarrow \underset{\text{(Soluble)}}{2KHCO_3(aq)}$$

Uses of carbon dioxide

- It is used as a fire extinguisher as it does not support burnng and is denser than air.
- It is used in the manufacture of effervescent drinks. This is because the solution of the gas in water has a pleasant taste (the taste of soda water).
- Solid carbon dioxide (dry ice) is used as a refrigerating agent for perishable goods.
- Carbon dioxide is used in the manufacture of baking powder.
- Solid carbon dioxide (dry ice) is fired into the cloud where it cools to form rain. This happens in places where there is unreliable rain fall.

Test for carbon dioxide

When carbon dioxide gas is passed through a solution of calcium hydroxide (lime water), the solution becomes milky. This is due to the formation of calcium carbonate which is insoluble.

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow \underset{\text{(White precipitate)}}{CaCO_3(s)} + H_2O(l)$$

CARBONATES AND HYDROGEN CARBONATES

These are salts derived from carbonic acid.

a) A part from potassium, sodium and ammonium carbonates, all other carbonates are insoluble in water.

b) All solid hydrogen carbonates are soluble in water except sodium hydrogen carbonate which is only slightly soluble. Hydrogen carbonates of metals lower than magnesium in the reactivity series do not exist. Hydrogen carbonates of calcium and sometimes magnesium only exist in solution form.

2. Action of heat on carbonates and hydrogen carbonates

a) Carbonates of sodium, potassium and lithium are not decomposed by heat. However if they are hydrated, they only lose their water of crystallization. E.g.

$$\underset{\text{(White crystals)}}{Na_2CO_3.10H_20(s)} \longrightarrow \underset{\text{(White powder)}}{Na_2CO_3(s)} + 10H_2O(l)$$

b) All the other metal carbonates decompose when heated to form the oxide of the metal and a colorless gas that turns lime water milky (carbon dioxide) e.g.

$$\underset{\text{(Green powder)}}{CuCO_3(s)} \longrightarrow \underset{\text{(Black powder)}}{CuO(s)} + CO_2(g)$$

$$\underset{\text{(White powder)}}{ZnCO_3(s)} \longrightarrow \underset{\text{(Yellow-hot; white-cold)}}{ZnO(s)} + CO_2(g)$$

c) Ammonium carbonate sublimes when heated and forms ammonia, carbon dioxide and water vapor.

$$(NH_4)_2CO_3(s) \longrightarrow 2NH_3(g) + H_2O(g) + CO_2(g)$$

d) All hydrogen carbonates decompose to give corresponding carbonates, carbon dioxide gas and water vapor. E.g.

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

3. Action of dilute acids

All carbonates and hydrogen carbonates react with dilute acids to liberate carbon dioxide gas, water and corresponding salts. E.g.

$$2HCl(aq) + CaCO_3(s) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

$$H_2SO_4(aq) + 2NaHCO_3(s) \longrightarrow Na_2CO_3(aq) + CO_2(g) + H_2O(l)$$

N.B. Dilute sulphuric acid reacts with calcium carbonate and lead (II) carbonate at a very slow rate as the resulting salts formed (calcium sulphate and lead(II) sulphate) are insoluble and there fore tend to form coatings around the carbonates inhibiting further reactions between the carbonates and the acid.

The reaction between dilute hydrochloric acid and lead (II) carbonate also forms an insoluble salt (lead (II) chloride) which also forms coating stopping further reactions.

Dilute nitric acid reacts with all carbonates to form soluble nitrate salts.

Testing for carbonates and hydrogen carbonates

a) If the sample is a solid

Add a little dilute acid to the solid substance under test.

Observation

Effervescence occurs and a colorless gas that turns lime water milky (carbon dioxide) is given off.

Conclusion

This shows that a carbonate or hydrogen carbonate is present.

b) Carbonate in solution form

To 2cm^3 of test solution add 3 drops of lead (II) nitrate solution followed by excess dilute nitric acid.

Observation

White precipitate is formed and the precipitate dissolves in excess nitric acid with effervescence.

N.B. Sometimes barium nitrate solution or barium chloride solution may be used in place of lead (II) nitrate solution.

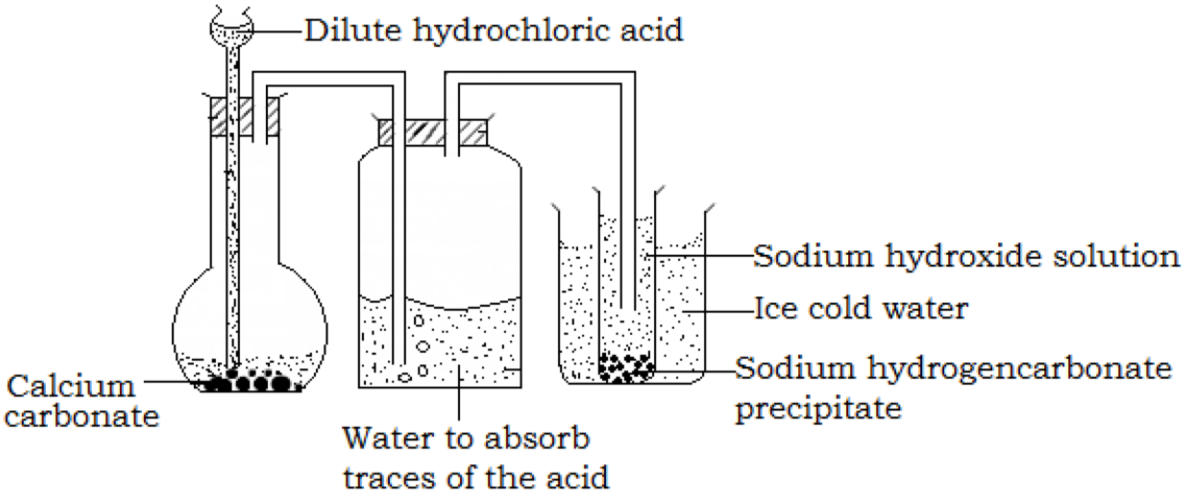
$Na_2CO_3(s) + Pb(NO_3)_2(aq) \longrightarrow PbCO_3(s) + 2NaNO_3(aq)$

$PbCO_3(s) + HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

Sodium carbonate

Laboratory preparation

Set up



Procedure

- Bubble carbon dioxide through a fairly concentrated sodium hydroxide solution until in excess. White precipitate of sodium hydrogen carbonate is formed according to the equation below

$$2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

$$Na_2CO_3(aq)+H_2O(l)+CO_2(g) \longrightarrow 2NaHCO_3(s) \text{ (White precipitate)}$$

The white precipitate of sodium hydrogen carbonate is filtered off, washed and dried.

- The sodium hydrogen carbonate is heated strongly until no further water vapor and carbon dioxide are given off. This leaves a white powder of calcium carbonate.

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

Commercial preparation of sodium carbonate (Solvay process)

Very concentrated brine (28% sodium chloride) is saturated with ammonia gas in a tower to form ammonia gas in a tower to form ammoniacal brine. The ammoniacal brine is run downwards from the top of the tower while carbon dioxide (formed from decomposition of calcium carbonate) is forced to rise up the tower from the base of the tower.

The tower is fitted with perforated mushroom shaped baffles at intervals that delay the flow of ammoniacal brine and also offer surface for the reaction.

The ammoniacal brine reacts with carbon dioxide to form sodium hydrogen carbonate as precipitates since it is not very soluble in water. Precipitation is assisted by cooling the lowest third of the chamber.

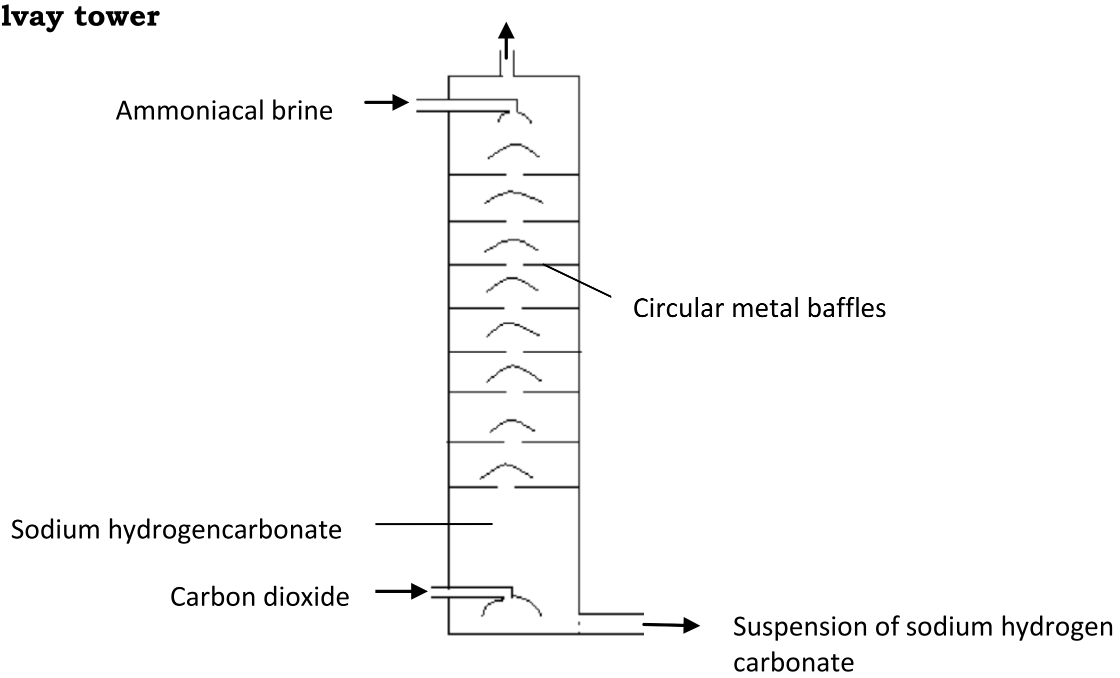
$$NaCl(aq) + NH_4OH(aq) + CO_2(g) \longrightarrow NaHCO_3(s) + NH_4Cl(aq)$$

Sodium hydrogen carbonate is filtered from the white sludge at the base of the tower and washed to remove ammonium compounds. The sodium hydrogen carbonate is the heated to form sodium carbonate.

$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$

The anhydrous sodium carbonate formed ha a wide market.

The Solvay tower



If crystalline form (washing soda) is required, the anhydrous solid is dissolved in hot water, crystallization takes place as the solution cools. The crystals are removed and allowed to dry.

$Na_2CO_3(aq) + 10H_2O(l) \longrightarrow Na_2CO_3.10H_2O(s)$

Sodium carbonate decahydrate, $Na_2CO_3.10H_2O$ are large translucent crystals. When the crystals are exposed to air, they lose mass and become coated with fine powder which makes is opaque. Each molecule of washing soda gives to the atmosphere 9 molecules of water of crystallization.

$Na_2CO_3.10H_2O(s) \longrightarrow Na_2CO_3.H_2O(s) + 9H_2O(g)$

(Sodium carbonate monohydrate)

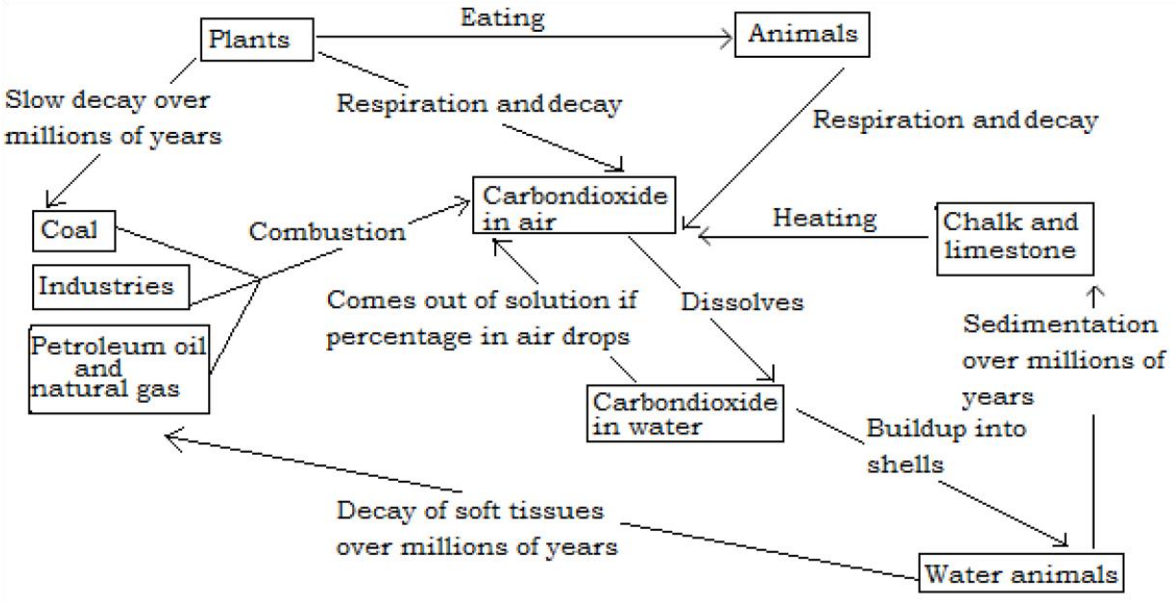
Such lose of water of crystallization to the atmosphere is termed as efflorescence. The substance that loses its water of crystallization is known as an efflorescent substance.

Uses of sodium carbonate

1. Manufacture of glass. Ordinary bottle glass is made by fusing together sodium carbonate, calcium carbonate, silicon dioxide (sand) and a little carbon (reducing agent). Broken glasses are added to assist fusion.
2. Manufacture of water glass that is used to preserve eggs, used in fire proofing and production of cement.
3. Sodium carbonate is used in the manufacture of soap powders.
4. Sodium carbonate in used in domestic water softening. Calcium ions (Ca^{2+}) which is the principal cause of hardness in water is precipitated from the water as calcium carbonate by adding sodium carbonate.

The carbon cycle

The balance of processes which give out carbon dioxide and those which use carbon dioxide is called the **carbon cycle.** Summary of the carbon cycle is given below.



Carbon dioxide in the atmosphere is added from: respiration of plants and animals; decay/ decomposition of plants and animals; heating of lime stone (to give quick lime) in lime kilns; water (if the percentage in air drops); and combustion of coal (formed as a result of slow decomposition of plants and animals over millions of years), petroleum oil and natural gas.

The processes that remove carbon dioxide from the atmosphere are: photosynthesis by green plants; and dissolution in water.

Animals obtain carbon from plants by feeding on food such as starch made by plants..

Effects of carbon dioxide on atmospheric temperature

The sun emits radiations that pass through the atmosphere of the earth with little absorption and warms up the surface of the earth (ground). The warm surface of the earth (ground) reflects back the radiations inform of infrared radiations. The infrared radiation is absorbed by gases like carbon dioxide and methane which radiate some heat back to the ground leading to warming up of the earth. The warming up of the earth is known as **green house effect**. The gases like carbon dioxide and methane that cause the warming up of the earth are referred to as **green house gases.**

When more green house gases are released into the atmosphere, much heat accumulates and this leads to general rise in temperature of the world. The general rise in temperature of the world is known as **global warming**.

Sample questions on Carbon and its compounds

1. *Define the terms allotrope and allotropy. Name two crystalline forms of the allotropes of carbon and describe their structures. Outline three differences between the two allotropes. Mention two uses of each allotrope. How do you prove that the allotropes consist solely of carbon?*
2. *Amorphous carbon exists in several forms name them and briefly explains how each form is obtained and give at least one use of each form of amorphous carbon.*
3. *With the aid of equations describe the reactions of carbon with oxygen, acids, steam and metallic oxides.*
4. *Describe carefully how you would prepare and collect carbon monoxide free from carbon dioxide. Give three ways in which this gas is different from carbon dioxide. Describe the reducing action of carbon monoxide on copper(II) oxide. Explain briefly how water gas and producer gas are obtained. How can exhaust fumes of cars act as poison?.*
5. *By means of a labeled drawing, and equation, show how a sample of carbon dioxide can be made and collected in the laboratory. Describe briefly without using diagrams how you can convert carbon dioxide to carbon monoxide. Mention two large scale uses of carbon dioxide.*
6. *Explain with equations the observations made when carbon dioxide is bubbled through calcium hydroxide (lime water) and sodium hydroxide until in excess.*
7. *Describe how sodium carbonate is prepared in the laboratory and in large scale from the Solvay process. Equations and drawings must be used to illustrate.*