TOPIC 1 : RATES OF REACTIONS

RATE OF REACTION

The rate of chemical reaction measures how quickly a reactant is used up, or a product is formed. It is defined as the change in the amount of a reactant or product per unit time.

THE COLLISION THEORY

For a chemical reaction to occur;

- the reactant particles must collide with each other
- the particles must have enough energy for the collision to be successful

A **successful collision** is the one that produces a reaction. The rate of reaction depends on how many successful collisions there are in a given time. The more successful collisions there are, the faster the reaction. The minimum amount of energy needed for a collision to be successful is called activation energy.

FACTORS AFFECTING THE RATE OF REACTION

Some of the factors that affect the rate of a reaction are:

a. Concentration of the reactants

The rate of a reaction increases when the concentration of a reactant is increased. When the concentration of the solution increases, the amount of the particles per unit volume also increases. Frequency of collisions between reacting particles increases.

b. Temperature

The rate of reaction increases with an increase in temperature. At higher temperatures, the reactant particles have more kinetic energy, they move faster, and collide more often with greater energy causing the rate of reaction to increase.

c. Surface area of solid reactants

When the size of particles of a solid decrease, surface area of the solid increases, allowing them to collide with other reacting particles more readily hence the rate of reaction of increases. This is why powdered solids react faster than lumps.

d. Use of a catalyst

A catalyst is a substance that speeds up a chemical reaction, but remains chemically unchanged itself. Catalysts work by lowering the activation energy for the reaction. Lowering the activation energy increases the frequency of successful collisions because a greater proportion of collisions now exceed this lower energy.

METHODS OF MEASURING RATES OF REACTIONS

There are different methods to measure the rate of reaction. The method chosen usually depends on the reactants and products involved and how easy it is to measure the changes in them.

a. Measuring the volume of gas produced.
b. Measuring the change in mass of reactants or products.
c. Measuring the time taken to for a given mass of reactant to disappear completely.
d. Measuring the time taken for a certain amount of product to form.

MEASURING THE VOLUME OF A GAS PRODUCED.

This method is used when a gas leaves the container. The volume is measured using a gas syringe.
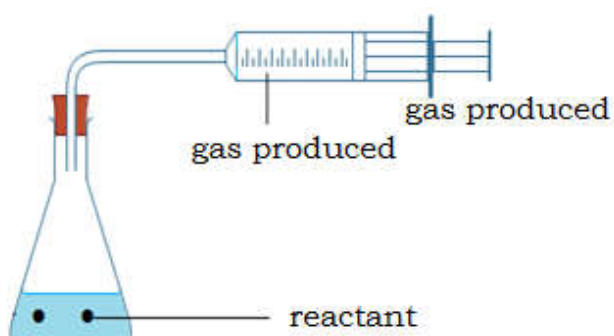
The volume of the gas given off is measured at regular time intervals, for example every half a minute. Using the syringe, the volume of the gas give off, is noted by checking how the plunger has been pushed. For example, if the plunger is pushed to the 20 cm^3 mark, it means 20 cm^3 of gas has been collected.

Instantaneous rate of reaction is calculated by the following formula:

$$\text{Rate of reaction} = \frac{\text{volume of gas at } t_2 - \text{volume of gas at } t_1}{t_2 - t_1}$$

Average rate of reaction is calculated by the following formula:

Average rate of reaction $=$ $\dfrac{\text{total volume of gas produced}}{\text{reaction time}}$

Worked example

In an experiment, some magnesium and excess of dilute hydrochloric acid were reacted together. The volume of hydrogen produced was recorded every half a minute.

Time (minutes)	Volume of hydrogen (cm^3)
0	0
0.5	8
1.0	14
1.5	20
2.0	25
2.5	29
3.0	33
3.5	36
4.0	38
4.5	39
5.0	40
5.5	40
6.0	40

a. Plot a graph of volume of hydrogen against time.
b. What is the rate of reaction (in cm^3 of hydrogen per minute) during, the 1st minute, the 2nd minute and the 3rd minute?
c. How much hydrogen was produced in total?
d. How long does the reaction last?
e. Calculate the average rate of reaction

Solutions

a. The graph of volume of hydrogen against time is shown below

The reaction between magnesium and dilute hydrochloric acid
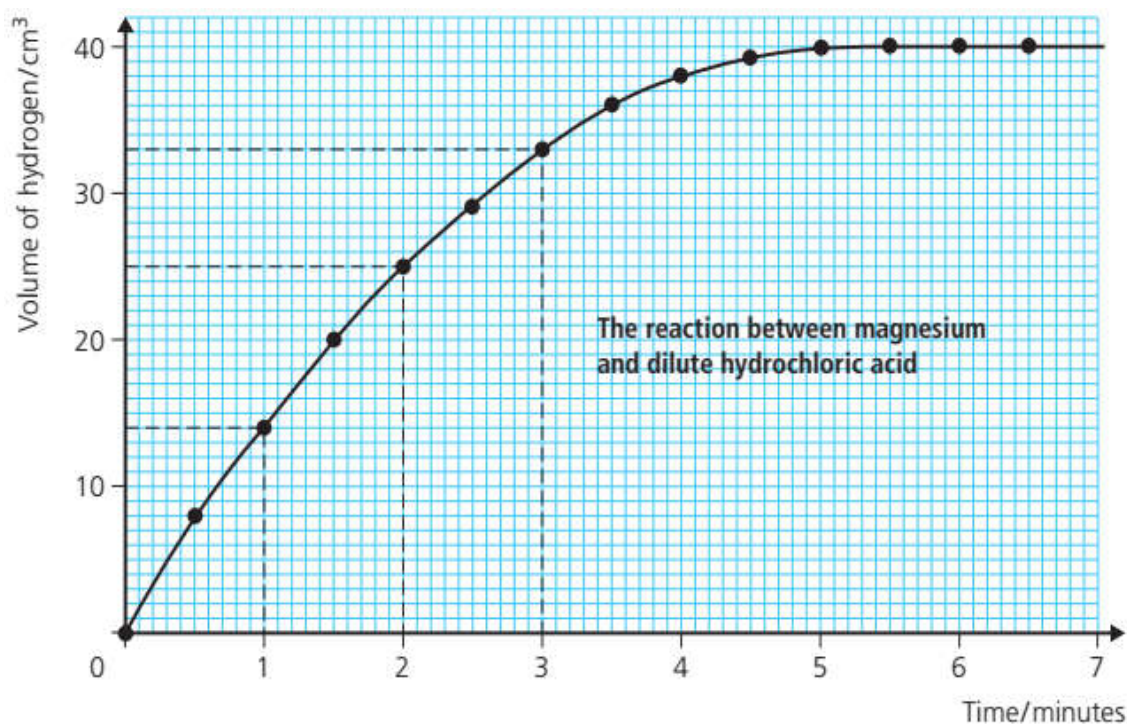
b. Rate of reaction in the 1st minute $\quad = \quad$ 14 cm^3 of hydrogen per minute

Rate of reaction in the 2nd minute $\quad = \quad \dfrac{25-14}{1}$

$= \quad$ 11cm^3 of hydrogen per minute

Rate of reaction in the 3rd minute $\quad = \quad \dfrac{33-25}{1}$

$= \quad$ 8cm^3 of hydrogen per minute

c. In total, 40 cm^3 of hydrogen was produced.

d. After 5 minutes no more hydrogen is produced. The reaction lasts 5 minutes

e. Average rate of reaction $\quad = \quad \dfrac{\text{total volume of hydrogen}}{\text{total time for the reaction}}$

$= \quad \dfrac{40\text{cm}^3}{5\text{minutes}}$

$= \quad$ 8cm^3 of hydrogen per minute

REVERSIBLE AND IRREVERSIBLE REACTIONS
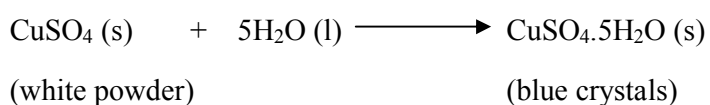
a. Reversible reaction

A reversible reaction is a chemical reaction in which the products can be converted back to reactants under suitable conditions.
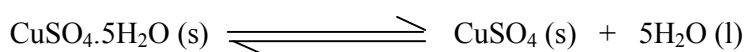
Example

When blue copper hydrated copper (II) sulphate is heated, it loses it water of crystallization. A white anhydrous copper sulphate is formed. The equation for the dehydration of blue copper sulphate is:

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

(blue crystals)   (white powder)
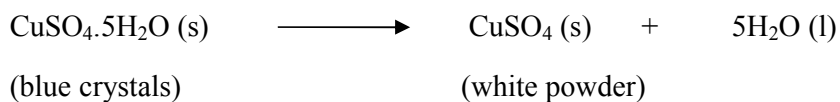
When water is added to the white anhydrous copper (II) sulphate powder, the colour turns back to blue. This is because the water of crystallization is regained. The equation for hydration of white anhydrous copper sulphate is:

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

(white powder)   (blue crystals)

The overall reaction is given as

$CuSO_4.5H_2O$ (s) $\rightleftharpoons$ $CuSO_4$ (s) + $5H_2O$ (l)

b. Irreversible reaction

An irreversible reaction is a reaction in which products cannot be converted back into reactants.

Example

When calcium carbonate is decomposed, calcium oxide and carbon dioxide are produced. The reaction cannot be reversed. The carbon dioxide escapes into the air.

Calcium carbonate $\xrightarrow{\text{heat}}$ Calcium oxide + Carbon dioxide

CaCO (s) $\xrightarrow{\text{heat}}$ CaO (s) + CO (g)
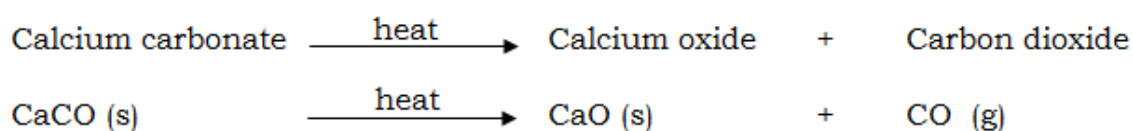
TOPIC 2 : ACIDS AND BASES

THE BRØNSTED-LOWRY THEORY OF ACIDS AND BASES

The BrØnsted-Lowry theory of acids and bases was proposed by a Danish chemist, Johnnes BrØnsted and an English chemist, Thomas Lowry. Their definition centres on the proton, H^+. The proton is what remains after the normal hydrogen atom loses an electron. According to the BrØnsted-Lowry theory:

An acid is a proton donor.

$$HCl \longrightarrow H^+ + Cl^-$$
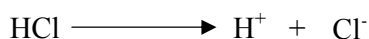
Acid   **Proton**

A base is a proton acceptor.

$$H_2O + H^+ \longrightarrow H_3O^+$$

Base   **Proton**

The overall acid-base reaction is: $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$

                                    Acid   Base

FORMATION OF THE HYDRONIUM ION, H_3O^+

The hydronium ion is formed when an acid dissolves in water, and the water gains the hydrogen ions which are donated by the acid.

For example, when hydrochloric acid dissolves in water, it dissociates into hydrogen ions, H^+ and chloride ions, Cl^- as follows.
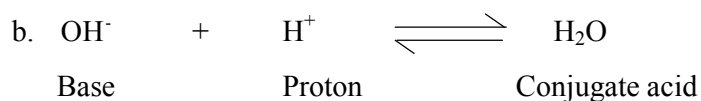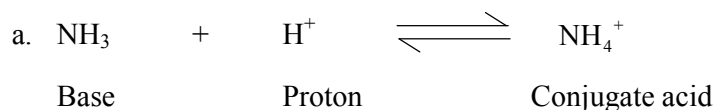
$$HCl \longrightarrow H^+ + Cl^-$$
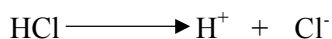
The protons are then gained by the water to form the hydronium ion, H_3O^+ as follows:

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$
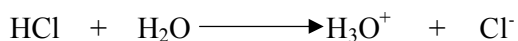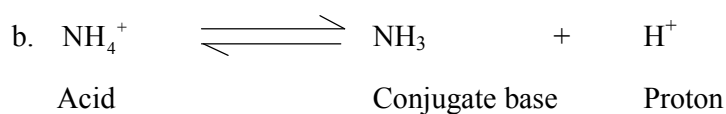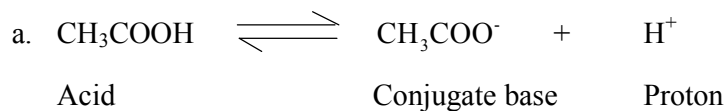
CONJUGATE ACIDS AND CONJUGATE BASES

The terms **conjugate acid** and **conjugate base** are used to decribe products of acid-base reactions.

A conjugate acid is the ion that is formed when a base gains a proton. Examples are:

a. NH_3 + H^+ $\rightleftharpoons$ NH_4^+

    Base   Proton   Conjugate acid

b. OH^- + H^+ $\rightleftharpoons$ H_2O

    Base   Proton   Conjugate acid

The ion formed is called the conjagate acid because, if we imagine the reverse of the reaction, it can give up a proton.

A conjugate base is the ion that is formed when an acid loses a proton. Examples are:

a. CH_3COOH $\rightleftharpoons$ CH_3COO^- + H^+

    Acid   Conjugate base   Proton

b. NH_4^+ $\rightleftharpoons$ NH_3 + H^+

    Acid   Conjugate base   Proton

The ion formed is called the conjugate base because if we imagine, the reverse of the reaction, it can accept a proton.

Examples
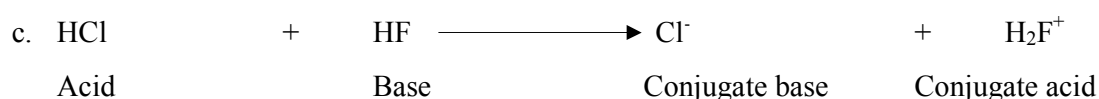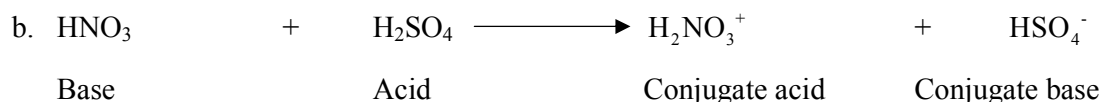
Identifying conjugate acid and conjugate base in acid-base reactions

a. CH_3COOH + HCl $\longrightarrow$ CH_3COOH_2 + Cl^-

    Base   Acid   Conjugate acid   Conjugate base

b. HNO_3 + H_2SO_4 $\longrightarrow$ $H_2NO_3^+$ + HSO_4^-

    Base   Acid   Conjugate acid   Conjugate base

c. HCl + HF $\longrightarrow$ Cl^- + H_2F^+

    Acid   Base   Conjugate base   Conjugate acid

CONJUGATE ACID-BASE PAIRS

A conjugate acid-base pair consists of molecules or ions which are connected by the **loss** of one proton, H^+ by an acid and by the **gain** of one proton, H^+ by a base.

Every acid-base reactions contains two conjugate acid-base pairs, because the H^+ is transferred in both the forward and reverse reactions.

Examples

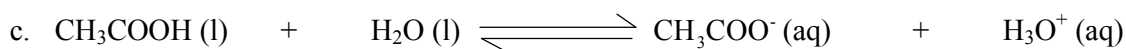Write down the conjugate acid-base pairs in the following equations.

a. NH_3 (g)   + H_2O (l) $\rightleftharpoons$ NH_4^+ (aq)   + OH^- (aq)
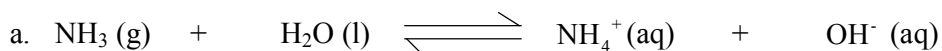
Solution

The conjugate acid-base pairs are NH_4^+ (aq)/NH_3 (g) and H_2O (l)/ OH^- (aq)

b. HNO_3 + H_2SO_4 $\longrightarrow$ $H_2NO_3^+$ + HSO_4^-

Solution

The conjugate acid-base pairs are: H_2SO_4/ HSO_4^- and $H_2NO_3^+$ / HNO_3

c. CH_3COOH (l)   + H_2O (l) $\rightleftharpoons$ CH_3COO^- (aq)   + H_3O^+ (aq)

Solution

The conjugate acid-base pairs are: CH_3COOH (l)/ CH_3COO^- (aq) and H_3O^+ (aq)/H_2O (l)
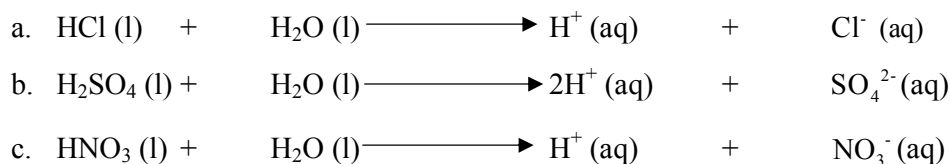
STRENGTH OF ACIDS AND BASES

The strength of a particular acid or base is based on the degree to which it dissociate into ions in aqueous solutions.
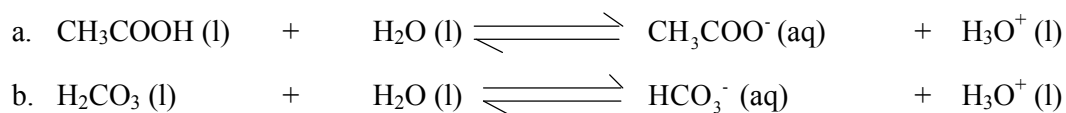
STRONG AND WEAK ACIDS

A **strong acid** is the one which ionizes completely in water to give all its hydrogen ions. Examples of strong acids are hydrochloric acid (HCl), sulphuric acid (H_2SO_4) and nitiric acid (HNO_3), etc.

The ionization equations for the acids are shown below:

a. HCl (l)   + H_2O (l) $\longrightarrow$ H^+ (aq)   + Cl^- (aq)
b. H_2SO_4 (l) + H_2O (l) $\longrightarrow$ $2H^+$ (aq)   + SO_4^{2-} (aq)
c. HNO_3 (l) + H_2O (l) $\longrightarrow$ H^+ (aq)   + NO_3^- (aq)

A **weak acid** is the one which ionizes partially in water to produce few hydrogen ions. Examples of weak acids are ethanoic acid (CH_3COOH), carbonic acid (H_2CO_3) etc.

The ionisation equations for the acids are shown below:

a. CH_3COOH (l) $+$ H_2O (l) $\rightleftharpoons$ CH_3COO^- (aq) $+$ H_3O^+ (l)

b. H_2CO_3 (l) $+$ H_2O (l) $\rightleftharpoons$ HCO_3^- (aq) $+$ H_3O^+ (l)

The double arrow $\rightleftharpoons$ means that the reaction is reversible.

STRONG AND WEAK BASES

A **strong base** is the one which ionizes completely in water to produce more hydroxide ions. Examples of stong bases are sodium hydroxide and potassium hydroxide.

The ionization equations for the bases are:

a. $NaOH$ (s) $+$ H_2O (l) $\longrightarrow$ Na^+ (aq) $+$ OH^- (aq)

b. KOH (s) $+$ H_2O (l) $\longrightarrow$ K^+ (aq) $+$ OH^- (aq)
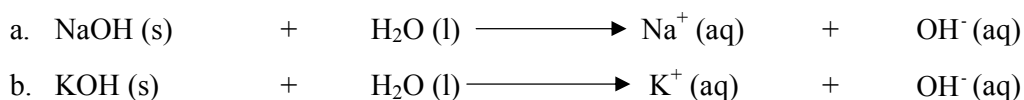
A **weak base** is the one which ionizes partially in water to produce few hydroxide ions. A typical example of a weak base is ammonia.

NH_3 (g) $+$ H_2O (l) $\rightleftharpoons$ NH_4^+ (aq) $+$ OH^- (aq)
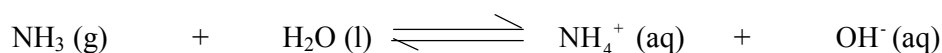
DETERMINING THE STRENGTH OF ACIDS AND BASES USING CONDUCITIVITY APPARATUS

Acids and bases are able to conduct electricity because they ionize when dissolved in water. However, the electrical conductivirty depends on the acid or base being used as electrolyte.

Stong bases and acids have higher electrical conductivity as compared to weak acids and bases. This is because stong acids and bases ionize completely releasing all the ions which conduct electricity when dissolvee in water.

On the other hand, weak acids and bases are weak electrolytes because they do not ionize completely when dissolve din water and therefore release few ions to conduct electricity.

COMPARING THE ELECTRICAL CONDUCTIVITY OF STRONG AND WEAK ACIDS AND BASES.

Materials: a beaker, 2 carbon rods, 2 cells, connecting wires, distilled water, ammeter, solutions of ethanoic acid, hydrochloric acid, sodium hydroxide and ammonia (of equal concentration)

Procedure:

a. Put 50 ml of ethanoic acid solution into the beaker.

b. Arrange the apparatus as shown below.
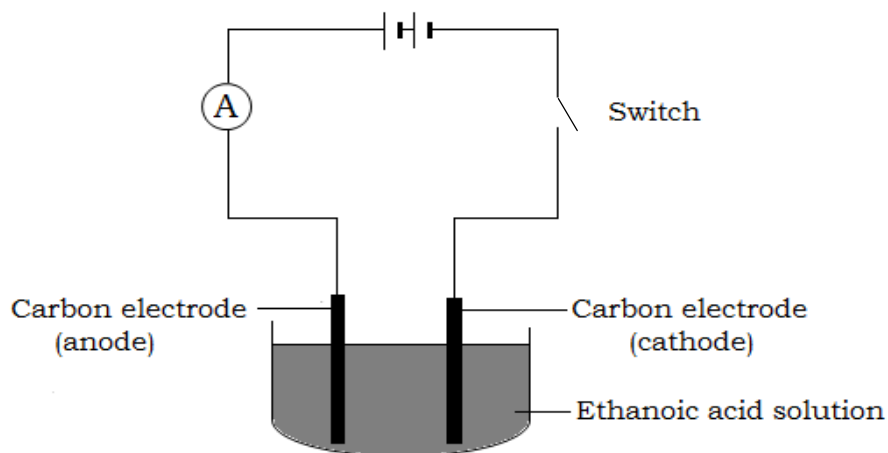
c. Close the switch and record the ammeter reading in the table of results.

d. Rinse the beaker and the carbon electrodes with the distilled water.

e. Repeat steps **a** to **d** using 50 ml of the solutions of hydrochloric acid, sodium hydroxide and ammonia respectively.

Table of results

Solution	Ammeter reading
Ethanoic acid	
Hydrochloric acid	
Sodium hydroxide	
Ammonia	

Conclusion

The acid or base whose current reading is higher is a strong acid or base, and the acid or base whose current reading is lower is a weak acid or base.

CONTROL VARIABLES DURING CONDUCTIVITY MEASUREMENTS

- Temperature of the solution
- Electrode surface area

- Electrode spacing
- Concentration of the solution

THE DIFFERENCE BETWEEN STRENGTH AND CONCENTRATION OF AN ACID OR A BASE

The strength of an acid or tells us how easily it ionizes to produce hydrogen ions while concentration of an acid indicates the proportion of water water and acid present in aqueous solution.

The strength of a base indicates how easily it ionizes to produce hydroxide ions while concentration tells us the proportion of water and base present in aqueous solution.

OXIDES

Oxides are compounds containing oxygen and another element.

TYPES OF OXIDES

Oxides can are classified as basic, acidic or amphoteric.

1. BASIC OXIDES
- These are formed when oxygen react with metals. Examples are magnesium oxide (MgO), iron (II) oxide (Fe_2O_3) and copper oxide (CuO) etc.
- These compounds are usually ionic in nature.
- Most of them dissolve in water to form alkalis. Others are insoluble in water.
- They can neutralize an acid.

2. ACIDIC OXIDES
- These are formed when oxygen react with non-metals. Examples are carbon dioxide (CO_2), sulphur dioxide (SO_2) and phosphorous pentaoxide (P_2O_5). They form carbonic acid, sulphurous acid and phosphoric acid.
- When dissolve in water, they form a solution that turns blue litmus paper red.

3. AMPHOTERIC OXIDES
- These oxides have both acidic and basic properties.
- They react with both bases and acids to form salt and water.
  Examples include aluminium oxide (Al_2O_3), zinc oxide (ZnO) and lead (II) oxide (PbO)

WAYS OF REGULATING pH IN DIFFERENT ENVIRONMENTS

The human digestive system, soils and water are some environments whose pH must be maintained for them to function efficiently.

a. THE HUMAN DIGESTIVE SYSTEM

The hydrochloric acid present in the stomach helps in protein digestion. If excess hydrochloric acid is produced, it causes the stomach to ache. To deal with this problem, an antacid containing magnesium hydroxide solution (milk of magnesia) is taken in order to neutralize the acid.

b. SOIL AND WATER ACIDITY

Acid levels in the soil increase when fertilizers are used for a long time. When acidic gases dissolve in rain water, they make water more acidic. High acidity in water and soils can be controlled by adding lime (calcium hydroxide).

WAYS OF PREPARING SALTS

A salt is product formed when the hydrogen ion of an acid is replaced by a metal or ammonium radical. Salts that do not dissolve in water are said to be soluble salts while those that do not dissolve in water are said to be insoluble salts. The method chosen for the preparation of a particular salt depends on the solubility of the salt in water. The methods of preparing salts include the following:
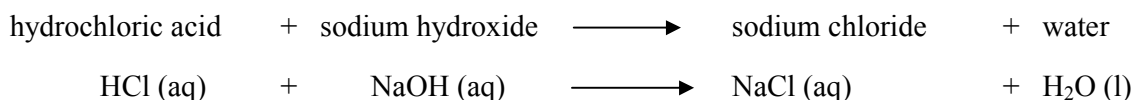
- Reaction of an acid with soluble bases.
- Reaction of acids with metals
- Reaction of acids with insoluble bases
- Reaction of acids with carbonates
- Precipitations

a. Reaction of acid with a soluble base (alkalis).

This method is used to prepare the salts of very reactive metals such as potassium or sodium. Titration is used in this reaction because both reactants are in solution. The acid is slowly and carefully added to a measured volume of alkali using a burette until the indicator, usually phenolphthalein, changes colour.
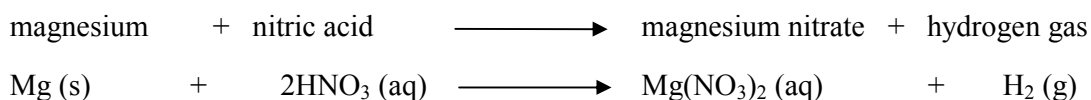
Once the end point is reached, the same volume of acid is added to the measured volume of alkali, but this time without the indicator. The solution which is produced can then be evaporated slowly to obtain pure dry crystals of the salt.

For example

hydrochloric acid   + sodium hydroxide ⟶   sodium chloride   + water

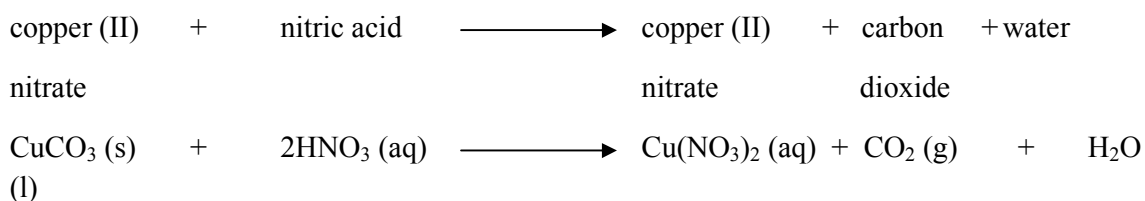HCl (aq)   +   NaOH (aq)   ⟶   NaCl (aq)   + H_2O (l)

b. Reaction of an acid with a metal

This method can only be used with less reactive metals such as magnesium, aluminium, zinc, iron and tin. Excess metal is added to dilute acid. The metal dissolves in then acid forming a salt solution. The salt solution formed is evaporated slowly to form a saturated solution of the salt. For example

magnesium   + nitric acid   ⟶   magnesium nitrate + hydrogen gas

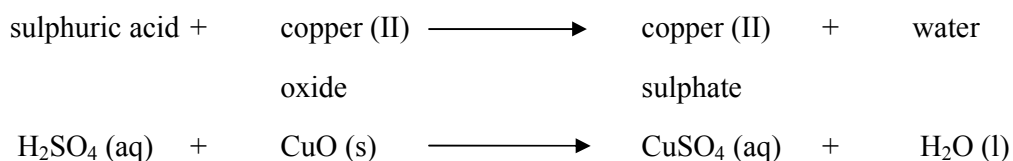Mg (s)   + $2HNO_3$ (aq)   ⟶ $Mg(NO_3)_2$ (aq)   + H_2 (g)

c. Reaction of acid with carbonate

This method is used with any metal carbonate and any acid, providing the salt is soluble. The procedure is similar to that of an acid and a metal. For example, copper (II) carbonate would be added in excess to dilute nitric acid to produce copper (II) nitrate solution.

copper (II)   +   nitric acid   ⟶   copper (II)   + carbon   +water

nitrate   nitrate   dioxide

$CuCO_3$ (s)   + $2HNO_3$ (aq)   ⟶ $Cu(NO_3)_2$ (aq) + CO_2 (g)   + H_2O (l)
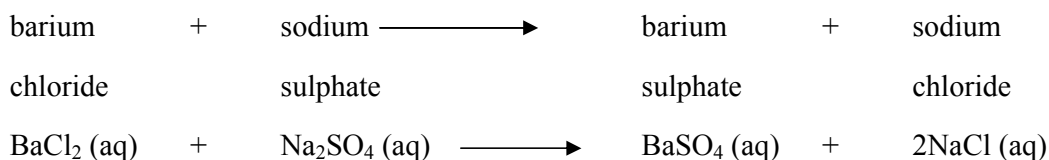
d. Reaction of an acid with an insoluble base

This method can be used to prepare a salt of an unreactive metal, such as lead or copper. In this case, it is possible to use direct reaction of the metal with an acid so the acid is neutralized using particular metal oxide. The reactants are warmed. A typical example is the neutralisation of sulphuric acid by copper (II) oxide to produce copper (II) sulphate.

sulphuric acid +   copper (II)   ⟶   copper (II)   +   water

oxide   sulphate

H_2SO_4 (aq)   +   CuO (s)   ⟶ $CuSO_4$ (aq)   + H_2O (l)

e. Precipitation

This method is used to prepare an insoluble salt. Solutions of the two chosen soluble solutions are mixed. The insoluble salt formed can be filtered off, washed with distilled water and dried. For example, to produce barium sulphate, barium chloride and sodium sulphate can be used.

barium + sodium $\longrightarrow$ barium + sodium

chloride   sulphate   sulphate   chloride

$BaCl_2 (aq)$ + $Na_2SO_4 (aq)$ $\longrightarrow$ $BaSO_4 (aq)$ + $2NaCl (aq)$

SOLUBILITY RULES

Solubility rules guide the choice of a method to be used when preparing salts.

- All potassium (K), sodium (Na) and ammonium (NH_4) are soluble.
- All nitrates and hydrogen carbonates (HCO_3) are soluble.
- All sulphates are soluble except lead sulphate ($PbSO_4$) and barium sulphate ($BaSO_4$). Calcium sulphate ($CaSO_4$) is sparingly soluble.
- All chlorides are soluble except silver chloride (AgCl). Lead chloride ($PbCl_2$) is only soluble in hot water.
- All carbonates are insoluble except potassium, sodium and ammonium carbonates.

APPPLICATIONS OF PRECIPITATIONS REACTIONS

a. Treatment of water for domestic use

Sodium carbonate can be used to treat water for domestic use. The toxic cations are precipitated as insoluble lead carbonate or mercury carbonates which are then filtered off.

Sodium carbonate is also used to make hard water soft by precipitating the ions that cause water hardness.

b. Treatment of industrial effluents

Effluents refer to liquid wastes discharged from a sewage system, a factory, a nuclear powers station or an industrial plant.

When treating industrial effluents, precipitation is carried out to remove toxic ions in the water solution. Ions to be removed are reacted with other ions to produce insoluble solids that can be removed by sedimentation or filtration.
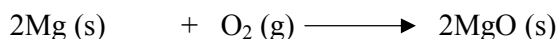
TOPIC 3 : OXIDATION AND REDUCTION

1. OXIDATION

a. It is the addition of oxygen to a substance.

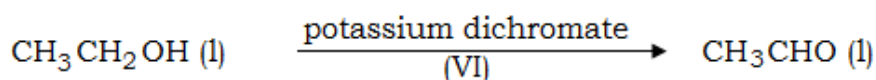Consider the reaction between magnesium (Mg) metal and oxygen.

magnesium + oxygen $\longrightarrow$ magnesium oxide

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

The magnesium has gained oxygen. We say that it has been oxidized.

b. It is removal of hydrogen from a substance.

For example, ethanol can be changed to ethanal by an oxidising agent potassium dichromate (VI), in the presence of an acid.

$$CH_3CH_2OH\ (l) \xrightarrow{\text{potassium dichromate (VI)}} CH_3CHO\ (l)$$

The ethanol molecule has lost two hydrogen atoms to become ethanal. It has been oxidized.

c. It is the loss of electrons by a substance.

Consider a reaction between magnesium and oxygen to form magnesium oxide.

$2Mg\ (s)\ +\ O_2\ (g) \longrightarrow 2MgO\ (s)$
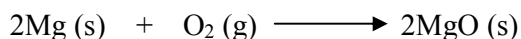
In this reaction, each magnesium atom loses two electrons. Therefore magnesium has been oxidized.

The oxidation reaction equation will be written as:

$Mg \longrightarrow Mg^{2+}\ +\ 2e^-$
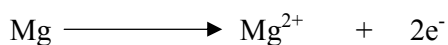
During oxidation, electrons are lost, so they are written on the side of products of the equation.

d. It is the increase in charge or oxidation number.

Consider the reaction between sodium and chlorine to form sodium chloride.

$2Na\ (s)\quad +\quad Cl_2\ (g) \longrightarrow 2NaCl\ (s)$
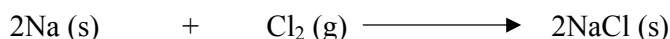
Each sodium atom gives one electron to the chlorine atom and increases its charge form (0) to (+1) according to the equation below.
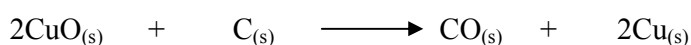
$$Na \longrightarrow Na^+ + e^-$$
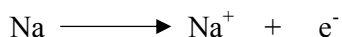
It is also correct to write: $Na - e^- \longrightarrow Na^+$

Since sodium has increased its charge from (0) to (+1), it has been oxidized.

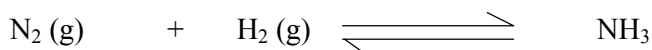## REDUCTION

a. It is the removal of oxygen from a substance.

For example when copper oxide is heated with carbon, the following reaction takes place.

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

Copper oxide is losing oxygen. It is being reduced.

b. It is the addition of hydrogen to a substance.

For example during the chemical equation for the formation of ammonia gas is shown below.

$$N_2 (g) + H_2 (g) \rightleftharpoons NH_3$$

Nitrogen is gaining hydrogen. It is being reduced.

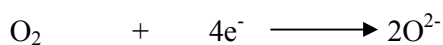**c. It is the gaining of electrons by a substance.**

Consider a reaction between magnesium and oxygen to form magnesium oxide.

$$2Mg (s) + O_2 (g) \longrightarrow 2MgO (s)$$
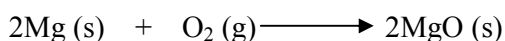
In this reaction, each oxygen atom gains two electrons. Therefore, oxygen is being reduced.

The reduction reaction equation will be written as:

$$O_2 + 4e^- \longrightarrow 2O^{2-}$$

During reduction, electrons are gained, so they are written on the reactants part of the equation.

d. It is the decrease in charge or oxidation number.

Consider the reaction between aluminium and chlorine to form aluminium chloride.

$$2Li (s) + Cl_2 (g) \longrightarrow 2LiCl (s)$$
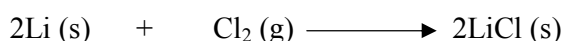
During the reaction, each chlorine atom receives one electron from the lithium atom and decreases its charge according to the equation below.

$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$
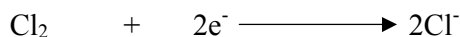
The charge on chlorine has decreased from (0) to (-1). Therefore, chlorine has been reduced

OXIDATION NUMBERS

An oxidation number is a whole number with a positive or negative sign of the charge on an atom, ion or a compound. It is also called the charge on an atom, ion or a compound.

It indicates the amount of electrons lost or gained by an element during a chemical reaction. For example Al^{3+} has an oxidation number of (+3) and O^{2-} has an oxidation number of (-2).

ASSIGNING OXIDATION NUMBERS

While it is now clear that oxidation numbers can be positive or negative, it is important to note that the calculation of such numbers follows a set of rules. The rules are:

1. The oxidation number of an uncombined element is equal to zero. For example the oxidation number of carbon, $C = 0$, $Cl_2 = 0$, $Fe = 0$ and $H_2 = 0$, etc.
2. The oxidation number of each hydrogen atom, H is +1.
3. The oxidation number of each oxygen atom, O is -2.
4. The algebraic sum of oxidation numbers of the elements in a compound is equal to zero. For example, $H_2O = 0$ means that:
   $2(H) + 1(O) = 2(+1) + 1(-2) = 2 + (-2) = 0$.
5. The oxidation number of monatomic ion is equal to the charge on the ion. For example, the oxidation numbers of ions of group I elements is +1, group II is +2, group III is +3, group VII is -1, etc. That is why the oxidation number of $Na^+ = +1$, $Al^{3+} = +3$, $Cl^- = -1$, $P^{3-} = -3$, etc.
6. For a polyatomic ion, the sum of the separate charges inside the ion must be equal to the final charge on the ion itself. For example, the oxidation number of NO_3^- is -1.

Example 1

Work out the oxidation number of sodium (Na) and chlorine (Cl) in sodium chloride (NaCl).

Solution

It means that Na + Cl = 0. (Rule number 4)

But Na is in group I and Cl is in group VII of the periodic table.

Therefore the oxidation numbers of Na $\quad$ = $\quad$ +1 and Cl $\quad$ = $\quad$ -1

Example 2

Work out the oxidation number of S in SO_4^{2-}.

Solution

SO_4^{2-} is a polyatomic ion with a charge of -2. The oxidation number of each oxygen atom, O = -2. Let the oxidation number of sulphur be y.

So, $\qquad\qquad\qquad\qquad$ y + 4(O) $\qquad$ = $\qquad$ -2 $\qquad\qquad$ (Rule number 6)

$\qquad\qquad\qquad\qquad\qquad\quad$ y + 4(-2) $\qquad$ = $\qquad$ -2

$\qquad\qquad\qquad\qquad\qquad\quad$ y – 8 $\qquad\quad$ = $\qquad$ -2

$\qquad\qquad\qquad\qquad\qquad\quad$ y $\qquad\qquad\quad$ = $\qquad$ +6

Therefore the oxidation number of S in SO_4^{2-} is +6.

Example 3

In the equation below, work out the oxidation number of manganese (Mn) in MnO_4^- and state if it is being reduced or oxidized.

$$MnO_4^- \quad + \quad 8H^+ \longrightarrow Mn^{2+} \quad + 4H_2O$$

Solution

MnO_4^- is a polyatomic ion with a charge of -1 and the oxidation number of each oxygen atom is -2.

Let the oxidation number of manganese (Mn) = z

So, $\qquad\qquad\qquad\qquad$ z + 4(O) $\qquad$ = $\qquad$ -1 $\qquad\qquad$ (Rule number 6)

$\qquad\qquad\qquad\qquad\qquad\quad$ z + 4(-2) $\qquad$ = $\qquad$ -1

$\qquad\qquad\qquad\qquad\qquad\quad$ z - 8 $\qquad\quad$ = $\qquad$ -1

$$z \quad = \quad -1 + 8$$

$$= \quad +7$$

Therefore the oxidation number of Mn is +7. The oxidation number of Mn decreases from (+7) to (+2), hence manganese (Mn) is being reduced.

HALF EQUATIONS IN REDOX REACTIONS

Consider the equation for the reaction between sodium (Na) and chlorine (Cl)

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

In this reaction the following has happened:

- Each sodium atom has donated one electron to chlorine atom. Therefore sodium has been oxidized according to the equation below.

  $$Na \ - \ e^- \longrightarrow Na^+ \ \dots\dots\dots\dots\dots\dots\dots\dots\dots\text{[oxidation reaction]}$$

- Each chlorine atom has received one electron from sodium atom. Therefore, chlorine has been reduced according to the equation below.

  $$Cl_2 \ + \ 2e^- \longrightarrow 2Cl^- \ \dots\dots\dots\dots\dots\dots\dots\dots\text{[reduction reaction]}$$

Reduction of chlorine is taking place side by side with the oxidation of sodium. Since reduction and oxidation occur simultaneously, this reaction is called **red**uction-**ox**idation reaction or in short **redox reaction**. Each of the reactions is called a **half-reaction** because the electrons released by one reaction are accepted by another in order to complete the redox reaction.
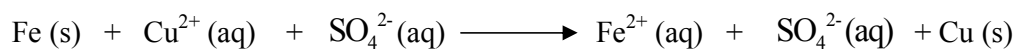
SPECTATOR IONS AND NET IONIC EQUATIONS

Spectator ion: This is an ion which does not take part in a chemical reaction.

Ionic equation: This is an equation in which the ions which are involved in a reaction are shown. The equation could be an overall reaction or half reaction. In a net ionic equation spectator ions are excluded.

The equation for the reaction between iron (Fe) and copper sulphate ($CuSO_4$) solution is shown below.

$$Fe \ (s) \ + \ CuSO_4 \ (aq) \longrightarrow Cu \ (s) \ + \ FeSO_4 \ (aq)$$
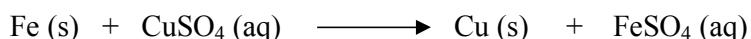
Writing the equation in terms of the ions present in solution we get:

$$Fe \text{ (s)} + Cu^{2+} \text{ (aq)} + SO_4^{2-} \text{ (aq)} \longrightarrow Fe^{2+} \text{ (aq)} + SO_4^{2-} \text{ (aq)} + Cu \text{ (s)}$$

In this equation, the sulphate ion, SO_4^{2-} (aq) is found on both sides of the equation and nothing has changed on it. This means the SO_4^{2-} ion did not take part in the reaction. It is the spectator ion. The ionic equation for the reaction between iron (Fe) and copper sulphate should therefore be written as follows:

$$Fe \text{ (s)} + Cu^{2+} \text{ (aq)} \longrightarrow Fe^{2+} \text{ (aq)} + Cu \text{ (s)}$$
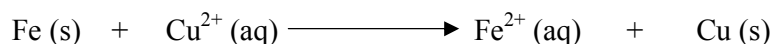
BALANCING REDOX EQUATIONS

There are two rules for writing redox equations

1. In any ionic equation, the charges must balance. This means the sum of charges on the left-hand side (LHS), must be equal to the sum of charges on the right-hand side (RHS) of the equation.

This rule applies to both overall (full) reaction and half-reaction equations. Examples are

Equation	Charges on LHS	Charges on RHS
$Zn \text{ (s)} + Pb^{2+} \text{ (aq)} \longrightarrow Zn^{2+} \text{ (aq)} + Pb \text{ (s)}$	+2	+2
$Br_2 \text{ (aq)} + 2e^- \longrightarrow 2 Br^- \text{ (aq)}$	-2	-2
$Na \text{ (s)} \longrightarrow Na^+ \text{ (aq)} + e^-$	0	0

2. The number of electrons lost in the oxidation half reaction must be equal to the number of electrons gained in the reduction half reaction.

For example, in the reaction between iron (Fe) and copper sulphate ($CuSO_4$) solution, the number of electrons lost in the oxidation half reaction is two according to the equation below:

$$Fe(s) \longrightarrow Fe^{2+} \text{ (aq)} + 2e^- \dots\dots\dots\dots\dots\dots\dots\text{[oxidation half equation]}$$

Therefore, the number of electrons gained in the reduction half reaction must also be two.

$$Cu^{2+} \text{ (aq)} + 2e^- \longrightarrow Cu \text{ (s)}\dots\dots\dots\dots\dots\dots\dots \text{[reduction half equation]}$$

WRITING OVERALL EQUATIONS FROM TWO HALF REACTION EQUATIONS

To write an overall equation from two half reaction equations, the steps to be followed are:

1. Balance the two half equations separately.
2. Combine the two half equations by adding using the rules of algebra. The electrons should cancel each other out.

Example 1

The following are half equations for the reaction between magnesium (Mg) and copper sulphate ($CuSO_4$) solution.

$$Mg^o(s) \longrightarrow Mg^{2+}(aq) + 2e^-$$
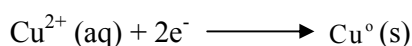
$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu^o(s)$$

Write a full equation for the reaction.

Solution

The oxidation half reaction equation is: $Mg^o(s) \longrightarrow Mg^{2+} + 2e^-$

            Charges   0   +2 + -2

The reduction half equation is: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu^o(s)$

            Charges 2+ + 2-   0

Each of the half equations is balanced. Each side has zero charge (Rule number 1). In the reaction, magnesium (Mg) loses two electrons in an oxidation reaction, which are gained by the copper ions (Cu^{2+}) in a reduction reaction (Rule number 2).
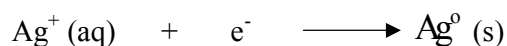
Adding the two half equations we get:

$$Mg^o(s) + Cu^{2+}(aq) + 2e^- \longrightarrow Mg^{2+}(aq) + Cu^o(s) + 2e^-$$
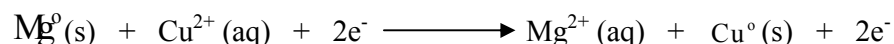
Cancelling out the electrons, the full equation is

$$Mg^o(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu^o(s)$$
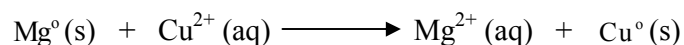
Example 2

The half equations for the reaction between zinc (Zn) and silver nitrate ($AgNO_3$) solution are:

$$Zn^0(s) - 2e^- \longrightarrow Zn^{2+}(aq)$$

$$Ag^+ (aq) \quad + \quad e^- \quad \longrightarrow \quad Ag^o \ (s)$$

Write an overall equation for the reaction.

Solution

Each side of the half equations has zero charge. Therefore each half equation is balanced. However, zinc (Zn) loses two electrons in an oxidation reaction but silver gains one electron in a reduction reaction.
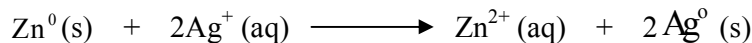
To equalize the number of electrons lost and gained, the reduction half equation must be doubled. Thus

$$Zn^0 (s) \ - \ 2e^- \longrightarrow \ Zn^{2+} (aq)\ldots\ldots\ldots\ldots\ldots\ldots\text{[oxidation half equation]}$$

$$2Ag^+ (aq) + 2e^- \longrightarrow 2Ag^o \ (s)\ldots\ldots\ldots\ldots\ldots\text{[reduction half equation]}$$

Adding the two half equations we get:

$$Zn^0 (s) \ - \ 2e^- \ + \ 2Ag^+ (aq) \ + \ 2e^- \longrightarrow Zn^{2+} (aq) \ + \ 2Ag^o \ (s)$$
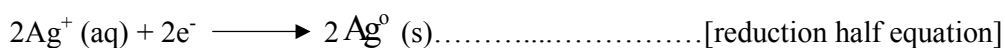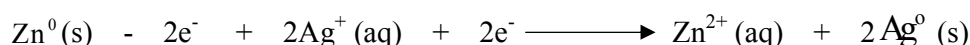
Cancelling out the electrons, the overall equation is:

$$Zn^0 (s) \ + \ 2Ag^+ (aq) \longrightarrow Zn^{2+} (aq) \ + \ 2Ag^o \ (s)$$
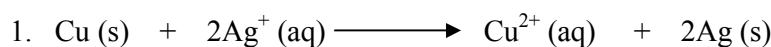
OXIDISING AGENTS AND REDUCING AGENTS

The terms **oxidising agent** and **reducing agent** are used to describe **reactants** in reduction–oxidation (redox) reactions.

An **oxidising agent** is a substance that causes oxidation to take place by accepting electrons from another substance. When accepting the electrons, the oxidising agent is itself reduced.

A **reducing agent** is a substance that causes reduction to take place by donating electrons to another substance. When donating the electrons, the reducing agent is itself oxidized.
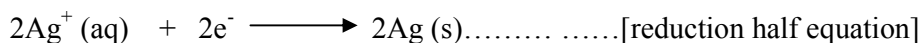
Examples

Write the oxidation and reduction half equations and identify the oxidising agent and the reducing agent in the following equations.

1. $Cu \ (s) \ + \ 2Ag^+ (aq) \longrightarrow Cu^{2+} (aq) \ + \ 2Ag \ (s)$

Solution

a. The half equations are:

$Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$………….....[oxidation half equation]

$2Ag^+_{(aq)} + 2e^- \longrightarrow 2Ag_{(s)}$……… ……[reduction half equation]
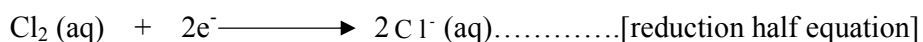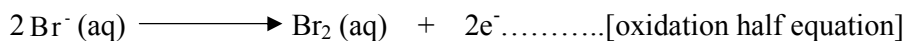
b. Reducing agent: $Cu_{(s)}$, it has increased its charge from 0 to +2

   Oxidising agent: $Ag^+_{(aq)}$, it has decreased its charge from +1 to 0.
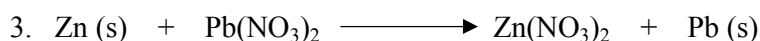
2. $Cl_{2(aq)} + 2Br^-_{(aq)} \longrightarrow 2Cl^-_{(aq)} + Br_{2(aq)}$

Solution

a. The half equations are:

$2Br^-_{(aq)} \longrightarrow Br_{2(aq)} + 2e^-$………..[oxidation half equation]

$Cl_{2(aq)} + 2e^- \longrightarrow 2Cl^-_{(aq)}$…………[reduction half equation]
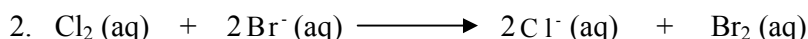
b. The reducing agent is $Br^-_{(aq)}$ while the oxidising agent is $Cl_{2(g)}$

3. $Zn_{(s)} + Pb(NO_3)_2 \longrightarrow Zn(NO_3)_2 + Pb_{(s)}$

Solution

a. The half equations are:

$Zn_{(s)} \longrightarrow Zn^{2+} + 2e^-$……………[oxidation half equation]

$Pb^{2+}_{(aq)} + 2e^- \longrightarrow Pb_{(s)}$……….…..[reduction half equation]
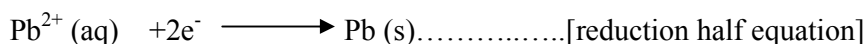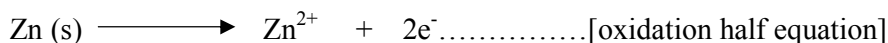
b. The reducing agent is $Zn_{(s)}$ while the oxidising agent is $Pb^{2+}_{(aq)}$

ORDER OF REACTIVITY OF METALS

The order of reactivity of metals can be based on three criteria: reactions of metals with water, reaction of metals with acid, and metals displacement reactions.

a. REACTIONS OF METALS WITH WATER

Some metals such as potassium, sodium, calcium, magnesium, zinc, copper, gold, silver and iron react with water to produce a basic solutions or oxides and hydrogen gas. When their reactions are compared, the order of reactivity of metals is generated as shown below:

Metal	Reaction	Order of reactivity
potassium (K)	very violent with cold water catches fire	most reactive
sodium (Na)	violent with cold water	
calcium (Ca)	less violent with cold water	
magnesium (Mg)	very slow with cold water but vigorous with steam	Increasing reactivity
zinc (Zn)	quite slow with steam	
iron (Fe)	slow with steam	
copper (Cu)	no reaction	
silver (Ag)	no reaction	
gold (Au)	no reaction	least reactive

b. REACTION WITH HYDROCHLORIC ACID

Most metals react with dilute acids to form a salt solution and hydrogen gas. The order of reactivity of metals with hydrochloric acid is shown below:

| Metal	Reaction with hydrochloric acid	Order of reactivity
Potassium (K)	explosively fast	most reactive
Sodium (Na)	explosively fast	
magnesium (Mg)	vigorous	
zinc (Zn)	quite slow	
iron (Fe)	slow	Increasing reactivity
lead (Pb)	slow, and only if the acid is concentrated	
copper (Cu)	no reaction even with concentrated acid	
silver (Ag)	no reaction	
gold (Au)	no reaction	least reactive

c. DISPLACEMENT REACTIONS OF METALS

A metal displacement reaction is the reaction in which one metal substitutes another metal from a solution. A metal displacing the other is in solid form while the metal being displaced is in aqueous state in solution.

Example

Iron (Fe) metal is placed in an aqueous solution of copper (II) sulphate [$CuSO_4$ (aq)] and the experiment is left for about 5 minutes. Observations are made and recorded.


Copper (II) sulphate solution
[$CuSO_4$(aq)]
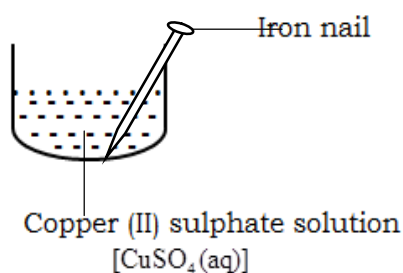
Observations

After 5 minutes, the following should have been observed

- Iron (nail) began to dissolve.
- A black-brown coating of copper metal formed on the nail.
- The colour of the solution began to turn from blue to pale green.

Explanation

When the iron (Fe) was placed in the solution, it started to dissolve and form iron ions (Fe^{2+}). This is oxidation reaction:

$$Fe\ (s) \longrightarrow Fe^{2+}\ (aq)\ +\ 2e^-$$
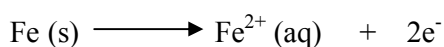
The copper ions (Cu^{2+}) present in the solution captured the free electrons. This is reduction reaction

$$Cu^{2+}\ (aq)\ +\ 2e^- \longrightarrow Cu\ (s)$$
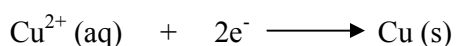
But since oxidation and reduction occur simultaneously, iron (Fe) goes into the solution as Fe^{2+} ions to take the place of copper ions [Cu^{2+} (aq)]. The copper ions (Cu^{2+}) come out of the solution as copper metal [$Cu^0(s)$] and coat on the iron nail.

Therefore, in this reaction iron is displaces copper. Since iron can displace copper, we say that iron (Fe) is more reactive than copper (Cu).

The colour of copper (II) sulphate is blue because of the presence of the copper ions. In this reaction, the copper ions are constantly being removed from solution to form copper metal. As time goes by, the concentration of copper ions decreases causing the change in the colour of solution from blue to pale green.

Displacement series

The displacement series is the arrangement of metals in order of their relative chemical reactivity. The displacement series is also called *activity series*, *reactivity series*, or *redox series*.

The activity series is as follows:

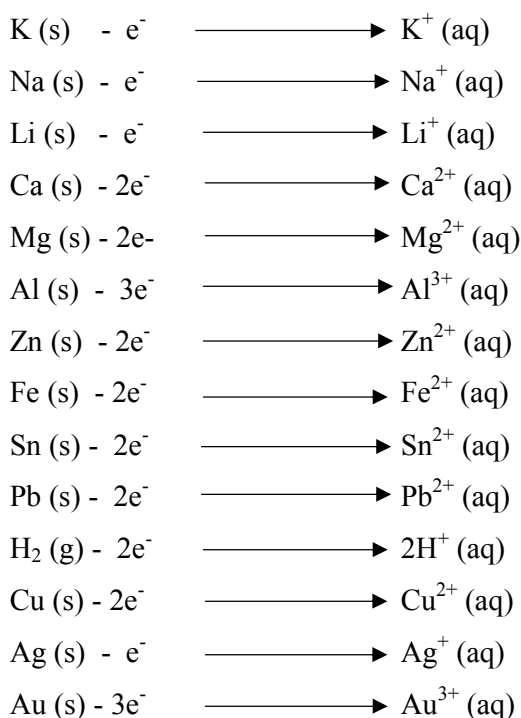| Most reactive (reducing) | potassium, K | |
	sodium, Na	
	lithium, Li	
	calcium, Ca	
	magnesium, Mg	
	aluminium, Al	
	zinc, Zn	
Increasing reactivity	iron, Fe	Decreasing reactivity
	tin, Sn	
	lead, Pb	
	hydrogen, H	
	copper, Cu	
	silver, Ag	
Least reactive (reducing)	gold, Au	

THINGS TO NOTE ABOUT THE ACTIVITY SERIES

1. The most reactive metal is at the top while the least reactive metal is at the bottom.
2. The metals at the top of the series can displace all other metals below them from solutions of their compounds by forcing them to accept electrons. For example, Potassium (K) can displace all the other metals below. Gold cannot displace any of the metals in the list.

3. The metals high on the series can easily donate electrons to metals lower down the series. For this reason, the metals at the top are said to be the most reducing or **most electropositive**.

THE ELECTROCHEMICAL SERIES

This is a list of half equations that shows the ability of one metal to donate electrons to other metals in solution. The series is as follows.

$$K (s) - e^- \longrightarrow K^+ (aq)$$
$$Na (s) - e^- \longrightarrow Na^+ (aq)$$
$$Li (s) - e^- \longrightarrow Li^+ (aq)$$
$$Ca (s) - 2e^- \longrightarrow Ca^{2+} (aq)$$
$$Mg (s) - 2e- \longrightarrow Mg^{2+} (aq)$$
$$Al (s) - 3e^- \longrightarrow Al^{3+} (aq)$$
$$Zn (s) - 2e^- \longrightarrow Zn^{2+} (aq)$$
$$Fe (s) - 2e^- \longrightarrow Fe^{2+} (aq)$$
$$Sn (s) - 2e^- \longrightarrow Sn^{2+} (aq)$$
$$Pb (s) - 2e^- \longrightarrow Pb^{2+} (aq)$$
$$H_2 (g) - 2e^- \longrightarrow 2H^+ (aq)$$
$$Cu (s) - 2e^- \longrightarrow Cu^{2+} (aq)$$
$$Ag (s) - e^- \longrightarrow Ag^+ (aq)$$
$$Au (s) - 3e^- \longrightarrow Au^{3+} (aq)$$

The working principle of the electrochemical series is the same as that of the displacement series. Atoms of an element above can displace ions of an element below. Atoms of an element below will not displace ions of an element above.

POSITIONS OF THE METALS IN THE ELECTROCHEMICAL SERIES

The position of the metals in the series is important in predicting a displacement reaction as well as the rate of reaction.

**a. Whether a displacement reaction will take place or not**

A reaction will take place if the atom is of an element which is high in the series while the ion is of an element which is lower in the series.

For example, a reaction takes place when zinc (Zn) metal is mixed with a solution of copper (II) sulphate.

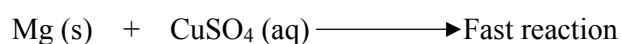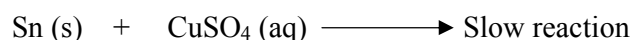$$Zn \text{ (s)} + CuSO_4 \text{ (aq)} \longrightarrow ZnSO_4 \text{ (aq)} + Cu \text{ (s)}$$
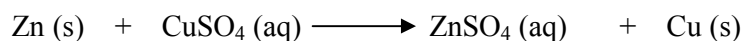
However, if the atom is of an element which is lower in the series while the ion is of an element high in the series, a reaction will not take place.

For example, a reaction does not take place if copper (Cu) metal is mixed with a solution of zinc sulphate ($ZnSO_4$).

$$Cu \text{ (s)} + ZnSO_4 \text{ (aq)} \longrightarrow No \text{ reaction}$$
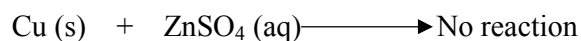
b. The rate of reaction

As the gaps between the elements become bigger, the reactions occur easily and fast. As the gaps decrease, the reactions occur slowly. For example

$$Sn \text{ (s)} + CuSO_4 \text{ (aq)} \longrightarrow Slow \text{ reaction}$$

$$Mg \text{ (s)} + CuSO_4 \text{ (aq)} \longrightarrow Fast \text{ reaction}$$

Experiment

Aim: To investigate the displacement reactions of metals

You are provided with four beakers, distilled water, a measuring cylinder, sand paper, and solutions of copper sulphate, zinc sulphate, iron sulphate and magnesium sulphate. You are also provided with pieces of copper, zinc, iron and magnesium metals.

a. Pour about 2 cm^3 of copper sulphate solution into each of the four beakers.

b. Clean the copper, zinc, iron and magnesium metals using sand paper.

c. Put a piece of each metal into each of the four beakers containing copper sulphate solution.

d. Record the in the table below by indicating "Reaction" or "No reaction".

e. Rinse the beakers with distilled water.

f. Repeat steps (a) to (e) using solutions of zinc sulphate, iron sulphate and magnesium sulphate respectively.

Metal salt solution	Metal			
	Copper	Zinc	Iron	Magnesium
Copper sulphate				
Zinc sulphate				
Iron sulphate				
Magnesium sulphate				

g. Use the results to arrange the metals in order of increasing reactivity.

HINT:

- The metal that reacts with all the given solutions is the most reactive.

- The metal that reacts with none of the given solutions is the least reactive.

- The metal that reacts with all the given solutions except the solution of the most reactive metal is the second-most reactive.

- The metal that reacts with all the solutions except the solutions of the first two most-reactive is the third-most reactive metal.

THE ELECTROCHEMICAL CELL

The electrochemical cell is a system in which metallic electrodes are placed into an electrolyte where a chemical reaction uses or generates electricity.

Electrochemical cells which generate electric current are called **voltaic** or **galvanic cells**.
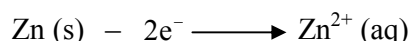
THE VOLTAIC CELLS

The voltaic cells consist of two separate compartments called half-cells containing electrolyte solutions and electrodes that can be connected in a circuit. Two dissimilar metals (e.g. copper and zinc) are dipped in an electrolyte. If the metals are connected by an external circuit, one metal is reduced while the other metal is oxidized.
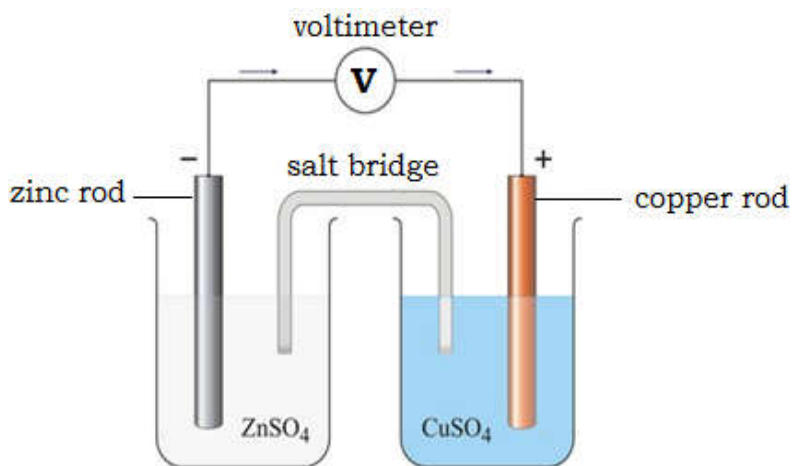

HOW THE VOLTAIC CELLS WORK

- Electrons flow from the more reactive metal (anode) to the less reactive metal (cathode). For example, from the left to the right in the figure above.

- Oxidation occurs at the anode. The name refers to the flow of anions towards it. For example, in the above voltaic cell zinc atoms give up electrons and go into the solution as ions. This is oxidation.

$$Zn\,(s)\; -\; 2e^- \longrightarrow Zn^{2+}\,(aq)$$

- Reduction occurs at the cathode. The name refers to the flow of cations towards it. For example, in the voltaic cell above copper (II) ions, Cu^{2+} (aq) accept electrons and copper is deposited. This is reduction.

$$Cu^{2+}\; +\; 2e^- \longrightarrow Cu\,(s)$$
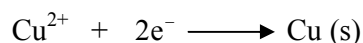
The overall reaction is as follows: $\quad Zn\,(s)\; +\; Cu^{2+} \longrightarrow Zn^{2+}\,(aq)\; +\; Cu\,(s)$

- The salt bridge acts as a transfer medium that allows ions to flow through out without allowing the different solutions to mix and react.
- Since electricity is defined as the flow of electrons, we can say that the redox reaction is giving out electrical energy.
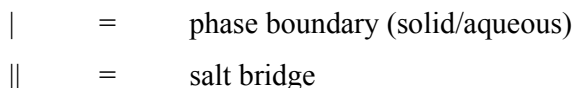
CELL NOTATION OF GALAVANIC CELLS

The standard notation used for all voltaic cells is:

anode|anode.electrolyte||cathode.electrolyte|cathode

For example $\qquad\qquad\qquad Zn(s)|Zn^{2+}\,(aq)||Cu^{2+}\,(aq)|Cu(s)$

where $\qquad\qquad\qquad$ | $\qquad$ = $\qquad$ phase boundary (solid/aqueous)

$\qquad\qquad\qquad\qquad\qquad$ || $\qquad$ = $\qquad$ salt bridge

STANDARD ELECTRODE POTENTIALS

The cell has an Electromotive Force (EMF) which is defined as maximum potential difference between two electrodes of a voltaic cell. EMF is the driving force of the cell reaction or any redox reaction i.e. the higher the EMF the stronger the reaction. The EMF is measured by a voltimeter, in volts.

Experiments show that each half reaction has its own special voltage, which is called the **electrode potential**; it is the potential difference between an electrode and the solution in which it is immersed.

The electrode potentials of metals are different due to differences in reactivity. Metals high in the electrochemical series tend to have high electrode potentials due to a greater ability to donate electrons, which also means a greater voltage.

Metals lower in the series are made to accept electrons from those higher up, causing them to have a lower voltage. This idea leads to the development of the **voltage series**.

THE VOLTAGE SERIES

The voltage is a list of standard electrode potentials for metals. It is written as follows:

$K (s) - 1e^- \longrightarrow K^+ (aq)$ -2.92 volts

$Na (s) - 1e^- \longrightarrow Na^+ (aq)$ -2.71 volts

$Mg (s) - 2e^- \longrightarrow Mg^{2+} (aq)$ -2.31 volts

$Al (s) - 3e^- \longrightarrow Al^{3+} (aq)$ -1.67 volts

$Zn (s) - 2e^- \longrightarrow Zn^{2+} (aq)$ -0.76 volts

$Fe (s) - 2e^- \longrightarrow Fe^{2+} (aq)$ -0.44 volts

$Sn (s) - 2e^- \longrightarrow Sn^{2+} (aq)$ -0.14 volts

$Pb (s) - 2e^- \longrightarrow Pb^{2+} (aq)$ -0.12 volts

$H_2 (g) - 2e^- \longrightarrow 2H^+ (aq)$ 0.00 volts

$Cu (s) - 2e^- \longrightarrow Cu^{2+} (aq)$ +0.34 volts

$Ag (s) - 1e^- \longrightarrow Ag^+ (aq)$ +0.80 volts

USING THE VOLTAGE SERIES

- The electromotive force (EMF) of the voltaic cell depends on the difference in the half cell voltages.

EMF = Oxidation electrode potential + Reduction electrode potential

       = $E_{reduction} + E_{oxidation}$

- Metals far apart in the voltage series would give a larger EMF, while metals close together would give a small EMF.
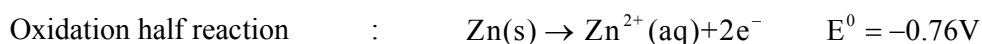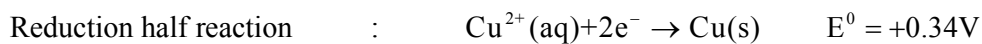
CALCULATING ELECTROMOTIVE FORCE (EMF) OF A CELL

- Write the half equation for each process.
- Look up the standard electrode potentials for each reaction.
- Reverse the sign on the oxidation electrode potential.
- Add the electrode potentials to get the EMF.
- A negative sign on the EMF indicates a spontaneous reaction will occur between the two metals. A negative sign indicates no reaction.

Example

Calculate the electromotive force of the electrochemical cell, $Zn(s)|Zn^{2+}(aq)||Cu^{2+}(aq)|Cu(s)$.

Refer to the voltage series.

Solution

Reduction half reaction   : $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $E^0 = +0.34V$

Oxidation half reaction   : $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $E^0 = -0.76V$

$$EMF \quad = \quad E_{reduction} + E_{oxidation}$$
$$= \quad +0.34V - (-0.76)\ V$$
$$= \quad +1.10\ V$$

CORROSION AND RUSTING

**Corrosion** is the reaction which occurs when some metals react with atmospheric oxygen and water.

Corrosion is an example of a displacement reaction which results in degradation of metals by dissolving and wearing away. During corrosion, the metals get oxidized to form oxides. For example,
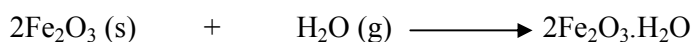
$$4Fe\ (s)\ +\ 3O_2\ (g) \longrightarrow 2Fe_2O_3\ (s)$$

Rusting means the oxidation of iron. Rust is the name of the iron oxide compound formed when iron corrodes (i.e. reacts with oxygen and water).

HOW RUSTING OCCURS

- Iron reacts oxygen to form (brown) iron (III) oxide.

  $$4Fe\ (s)\ +\ 3O_2\ (g) \longrightarrow 2Fe_2O_3\ (s)$$
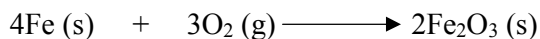
- At the same time, the iron (III) oxide reacts with water to form brown hydrated iron (III) oxide.

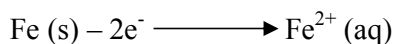 $$2Fe_2O_3\ (s)\ +\ H_2O\ (g) \longrightarrow 2Fe_2O_3.H_2O$$

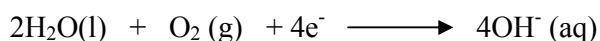 Iron (III) oxide   water vapour   hydrated iron (III) oxide

The hydrated iron (III) oxide is known as rust. Rust is a soft, crumbly brown solid.

RUSTING AS A REDOX REACTION

Firstly, iron dissolves in water and ionizes according to the equation:

$$Fe\ (s) - 2e^- \longrightarrow Fe^{2+}\ (aq)$$

Water and air capture the electrons and are reduced according to the equation:

$$2H_2O(l)\ +\ O_2\ (g)\ +\ 4e^- \longrightarrow 4OH^-\ (aq)$$

The hydroxyl ions react with the iron, Fe^{2+} (aq) ions and more water to form rust. The net ionic equation for the reaction is:

$$2Fe^{2+} (aq) + O_2 (g) \longrightarrow 4Fe^{3+} (aq) + 2O^{2-} (aq)$$

CONDITIONS FOR RUSTING

There are two conditions necessary for rusting to take place.
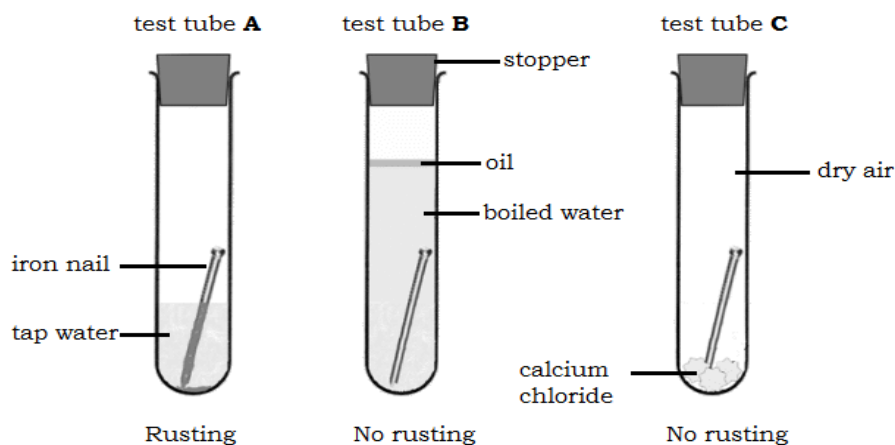
- Presence of oxygen
- Presence of water

If either oxygen or water is not present, then rusting will not take place.

INVESTIGATING CONDITIONS FOR RUSTING

Materials: test tubes, test tube rack, iron nails, cooking oil, anhydrous calcium chloride, tap water

Procedure:

a. The apparatus is set up as shown below.

b. The test tubes are left undisturbed for one week.

Expected results

The nail in test tube **A** will rust while those in test tubes **B** and **C** will not rust.

Explanation

In test tube **A**, rusting occurs because of the presence of both oxygen and water. The tap water contains air, hence oxygen. All the necessary conditions for rusting are present in test tube **A**. In test tube **B**, rusting will not occur because the boiled water does not contain dissolved oxygen. The oil prevents any atmospheric oxygen from entering. So, only water is

present. In test tube **C**, the anhydrous calcium chloride is used to remove water. Hence only air is present. As a result, rusting will not occur.

PREVENTION OF RUSTING

Rusting can be prevented using the following methods.

- Painting. It prevents water and oxygen coming into contact with metal.

- Greasing and oiling. It prevents water and oxygen from interacting with the metal.

- Alloying. An alloy is a mixture of metals. For example, stainless steel is made by mixing iron with chromium, nickel and manganese. The alloy is corrosion resistant.

- Galvanizing. This involves covering or coating iron with a molten layer of zinc. The zinc reacts with water and oxygen in preference in preference to iron.
  Even if the zinc is scratched, the iron does not rust. This is because zinc is higher in the activity series. Zinc coated iron is called galvanized iron. In galvanizing, zinc is being sacrificed to protect iron. Hence this type of protection is called **sacrificial protection**.

- Coating iron with plastic.

- Electroplating. This is the process of coating a metal with a thin layer of another metal by using electrolysis.
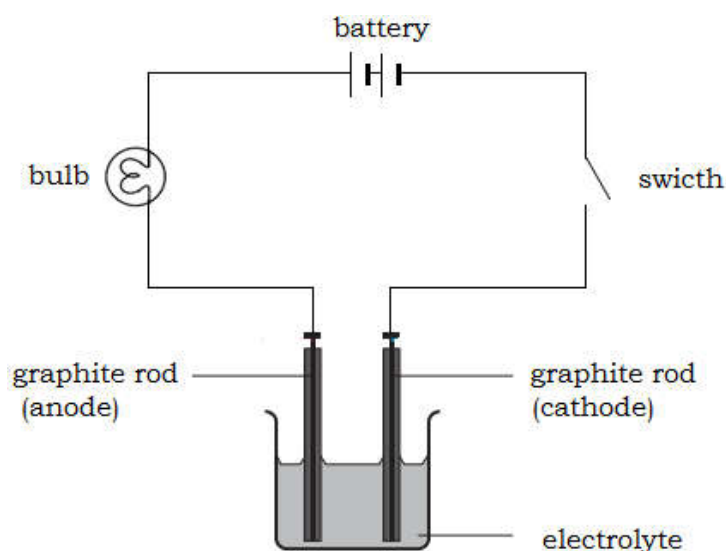
TOPIC 4 : ELECTROLYSIS

Electrolysis is the decomposition of a compound which is in molten or aqueous state using electrical energy. The process of electrolysis is carried out in an electrochemical cell. Electrolysis works for

- ionic compounds that are melted or dissolved in water, so that the ions are free to move.
- covalent compounds that form ions when they are dissolved in water.

COMPONENTS OF AN ELECTROCHEMICAL CELL

The electrochemical cell consists of a dry cell, an ammeter, bulb, graphite rods and an electrolyte.

- An **electrolyte** is a compound which when melted or when dissolved in water can conduct electricity and get decomposed by it.
- The graphite rods are called **electrodes**. The electrode attached to the positive terminal of the battery is called the **anode**.
- The electrode connected to the negative terminal of the cell is called the **cathode**.
- Dry cells are used to supply direct current.
- The bulb or ammeter is used to indicate that current is flowing in the circuit.

ELECTROLYSIS OF MOLTEN IONIC COMPOUNDS

Consider the electrolysis of molten lead bromide in the figure below
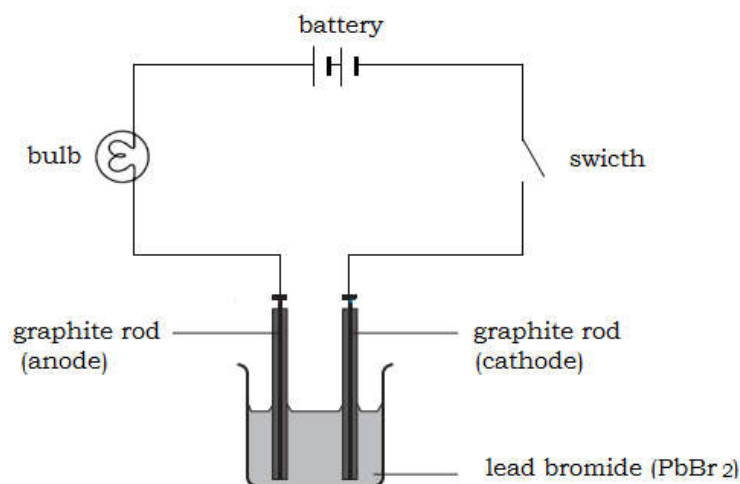
When the switch is closed this what happens:

- Electrons move from the negative terminal of the battery to the cathode.
- In the liquid, ions move to the electrode of opposite charge. Lead ions (Pb^{2+}) are attracted to the cathode while bromide ions (Br^-) are attracted to the anode. The moving ions carry electric current.
- At the cathode, lead ions each receive two electrons and are reduced according to the equation: Pb^{2+} (l) $+ 2e^- \longrightarrow$ Pb (l)

  Lead collects on the electrode and eventually drops off it.
- At the anode, bromide ions each gives up an electron and are oxidised according to the equation: $2Br^-$ (l) $- 2e^- \longrightarrow Br_2$ (g)
- The bromide gas bubbles off. Combining the anode and cathode reactions we get a redox reaction.

Pb^{2+} (l) $+$ $2Br^-$ (l) $\longrightarrow$ Pb (l) $+$ Br_2 (g)

So it must be noted that electrolysis decomposes a molten ionic compound. A metal is obtained at the cathode, and a non-metal at the anode.

ELECTROLYSIS OF AQUEOUS IONIC COMPOUNDS

When a salt such as sodium chloride (NaCl) dissolves in water its ions become free to move. So it can be electrolysed. However, the products are usually different from the products of the electrolysis of molten ionic compounds. This is so because the water itself also produces ions; hydrogen ions (H^+) and hydroxide ions (OH^-). A tiny proportion of its molecules split into ions.

$$\text{some water molecules} \longrightarrow \text{hydrogen ions} + \text{hydroxide ions}$$

$$H_2O \ (l) \longrightarrow H^+(aq) + OH^-(aq)$$
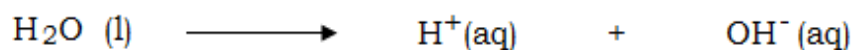
So in the aqueous solution, the H^+ and OH^- ions are from water compete with other ions to receive or give up electrons. To predict which ion will win the following rules apply during the electrolysis of aqueous ionic compound.

1. **At the cathode, either a metal or hydrogen is formed**.

- If a metal is more reactive than hydrogen, its ions stay in solution. The H^+ ions accept electrons, and hydrogen bubbles off.

- If the metal is less reactive than hydrogen, its ions will accept the electrons. The metal forms, leaving the H^+ ions in solution.

2. **At the anode, a non-metal other than hydrogen is formed**.

- If a halide ion (Cl^-, Br^-, I^-) is present in sufficient concentration, it will give up electrons more readily than the OH^- ions do. Molecules of chlorine, bromine and iodine.

- If no halide ion is present or if the halide solution is dilute, the OH^- ions will give up electrons and oxygen will be formed.

In either case, the ions are discharged selectively since only at each electrode only one type of ions gets discharged. This is called **preferential discharge**.

DEMOSTRATING PREFERENTIAL DISCHARGE

a. **Electrolysis of brine (**concentrated solution of sodium chloride, NaCl (aq)).
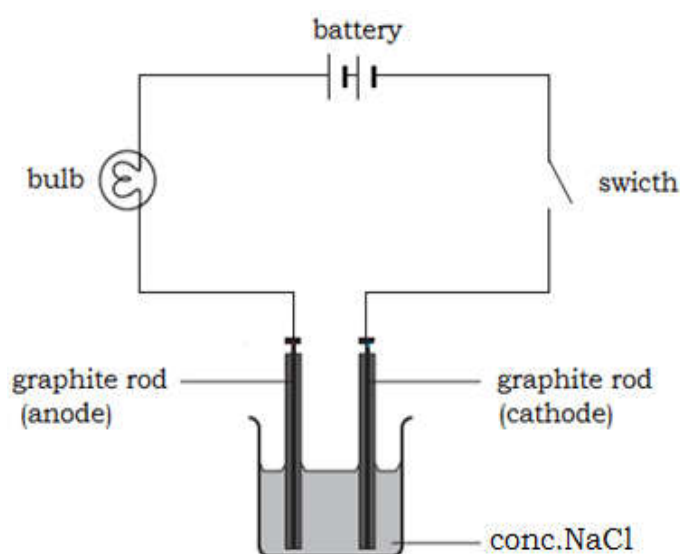
The ions present in solution are Na^+, Cl^- from the salt, and H^+ and OH^- from the water. The figure below shows the reactions at each electrode as the electrolysis proceeds.

- At the cathode, the H^+ accepts electrons since hydrogen is less reactive than sodium. The equation for the reaction is:

  $2H^+ (aq) + 2e^- \longrightarrow H_2 (g)$. This is **reduction**. The hydrogen gas bubbles off.
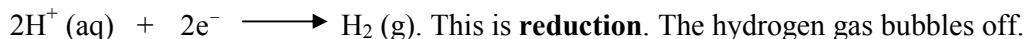
- At the anode, the Cl^- ions give up electrons more readily than the OH^- ions do. The equation for the reaction is:

  $2Cl^- (aq) - 2e^- \longrightarrow Cl_2 (g)$. This is **oxidation**. The chlorine gas bubbles off.
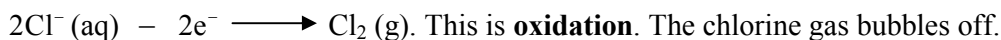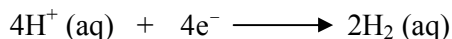
When the hydrogen and chlorine gases bubble off, Na^+ ions and OH^- ions remain behind, so a solution of sodium hydroxide is formed.

b. Electrolysis of dilute sodium chloride

The same ions are present as before. But now the proportion of Na^+ and Cl^- ions is lower since the solution is dilute.
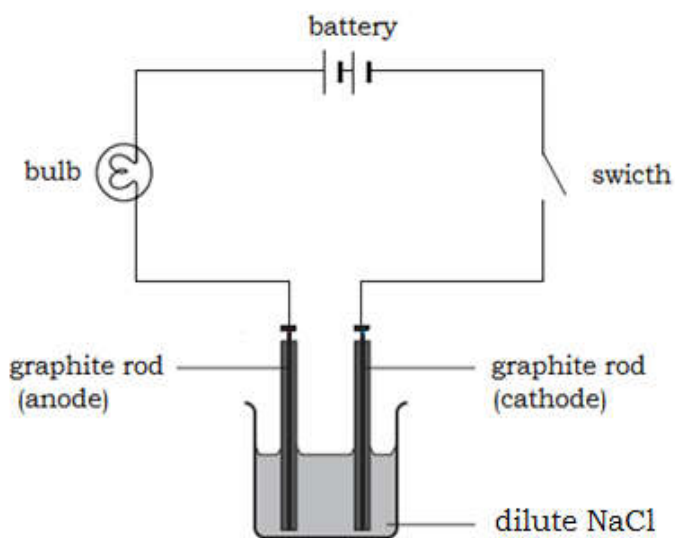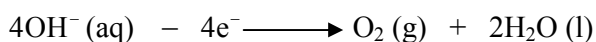

- At the cathode, hydrogen ions gain electrons according to the equation:

  $4H^+ (aq) + 4e^- \longrightarrow 2H_2 (aq)$

  This is **reduction**. (The 4 electrons are there to balance the half-equation at the anode)

- At the anode, the OH^- ions give up electrons, since the Cl^- ions are in low concentration. Oxygen gas bubbles off according to the equation:

 $4OH^- (aq) - 4e^- \longrightarrow O_2 (g) + 2H_2O (l)$

Instead of the dilute sodium chloride being decomposed, it is the water that gets decomposed. In a sense this is electrolysis of water.

**EFFECTS OF ELECTRODES ON ELECTROLYSIS**

Electrolysis of a solution will give different results depending on electrodes. A comparison is shown in the following **a** and **b**.

a. Electrolysis of copper sulphate aqueous copper (II) sulphate using inert electrodes.

Inert electrodes are the electrodes which allow passage of electric current into and out of electrolyte without themselves getting decomposed. Examples are carbon graphite and platinum.
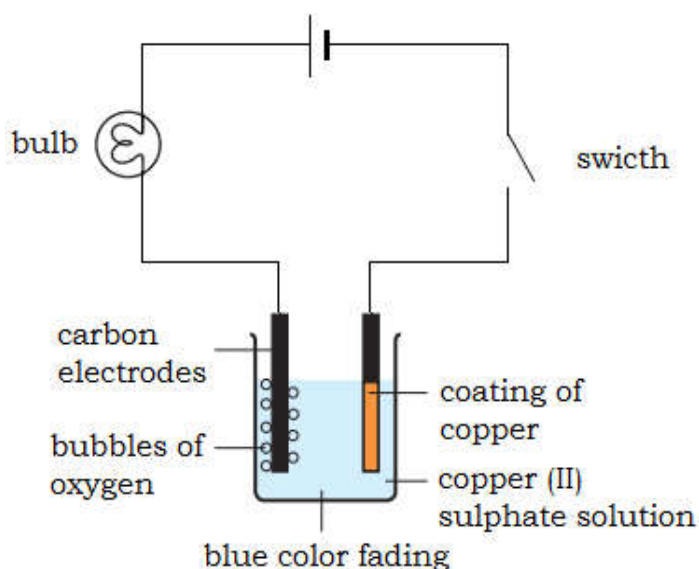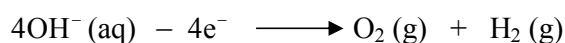

- At the cathode, copper ions get discharged.

  $$2Cu^{2+} (aq) + 4e^- \longrightarrow 2Cu (s)$$
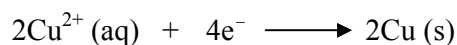
  The copper coats the electrode

- At the anode, oxygen bubbles off.

  $$4OH^- (aq) - 4e^- \longrightarrow O_2 (g) + H_2 (g)$$
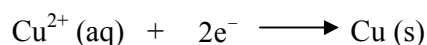
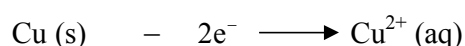The blue colour of copper sulphate solution fades as the copper ions are discharged.

b. Electrolysis of copper aqueous copper (II) sulphate using copper electrodes
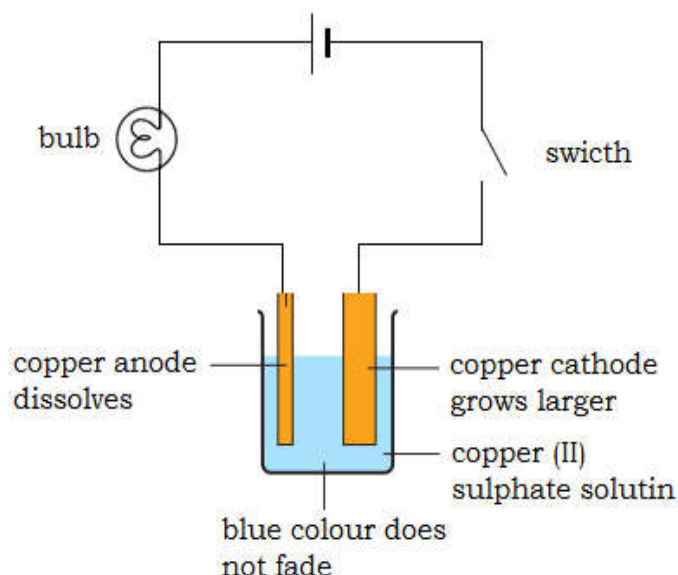
- At the cathode, copper is formed and coats the electrode.

$$Cu^{2+} (aq) + 2e^- \longrightarrow Cu (s)$$

- The anode dissolves, giving copper ions.

$$Cu (s) - 2e^- \longrightarrow Cu^{2+} (aq)$$

The copper ions move to the cathode to form copper. The colour of the solution does not fade.

WRITING HALF-EQUATIONS FOR ELECTRODE REACTIONS

To write the correct half-equations for electrode equations, follow the following steps:

a. Name the ions present and the products. Write each half-equation correctly.
b. Remember, positive ions go to the cathode and negative ions go to the anode.
c. Write the correct symbol of the element that forms. For example, Cl_2 (not Cl).
d. The number of electrons in the equation should be the same as the total charge on the ions in it.
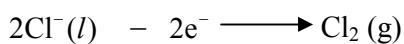
Example

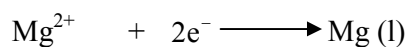Write the anode and cathode half-equations for the electrolysis of molten magnesium chloride.

Solution

The ions present in solution are magnesium ions, Mg^{2+} and chloride ions, Cl^- ions.
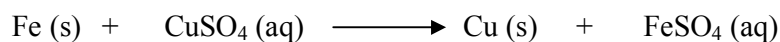
- At the cathode

$$2Cl^- (l) \quad - \quad 2e^- \longrightarrow Cl_2 (g)$$

- At the anode

$$Mg^{2+} \quad + \quad 2e^- \longrightarrow Mg (l)$$

USES OF ELECTROLYSIS

1. EXTRACTION OF METALS

Less reactive metals can be extracted from their ores by spontaneous redox reactions. The metal is extracted by dipping the more reactive metal into an aqueous solution of the metal to be extracted. For example, to extract copper (Cu) from its compound, an iron metal can be dipped into the aqueous solution of a compound of copper such as copper (II) sulphate. The reaction between iron and aqueous copper sulphate will be:

$$Fe (s) + \quad CuSO_4 (aq) \longrightarrow Cu (s) + \quad FeSO_4 (aq)$$
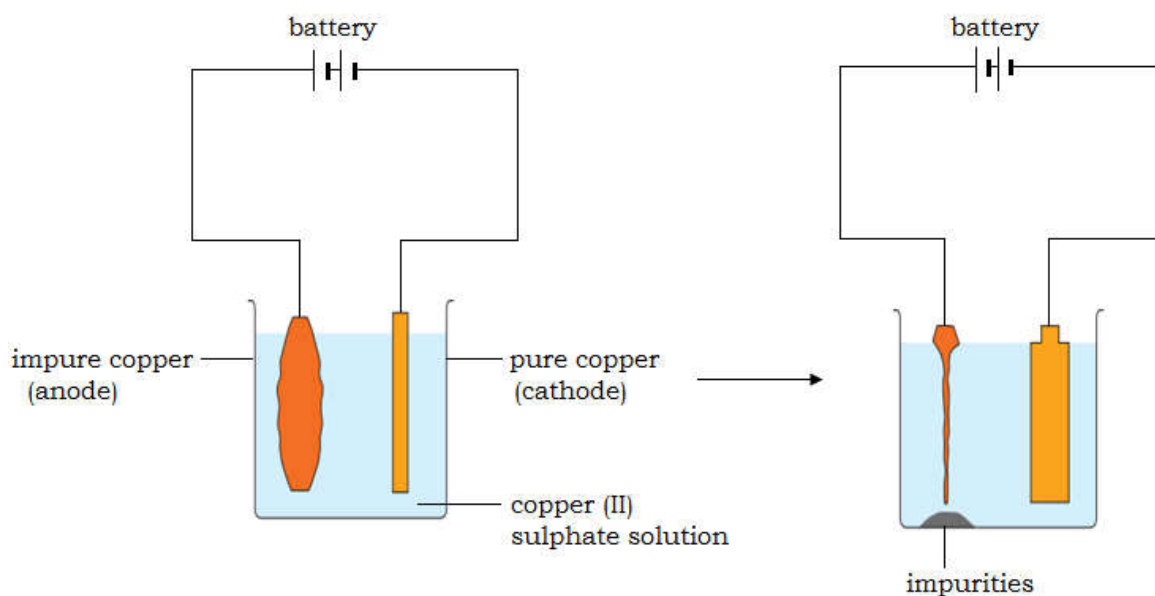
2. PURIFICATION OF METALS

To purify a metal, the mineral ore is dissolved to form an electrolyte and current is passed through it. The metal is extracted as a deposit on the cathode.

PURIFICATION OF COPPER

The apparatus is connected as shown below.

The anode is made of impure copper. The cathode is pure copper. The electrolyte is dilute copper (II) sulphate solution.

The copper in the anode dissolves into the solution. This is oxidation reaction. The half reaction equation at the anode is: $Cu\,(s) \longrightarrow Cu^{2+}\,(aq) + 2e^-$

The copper ions are reduced and deposited on the pure copper cathode. The half reaction equation at the cathode is: $Cu^{2+} + 2e^- \longrightarrow Cu\,(s)$

The impurities do not dissolve; they just drop to the bottom of the container.

3. ELECTROPLATING

Electroplating is the process of coating a metal with a thin layer of another metal by using electrolysis.

To electroplate a metal, the procedure is as follows.

- First, the metal to be electroplated is thoroughly cleaned.
- An electrolyte whose ions are required to be deposited is selected.
- Direct current (dc) is used instead of alternating current (ac).
- The metal to be electroplated must be made the cathode. Metals for coating must be made the anode so that when they dissolve into the electrolyte, they replace the ions which are being removed from the container.

ELECTROPLATING OF IRON WITH COPPER

The apparatus is set up as shown below
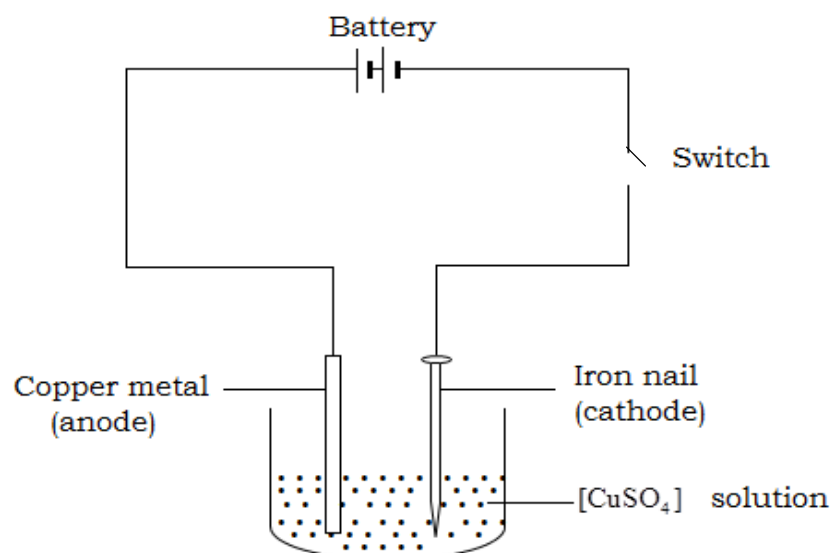
The switch is closed and the experiment is observed for about 10 minutes.

Observations are made and recorded.

Observations

- The iron nail will be coated with copper metal.

- The copper metal in the anode will be eaten up.

Explanation

When the switch is closed, the copper metal in the anode dissolves and is oxidised as follows:

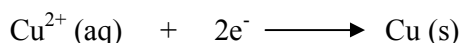$Cu \ (s) \longrightarrow Cu^{2+} \ (aq) \ + \ 2e^{-}$

The copper ions (Cu^{2+}) are attracted to the cathode. At the cathode, the copper ions are gain electrons and are reduced to copper atoms.

$Cu^{2+} \ (aq) \ + \ 2e^{-} \longrightarrow Cu \ (s)$

The copper metal coats the iron metal. The iron metal is said to be electroplated.

APPLICATIONS OF ELECTROPLATING AND ITS BENEFITS TO SOCIETY

The process of electroplating is used:

- To make objects look attractive. For example most necklaces and wristwatches are plated with silver to make them look shiny and beautiful.

- To protect metals from corrosion. For example, steel car bumpers are coated with chromium to protect them from reacting with water.

- To increase the strength or size of the metal. If a metal bar is small, it can be made bigger by adding a layer on it through electroplating.

- To extract and purify metals from ores e.g. copper.

TOPIC 5   :   ISOMERISM

Isomerism is the existence of compounds with the same molecular formula but different structural formulae.

ISOMERS

Isomers are the compounds having the same molecular formula but different structural formulae. Examples are

C_4H_{10} C_4H_{10}

```
  H  H  H  H                        H   H H
  |  |  |  |                        |   | |
H-C--C--C--C-H                   H-C--C--C-H
  |  |  |  |                        |   | |
  H  H  H  H                        H   | H
                                   H-C-H
                                       |
                                       H
```
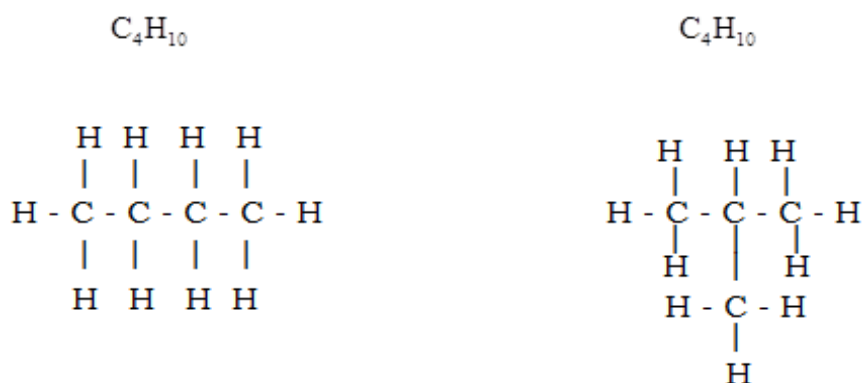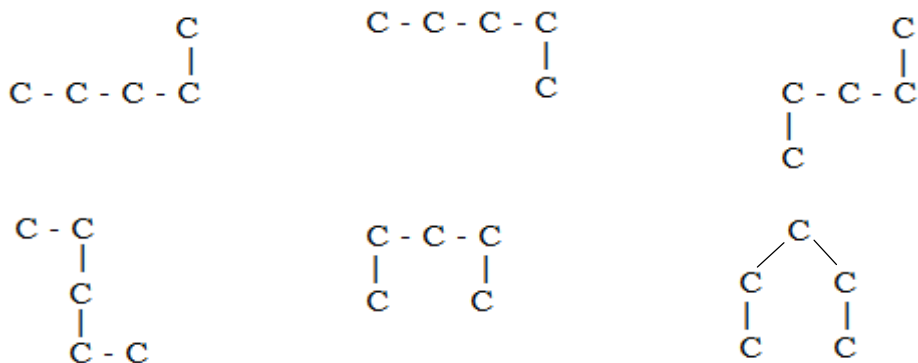
CONFORMERS

Conformations are compounds that differ from each other by the rotation of a single bond in the molecule.

```
          C              C-C-C-C                    C
          |                 |                       |
  C-C-C-C                    C                   C-C-C
                                                    |
                                                    C

  C-C              C-C-C                         C
    |              |   |                       C   C
    C              C   C                        |   |
    |                                           C   C
    C-C
```

All the compounds above have the same molecular formulae and same structural formulae, but they have different shapes due to bending, turning, or twisting of the carbon – carbon bonds. Bending, turning or twisting of the C – C bonds in the molecule does not change the structure of the molecule.

IDENTIFYING ISOMERS OR CONFORMERS

In order to decide whether given structures are isomers or conformation of each other, it is useful to check if there are any branches or junctions in the molecules the molecules have simply been turned, twisted or bent.
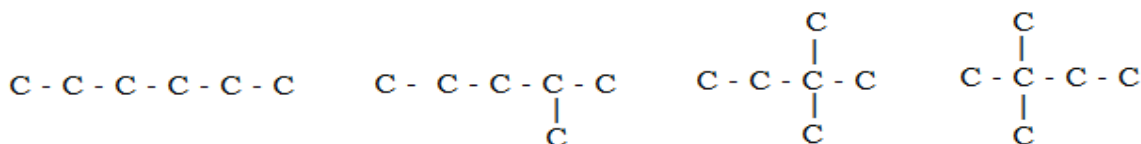
To do this properly, the following steps may assist.

a. Working out the carbon skeleton first. The carbon skeleton enables us to see important point about the structure of the molecule.

b. Choosing and numbering the longest chain of carbon atoms in the molecule. This chain acts like a reference line inside the molecule. It is called the *parent chain*. The parent chain allows us to locate accurately the number and position of the junctions in a molecule.

c. Naming the compounds. Isomers have the same molecular formulae but different names while conformations have the same molecular formulae and names.

For example, the following carbon structures represent conformations, because there are no junctions in the molecules.

C - C - C - C - C - C   C - C - C - C - C   C - C
                                               | |
                                               C   C - C - C - C
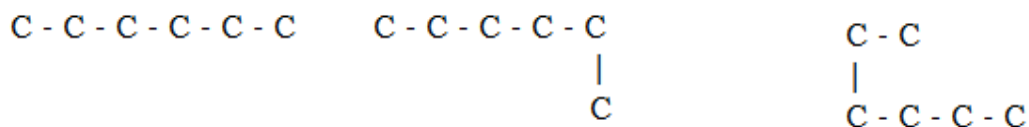
All of the conformations above have the same name: normal hexane (n – hexane).

```
                                                            C                    C
                                                            |                    |
C - C - C - C - C - C     C - C - C - C - C     C - C - C - C     C - C - C - C
                                       |                    |                    |
                                       C                    C                    C
```

The above compounds will have different names because they have different structural formulae even though they have the same molecular.

TYPES OF STRUCTURAL ISOMERISM

There two types of structural isomerism; chain and position isomerism.

a. Chain isomerism

This is the isomerism due to the differences in the way the carbon atoms are connected to each other in the molecule. For example

```
        C                             C
        |                             |
C = C - C - C                 C = C - C - C
```
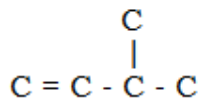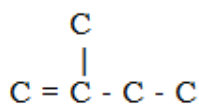
b. Position isomerism

This is the isomerism due to the difference in the position of a functional group in the molecule. For example

```
C = C - C - C - C             C - C = C - C - C
```
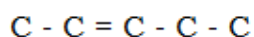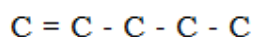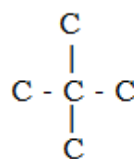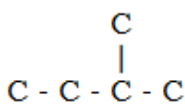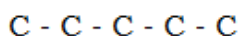
NAMING ISOMERS

Isomers are named systematically. The rules for naming the isomers of compounds vary from one homologous series depending on whether the compounds in the series show chain isomerism or position isomerism or both.
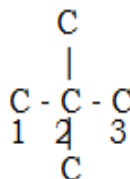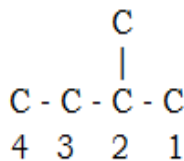
RULES FOR NAMING ISOMERS OF ALKANES

Alkanes do not have a functional group, so they only show chain isomerism. The rules for naming the isomers of alkanes are as follows:

a. Write down the carbon skeleton for each molecule. Using C_5H_{12}, the carbon skeletons are:

```
                              C                     C
                              |                     |
C - C - C - C - C     C - C - C - C         C - C - C
                                                    |
                                                    C
```

b. Number the parent chain in the carbon skeleton. Numbering starts the end of the parent chain nearest the first branch.

```
                              C                     C
                              |                     |
C - C - C - C - C     C - C - C - C         C - C - C
1   2   3   4   5     4   3   2   1         1   2   3
                                                    C
```
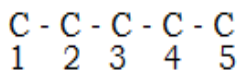
c. Count up all the carbon in the parent chain and use the number to give a basic name to the compound.

$$C - C - C - C - C$$
$$1 \quad 2 \quad 3 \quad 4 \quad 5$$

pentane
(5 C atoms)

$$
\begin{array}{c}
\text{C} \\
| \\
\text{C - C - C - C} \\
4 \quad 3 \quad 2 \quad 1
\end{array}
$$

butane
(4 C atoms)

$$
\begin{array}{c}
\text{C} \\
| \\
\text{C - C - C} \\
1 \quad 2 \quad 3 \\
| \\
\text{C}
\end{array}
$$
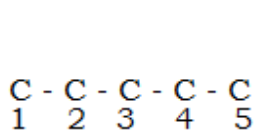
propane
(3 C atoms)

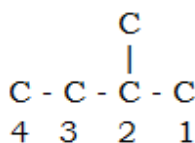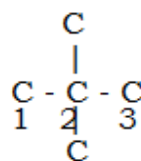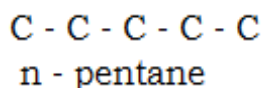d. Name the group that is attached to the parent chain. This group is called an **alkyl group**.

Alkyl group	Name of the alkyl group
CH_3-	Methyl
CH_2CH_3	Ethyl
$CH_2CH_2CH_3$	Propyl
$CH_2CH_2CH_2CH_3$	Butyl

For straight chain single bonded carbon – carbon isomers, there is no branched group. Such isomers start with the prefix 'n-' which stands for normal.

$$C - C - C - C - C$$
n - pentane

e. State the positions of the carbon atom on the parent chain where the branch (es) is/are occurring and give a final name to the compound. The number is written before the alkyl name. For example

$$
\begin{array}{c}
\text{C} \\
| \\
\text{C - C - C - C} \\
4 \quad 3 \quad 2 \quad 1
\end{array}
$$

2 - methyl butane
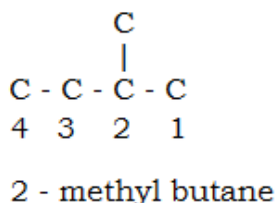
If more than one alkyl group is attached to the parent chain, the numbers indicating the positions of the branches are also written before the alkyl group. The numbers of the alkyl

groups having the same number of carbon atoms are indicated by prefixes di-, tri-, or tetra- to mean two, three or four respectively. For example

$$\begin{array}{c} \text{C} \\ | \\ \text{C} - \text{C} - \text{C} \\ 1 \quad 2 \quad 3 \\ | \\ \text{C} \end{array}$$

2, 2 - dimethyl propane

f.   Name the compound fully.

$$\begin{array}{c} \text{H} \quad \text{H} \ \text{H} \ \text{H} \ \text{H} \\ | \quad | \ | \ | \ | \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{H} \\ | \quad | \ | \ | \ | \\ \text{H} \quad \text{H} \ \text{H} \ \text{H} \ \text{H} \end{array}$$
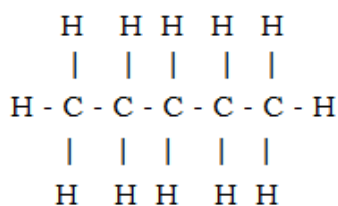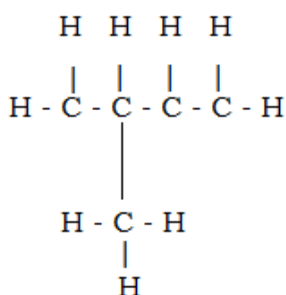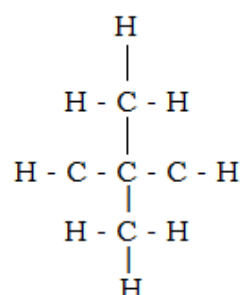
n-pentane

$$\begin{array}{c} \text{H} \ \text{H} \ \text{H} \ \text{H} \\ | \ | \ | \ | \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{H} \\ | \\ \text{H} - \text{C} - \text{H} \\ | \\ \text{H} \end{array}$$

2 - methyl butane

$$\begin{array}{c} \text{H} \\ | \\ \text{H} - \text{C} - \text{H} \\ | \\ \text{H} - \text{C} - \text{C} - \text{H} \\ | \\ \text{H} - \text{C} - \text{H} \\ | \\ \text{H} \end{array}$$
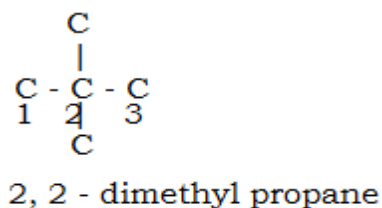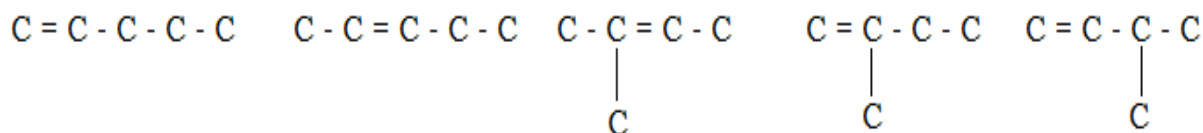
2, 2 - dimethyl propane

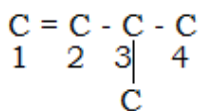**RULES FOR NAMING ISOMERS OF ALKENES**

Alkenes show both chain and position isomerism due to the presence of the functional group, the carbon – carbon double bond in their structures. The procedure for naming isomers of alkenes is as follows:

a.   Write the carbon skeleton using the number of carbon atoms given in the molecular formula. Place the functional group and branches in correct positions.

Let us consider the compound whose molecular formula is C_5H_{10}.

$$\text{C} = \text{C} - \text{C} - \text{C} - \text{C} \qquad \text{C} - \text{C} = \text{C} - \text{C} - \text{C} \qquad \begin{array}{c} \text{C} - \text{C} = \text{C} - \text{C} \\ | \\ \text{C} \end{array} \qquad \begin{array}{c} \text{C} = \text{C} - \text{C} - \text{C} \\ | \\ \text{C} \end{array} \qquad \begin{array}{c} \text{C} = \text{C} - \text{C} - \text{C} \\ | \\ \text{C} \end{array}$$

b.   Choose the parent chain and number it. Numbering starts at the end of the parent chain nearest the functional group.

C = C - C - C - C   C - C = C - C - C   C - C = C - C
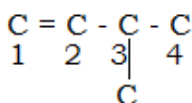1   2   3   4   5   1   2   3   4   5   1   2   3   4
                                                                   C

C = C - C - C   C = C - C - C
1   2| 3   4   1   2   3| 4
    C   C

c. Count up the number of carbon atoms in the parent chain and give a basic name to the compound.

C = C - C - C - C   C - C = C - C - C   C - C = C - C
1   2   3   4   5   1   2   3   4   5   1   2   3   4
                                                                   C

Pentene (5 C atoms)   Pentene (5 C atoms)   Butene (4 C atoms)
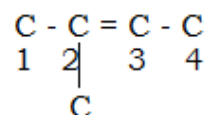
C = C - C - C   C = C - C - C
1   2| 3   4   1   2   3| 4
    C   C

Butene (4 C atoms)   Butene (4 C atoms)
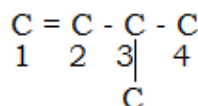
d. Show the position of the functional group by the lowest number of the carbon atom holding the functional group. Put it before –ene. For example pent-1-ene or hex-2-ene.

C = C - C - C - C   C - C = C - C - C   C - C = C - C
1   2   3   4   5   1   2   3   4   5   1   2   3   4
                                                                    C

Pent -1 - ene   Pent - 2 - ene   But - 2 - ene

C = C - C - C   C = C - C - C
1   2| 3   4   1   2   3| 4
    C   C

But -1 - ene   But - 1 - ene

e. State the position and type of the alkyl group attached. Name the compound fully.

```
      H  H  H  H  H              H  H  H  H  H             H           H
      |  |  |  |  |              |  |  |  |  |             |           |
  H - C = C - C - C - C - H   H - C - C = C - C - C - H   H - C - C = C - C - H
      |  |  |                      |        |  |             |         | |
      H  H  H                      H        H  H             H        H H
                                                                      |
                                                            H - C - H
     Pent - 1 - ene           Pent - 2 - ene                         |
                                                                      H

                                                        2 - methyl but - 2 - ene
```
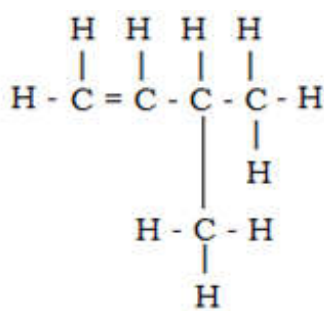
H H H
| | |
H - C = C - C - C - H
| |
H H

H - C - H
|
H

2 - methyl but - 1 - ene

H H H H
| | | |
H - C = C - C - C - H
|
H

H - C - H
|
H
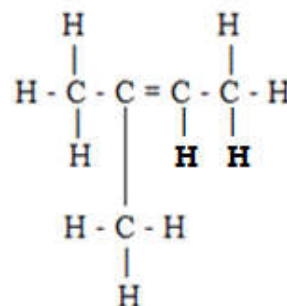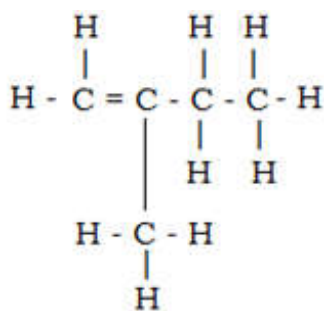
3 - methyl but -1 - ene

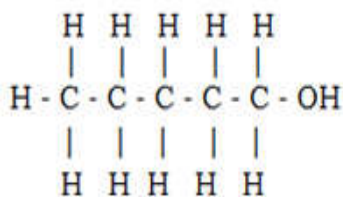## RULES FOR NAMING ISOMERS OF ALKANOLS

Alkanols contain the –OH functional group. For this reason, alkanols, just like the alkenes, show both chain and position isomerism.

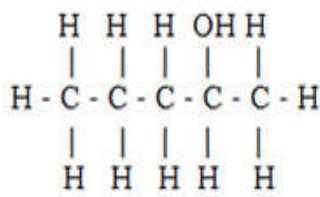The rules for naming isomers of alkanols are as follows:

a. Draw a straight chain carbon skeleton using the number of carbon atoms and give its name.
b. Place the functional group and branches in correct positions.
c. Number the parent chain. It is numbered in such a way that the carbon atom to which the –OH group is attached has the lowest number.
d. Indicate the position of the functional group using the number of the carbon atom to which the –OH group is attached. Put it before –ol. For example butan-2-ol.
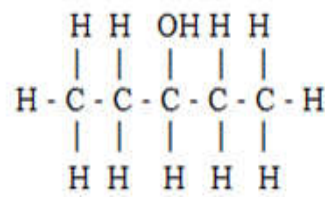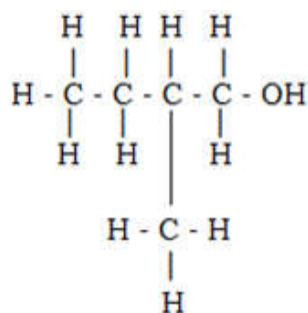
Example

Write down and name all the isomers of the pentanol ($C_5H_{11}OH$).

H H H H H
| | | | |
H - C - C - C - C - C - OH
| | | | |
H H H H H

Pentan -1- ol

H H H OH H
| | | | |
H - C - C - C - C - C - H
| | | | |
H H H H H

Pentan -2- ol

H H OH H H
| | | | |
H - C - C - C - C - C - H
| | | | |
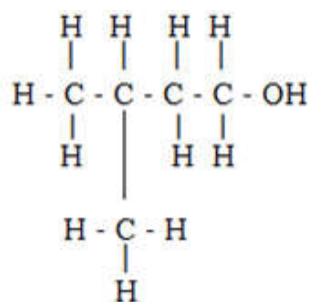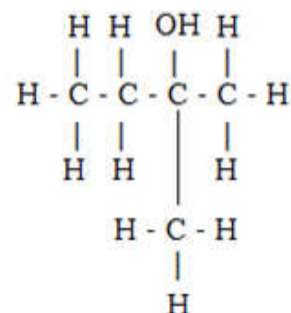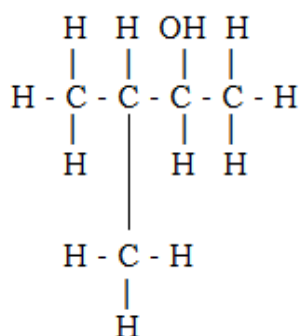H H H H H

Pentan -3- ol

2 - methyl butan -1- ol

3 - methyl butan -1- ol

2 - methyl butan -2- ol

3 - methyl butan -2- ol
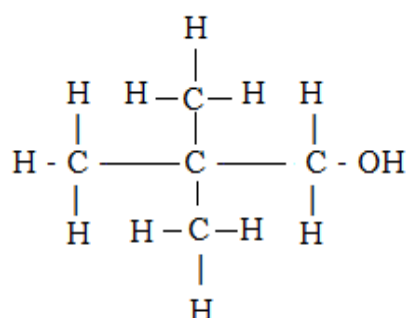
2, 2 - dimethyl propan -1- ol

RULES FOR NAMING ISOMERS OF ALKANALS

Alkanals are compounds with a functional group called the carbonyl group, **-CO**. Alkanals only exhibit chain isomerism. They do not show position isomerism because their functional group is always placed at the end of the alkanal molecule.

Example 1:

Write down all the structural isomers of butanal (C_3H_7CHO)

Solution

Butanal
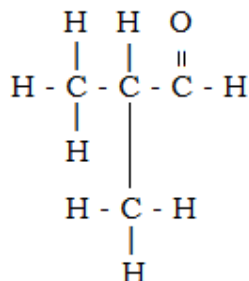
2 - methylpropanal

Example 2:

Write down and name all the structural isomers of pentanal (C_4H_9CHO).

Solution:

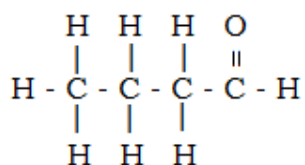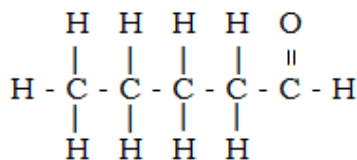Pentanal
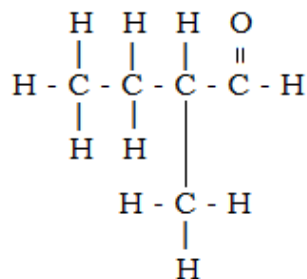
2 - methylbutanal
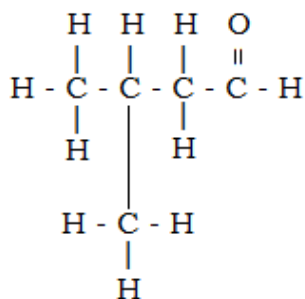
3 - methylbutanal
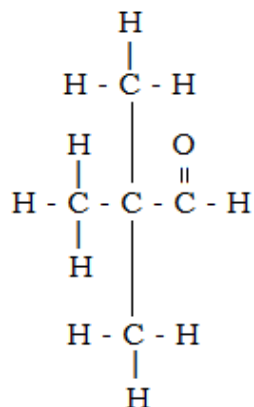
2,2 - dimethlypropanal

RULES FOR NAMING ISOMERS OF ALKANONES

Alkanones have a functional group similar to that of alkanals, the carboxyl group, **-CO**. Alkanones exhibit both chain and position isomerism. This is because the carboxyl group in alkanones can be shifted to different positions in the molecule.

The position of the functional group is indicated by the number of the carbon atom in the carboxyl group. It is put before –one. For example, butan-2-one, pentan-3-one, etc.

Example

Write down and name all the structural isomers of pentanone $(C_5H_{10}O)$.

Solution:
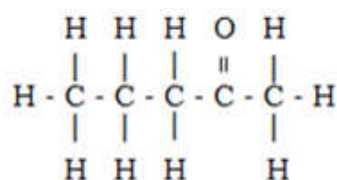
Pentan-2-one
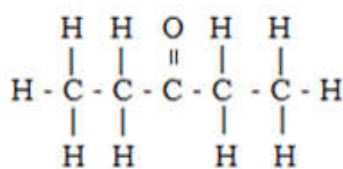
Pentan-3-one
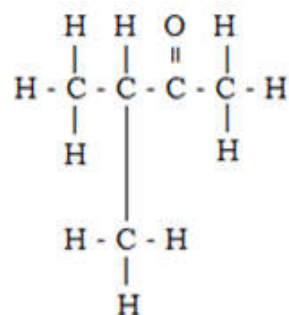
3-methylbutan-2-one

RULES FOR NAMING ISOMERS OF ALKANOIC ACIDS

Alkanoic acids are organic compounds whose functional group is the carboxyl group, –**COOH**. Alkanoic acids only exhibit chain isomerism.

Example 1

Write down and name all the structural isomers of pentanoic acid (C_4H_9COOH).

Solution:
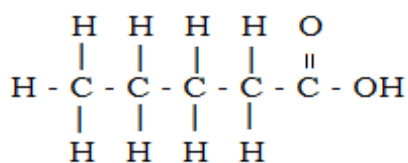
Pentanoic acid
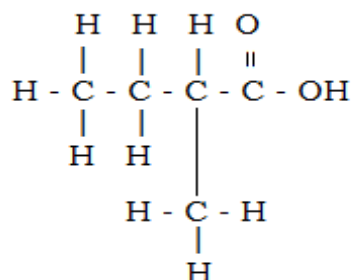
2-methylbutanoic acid
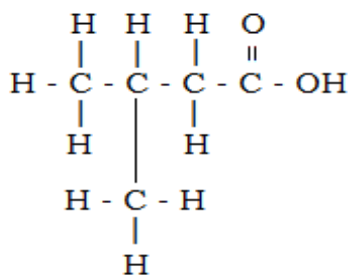
3-methylbutanoic acid
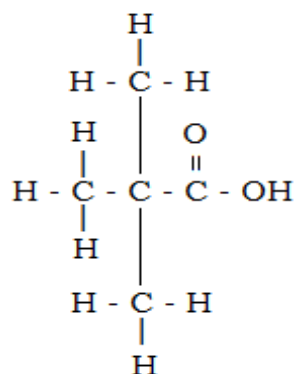
2,2-dimethlylpropanoic acid

TOPIC 6 : POLYMERISATION

Polymerisation is a chemical reaction in which small molecules join together to form long chain molecules called polymers. **Polymers** are long – chain molecules made by joining many small molecules called monomers. **Monomers** are the small molecules which join together to form polymers.

TYPES OF POLYMERS

There are two types of polymers: natural polymers and synthetic polymers.

a. **Natural polymers**

These are polymers found in all living things. Examples of natural polymers are proteins, carbohydrates, starch, wood, rubber and deoxyribonucleic acid (DNA).

b. **Synthetic polymers**

Synthetic polymers are manufactured in industries. They are normally called plastics. Examples of synthetic polymers are: Polythene, Polyvinylchloride (PVC), Terylene, Nylon, Perspex (artificial glass), Polytetraflouroethene (PTFE) or Teflon and Polyphenylethene (Polystyrene).

TYPES OF POLYMERISATION

a. **Addition polymerisation**

This is a process which involves repeated addition reactions of monomers of the same type. The monomers used in addition polymerisation are unsaturated compounds, for example ethene or derivatives of alkenes e.g. chloroethene.

Examples of addition polymerisation are

• The polymerisation of ethene

ethene   ethene   ethene   polyethene
(monomer)   (monomer)   (monomer)   (polymer)
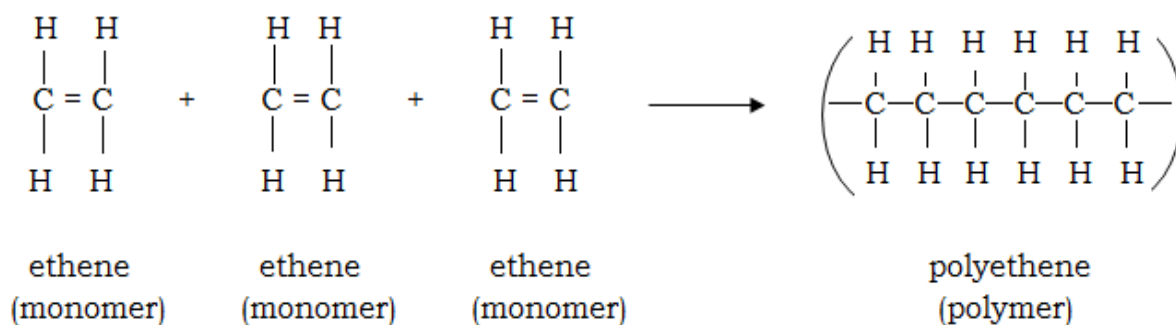
If **n** molecules of ethene combine, the process may be written as follows:

$$n \left(\begin{array}{c} H \quad H \\ | \quad\;\; | \\ C = C \\ | \quad\;\; | \\ H \quad H \end{array} \right) \longrightarrow \left(\begin{array}{c} H \quad H \\ | \quad\;\; | \\ -C - C- \\ | \quad\;\; | \\ H \quad H \end{array} \right)_n$$
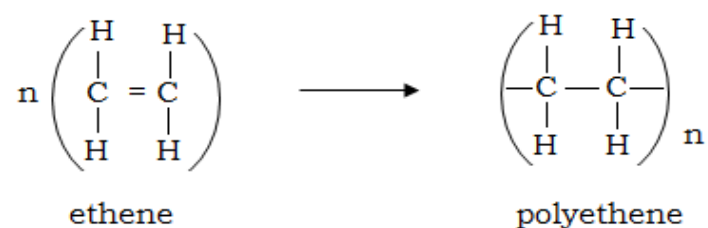
ethene   polyethene

HOW THE POLYMERISATION OF ETHENE OCCURS

During the polymerisation of ethene, ethene undergoes relatively high temperature and high pressure in the presence of a catalyst. The double bonds between the carbon atoms are broken, enabling many molecules of ethene to join to each other to form polyethene.

- The polymerisation of chloroethene

$$\begin{array}{c} H \quad H \\ | \quad\;\; | \\ C = C \\ | \quad\;\; | \\ H \quad Cl \end{array} + \begin{array}{c} H \quad H \\ | \quad\;\; | \\ C = C \\ | \quad\;\; | \\ H \quad Cl \end{array} + \begin{array}{c} H \quad H \\ | \quad\;\; | \\ C = C \\ | \quad\;\; | \\ H \quad Cl \end{array} \longrightarrow \left(\begin{array}{c} H\;H\;H\;H\;H\;H \\ | \;\; | \;\; | \;\; | \;\; | \;\; | \\ -C-C-C-C-C-C- \\ | \;\; | \;\; | \;\; | \;\; | \;\; | \\ H\;Cl\;H\;Cl\;H\;Cl \end{array} \right)$$

chloroethene   chloroethene   chloroethene   poly vinyl chloride (PVC)
(monomer)   (monomer)   (monomer)   (polymer)

$$n \left(\begin{array}{c} H \quad H \\ | \quad\;\; | \\ C = C \\ | \quad\;\; | \\ H \quad Cl \end{array} \right) \longrightarrow \left(\begin{array}{c} H \quad H \\ | \quad\;\; | \\ -C - C- \\ | \quad\;\; | \\ H \quad Cl \end{array} \right)_n$$

chloroethene   poly vinyl chloride
(vinyl chloride)   (PVC)
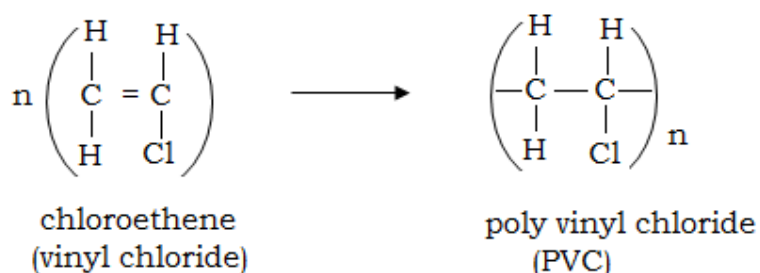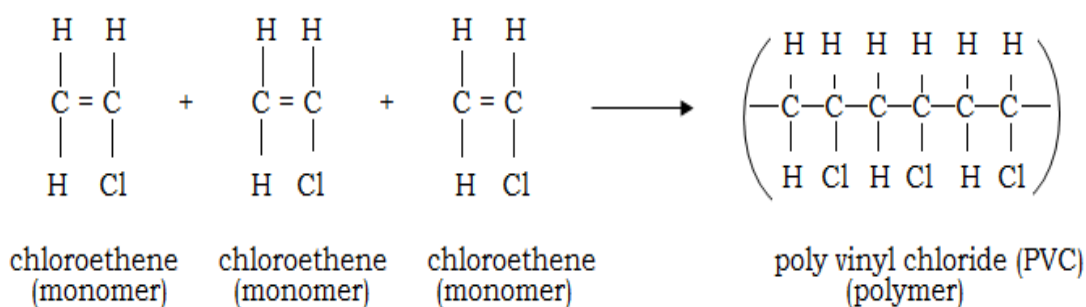
b. Condensation polymerisation

This is a reaction whereby a small molecule is eliminated between two different monomers.

During condensation polymerisation:

- two different monomers join.
- the monomers join by releasing small molecules e.g. water or hydrogen chloride

Page **56**

The best example of condensation polymerisation is the **esterification reaction.**

Carboxylic acid   +   Alkanol   ⟶   Ester   +   water
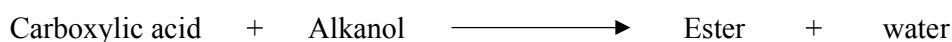
(Monomer 1)   (Monomer 2)   (Polymer)

Examples of condensation polymers are: Nylon, Terylene (polyester), etc.

PROPERTIES OF SYNTHETIC POLYMERS

Different polymers have unique characteristics. However, most of them will have the following characteristics.

- They do not corrode in air or water.
- They are excellent heat insulators.
- They are excellent electrical insulators.
- They do not decompose easily.
- They are relatively light.
- Some are transparent

USES OF SYNTHETIC POLYMERS

Polymer	Monomer	Uses
Polyethene	Ethene	Used in making: - pipes and plastic bags - bowls and bottles - packaging material - insulation for electrical wiring
Polyvinylchloride (PVC)	Chloroethene	Used in making: - crates and bottles - plastic ropes - artificial leather - water pipes - insulation for electrical wiring
Polystyrene	Styrene	Used in making: - insulation material

		• packaging (foam)
		• radio cabinets
		• pot handles
Polytetraflouroethene (PTFE) or Teflon	Tetraflouroethene	Used in making:
		• soles of iron
		• non-stick frying pans
Perspex		• used as a glass substitute
		• used in lenses

ADVANTAGES OF SYNTHETIC POLYMERS

- They are cheap to produce.

- They are resistant to corrosion from acids, alkalis, water and air.

- They are light and therefore portable.

- They are soft and easy to manufacture.

- They can be moulded into complicated shapes more easily.

- They are good heat insulators.

- They are good electrical insulators.

DISADVANTAGES OF SYNTHETIC POLYMERS

- They are non-biodegradable hence difficult to dispose.

- Some give off toxic gases such as carbon mono oxide when they burn.

- They are expensive to recycle.

- Some catch fire easily.

TYPES OF PLASTICS

There are two types of plastics: thermoplastics and thermosetting plastics.

a.   Thermoplastics (thermosoftening plastics)

These are plastics that soften when heated and harden into a new shape when cooled. This is because the intermolecular forces are weak and there are no cross-links or branches between the polymer chains.

Examples of thermoplastics are polythene, PVC, polystyrene, etc.

PROPERTIES OF THERMOPLASTICS

- They are flexible. For this reason they do not break easily.
- Their polymer chains are not cross-linked. They can be moulded into different shapes several times.
- They have low melting points. This is because they have weak intermolecular forces between the polymer chains.
- They can easily be recycled.

b. THERMOSETTING PLASTICS (THERMOSETS)

These are plastics which set hard and cannot be melted again once they are made.

This is because the polymers chains are held more strongly due to the presence of cross links between the polymer chains.

Examples of thermosetting plastics are melamine, Formica, Bakelite, among others.

PROPERTIES OF THERMOSETTING PLASTICS

- They are rigid. For this reason, they break rather than stretch under tension.
- Their polymer chains are cross-linked. Therefore, they cannot be moulded into new shapes after they are made.
- They do not melt, but decompose on heating.
- They cannot be recycled.

DIFFERENCE BETWEEN THERMOPLASTICS AND THERMOSETTING PLASTICS

Thermoplastics	Thermosets
They soften and melt when heated	They do not soften on heating
They can be re-moulded into new shape after they are made	They cannot be re-moulded into new shape after they are made.
They have no cross-links between polymer chains	They have cross-links between polymer chains
They can be recycled	They cannot be recycled

TOPIC 7 : WATER

NATURAL SOURCES OF WATER

The major natural sources of water can be classified into underground, surface and atmospheric water sources.

1. **Underground water sources**

These are sources of water found in water tables. They include:

a. Wells. These are holes dug in the ground in order to obtain water.
b. Boreholes. These are deep holes drilled in the ground using machines. They lift water using submersible pumps.
c. Springs of water. These are sources of water that flow out of the ground as small streams or pools.

2. **Surface water sources**

Examples of surface water sources include dams, lakes, streams, rivers, seas and oceans.

3. **Atmospheric water sources**

These sources include rainwater collected from rooftops during rainy season.

THE WATER CYCLE

The water cycle is also called the **hydrological cycle**. It describes how water circulates continuously between the earth's surface and the atmosphere.

- Heat from the sun causes evaporation from oceans, seas and lakes. It also causes transpiration. Both evaporation and transpiration cause water vapour to form.
- The water vapour rises, cools and condenses to form tiny droplets of water. These droplets from clouds.
- The clouds are moved by along by air currents. As they cool, the tiny droplets from larger droplets which fall as rain when they reach a certain size.
- The water that falls as rain runs into streams and rivers and then into lakes, seas and oceans.
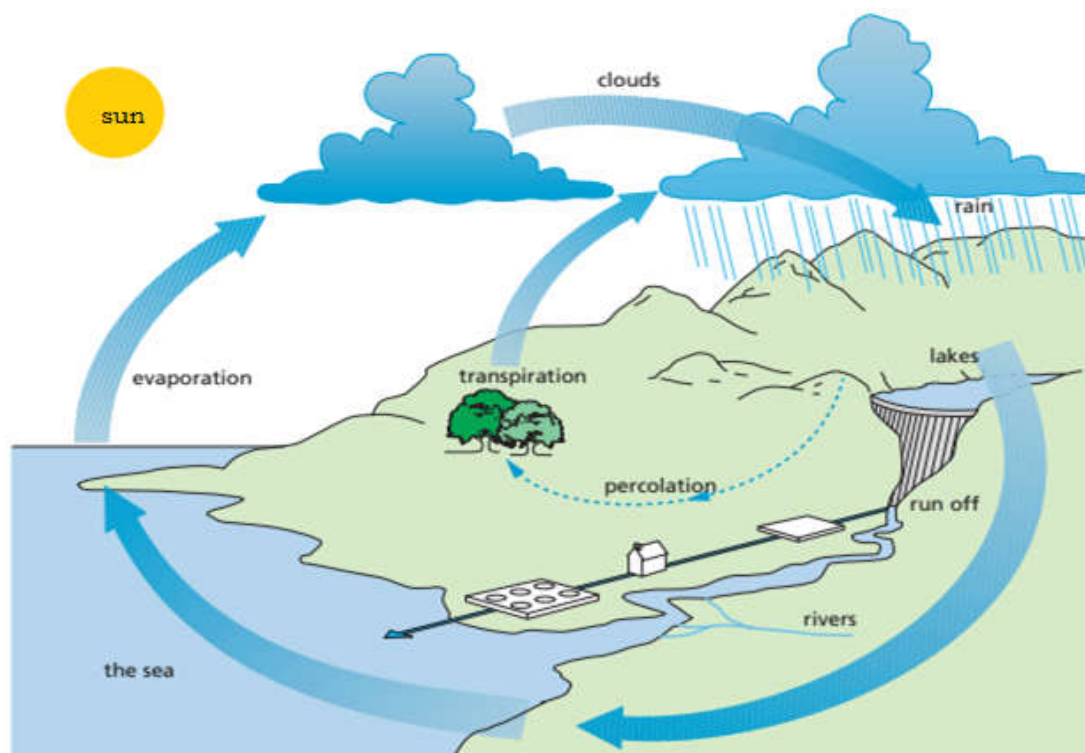
PHYSICAL PROPERTIES OF WATER

- Pure water is a colourless, tasteless and odourless liquid.
- It boils at 100ºC and freezes and 0ºC.
- It has a density of 1g/cm^3.

IMPORTANCE OF WATER

Water is used:

a. as a **coolant**. In car engines it is used to maintain low temperature while in simple distillation it is facilitates liquefaction process.

b. as a **solvent**. It is considered as universal solvent because it dissolves many things.

c. as a **chemical reactant**. It is use in many chemical reactions which occur in aqueous media.

d. in hydroelectric power generation. Moving water is used to generate electricity.

e. for **recreational purposes** such as in swimming pools.

f. as a **transport medium** for ships, ferries, boats and canoes when transporting people and goods.

HARD AND SOFT WATER

Soft water is the water that contains low mineral content and produces soapy bubbles easily. Examples of soft water are rainwater and distilled water. **Hard water** is the water that has a high mineral content and does produce soapy bubbles easily. An example of hard water is tap water.

TYPES OF WATER HARDNESS

a. Temporary hardness

Temporary hardness is caused by the presence of dissolved calcium or magnesium hydrogen carbonates. Temporary hardness can easily be removed by boiling.

b. Permanent hardness

Temporary harness is caused by the presence of calcium or magnesium sulphates. Permanent hardness is difficult to remove and cannot be removed by boiling. It can be removed by distillation or chemical treatment.

ADVANTAGES OF HARD WATER

- Calcium ions present in hard water are required by the body for bones and teeth.
- Coates lead pipes with a thin layer of lead sulphate or lead carbonate and prevents the lead from dissolving in water.
- It is good for the formation of shells in animals.
- Calcium and magnesium present in hard water are necessary for growth of both plants and animals.

DISADVANTAGES OF HARD WATER

- It wastes soap due to scum formation.
- Causes electrical appliances such as kettles and boilers to fur hence making them inefficient.
- Can cause hot water pipes to block.
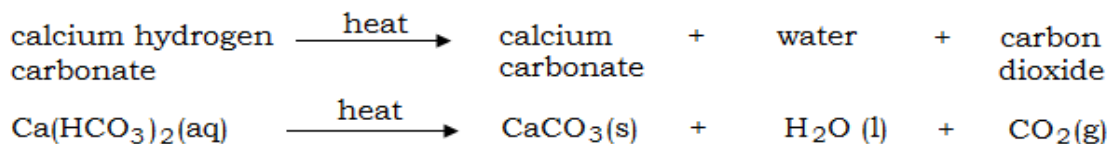- Spoils the finish of some fabrics.

METHODS OF REMOVING WATER HARDNESS

Removal of water hardness is based on whether the hardness is temporary or permanent. The method and can be physical or chemical.

1. REMOVING TEMPORARY HARDNESS

a. Boiling

When heated the hydrogen carbonates decompose to form insoluble carbonates. For example

calcium hydrogen carbonate $\xrightarrow{\text{heat}}$ calcium carbonate + water + carbon dioxide

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

b. Addition of calcium hydroxide or sodium hydroxide

Addition of calcium or sodium hydroxide to water causes precipitation of calcium and magnesium ions in form of carbonates. For example

calcium hydrogen carbonate + calcium hydroxide $\longrightarrow$ calcium carbonate + water

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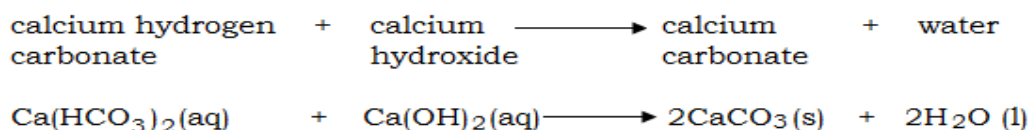
2. REMOVING PERMANENT HARDNESS

a. Distillation

The water is distilled away from the dissolved substances. This method is expensive to be used on a large scale.

b. Addition of sodium carbonate (soda)

Calcium and magnesium ions are removed as precipitates. For example

magnesium sulphate + sodium carbonate $\longrightarrow$ calcium carbonate + sodium sulphate

$$MgSO_4(aq) + Na_2CO_3(aq) \longrightarrow MgCO_3(aq) + Na_2SO_4(aq)$$
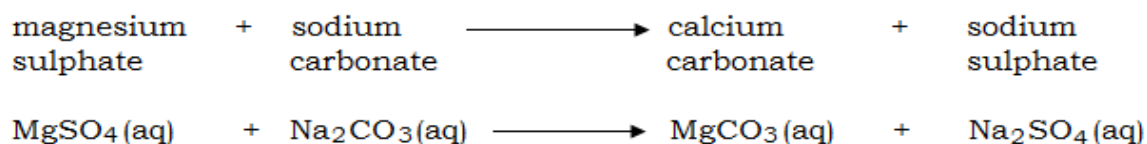
c. Ion exchange

The hard water is passed through a container filled with suitable **resin** containing sodium ions. The calcium or magnesium ions causing water hardness are exchanged for the sodium ions in the resin. Sodium ions do not cause water harness.
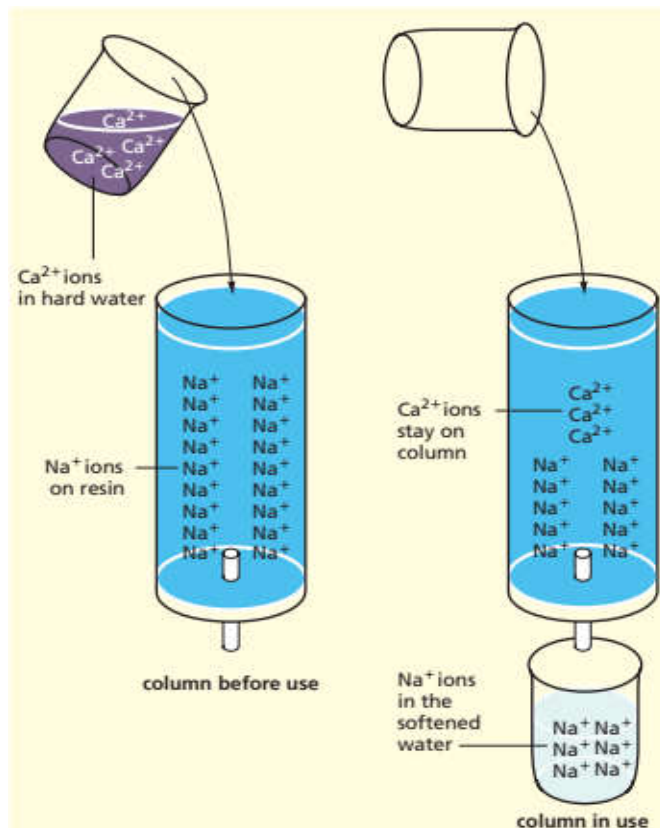
Sodium ions become part of the softened water and the calcium or magnesium ions become part of the resin.

WATER POLLUTION

It is defined as the addition of harmful substances into water as result of human activities. The harmful substances include toxic metals, plastics, pesticides, and fertilizers.

SOURCES OF WATER POLLUTION

a. Domestic sources.

This involves the release of human faeces and sewage into water bodies such as lakes and rivers.

b. Industrial wastes

These include effluents (liquid wastes) originate from breweries, tanneries, textiles and paper industries. The effluents contain toxic chemicals.

c. Agricultural wastes

These include excessive pesticides, herbicides, weedicides, and fertilizers. These may enter into rivers and lakes through run off after rains.

EFFECTS OF WATER POLLUTION

- It can cause diseases such as cholera, typhoid and hepatitis.
- It leads to destruction of ecosystems.
- It can severely affect aquatic life by decreasing amount of oxygen in water bodies.
- It can cause disruptions in food chains. This happens when toxins and pollutants in the water are consumed by aquatic animals which are then consumed by humans.

WAYS OF CONTROLLING WATER POLLUTION

- Avoiding excessive use of chemicals.
- Treatment of sewage before releasing it into water bodies. It ensures that the pollutants and disease-causing micro-organisms are destroyed.
- Encouraging farmers to use farmyard manure instead of chemical fertilizers.
- Constructing pit latrines away from water sources such as wells to avoid contamination of the water source. The slope of the land should also be considered in putting up a latrine.
- Proper legislation should be put by government to reduce improper damping of industrial and domestic wastes.
- Petroleum companies should put measures in place to avoid unnecessary accidents resulting from spillage of oils into water bodies.

TOPIC 8: GREENHOUSE GASES AND THE OZONE LAYER

NATURAL DISASTERS

Natural disasters are extreme, sudden events caused by environmental factors that injure people and damage property. Examples of natural disasters are drought, windstorm, earthquake, wild fires, floods, hurricanes, thunderstorms and volcanic eruptions.

IMPACT OF HUMAN ACTIVITIES ON NATURAL DISASTERS

Human activities can have an impact on natural disasters through floods following these activities;

- **Deforestation**. Hard surfaces such as sewers, parking lots, and concrete surfaces causes water to move faster as compared to natural channels. Cutting down natural vegetation and trees leaves land bare resulting in more erosion when water flows.
- **Urban development**. Urban development has an impact on natural disasters. It destructs the natural system of waterways.
- **Natural wetlands destruction**. The destruction of natural wetlands causes devastating floods during peak floods along river basins. More natural wetlands lower water level during heavy rains.
- **Agriculture**. Large scale practice of agriculture can result in the depletion of top soils causing weakening of vegetation and inability to protect against flooding.

EFFECTS OF NATURAL AND HUMAN ACTIVITIES ON THE ATMOSPHERE

- Destruction of buildings
- Death of animals and human beings
- Change in landscape
- Damage of vegetation
- Global warming due to production of harmful gases

GREENHOUSE GASES

Greenhouse gases are components of air whose accumulation in the atmosphere cause rise in global temperatures. Examples of greenhouse gases are carbon dioxide, sulphur dioxide, nitrogen oxides, and methane.

SOURCES OF GREENHOUSE GASES

Greenhouse gas	Sources
Carbon dioxide (CO_2)	- Combustion - Respiration - Fermentation - Decomposition
Methane (CH_4)	- It is the primary component of natural gas. - Wetlands - Oceans - Activities of termites
Nitrogen oxides (NO, NO_2)	- Natural processes e.g. lightning and biological processes that occur in plants and animals. - Combustion of fossil fuels in car engines and in industries.
Sulphur dioxide (SO_2)	Sources include fossil fuels combustion at power plants and extraction of metal ores

GLOBAL WARMING

Global warming is the continued rise in the average temperature of the earth due to the accumulative effect of the greenhouse gases in the atmosphere.

HOW DO GREENHOUSES CAUSE GLOBAL WARMING?

- The sun sends out energy as light and ultraviolet (UV) rays.
- These warm the earth, which reflect some of the energy away again, as heat.
- Some of this heat escapes from the atmosphere. However, some of reflected heat energy is absorbed by greenhouse gases in the atmosphere. So, the air and the earth are warmed.

CONTRIBUTION OF CARBON DIOXIDE AND METHANE TO GLOBAL WARMING

Two main greenhouse gases believed to be responsible for global warming are carbon dioxide and methane. The levels of both gases are rising.

- The level of carbon dioxide is rising because we burn more fossil fuels each year. The carbon dioxide from this goes into the atmosphere.
- The level of methane is increasing because there is an increase in animal farming and rice farming around the world, and more landfill sites.

WAYS OF MITIGATING EFFECTS OF GLOBAL WARMING

- Use of renewable sources of energy such as solar energy, wind power, and hydroelectric power.
- Automobiles can be fitted with catalytic converters to clean up exhaust gases before they are released into the atmosphere.
- Afforestation and deforestation in order to reduce the carbon dioxide levels in the atmosphere.

AIR QUALITY STANDARDS

Internationally there are standards of air quality which limit emissions of gases. The air quality standards being used currently are based on the World Health Organisation (WHO.) In line with the WHO guidelines, the Malawi government through the Malawi Bureau of standards (MBS) published guidelines on maximum allowable emissions of both gases and particulate pollutant.

The table below shows ambient air quality standard limits in Malawi.

Pollutant	Maximum concentration in ambient air	Averaging period
Suspended particulate matter	$25\mu g/m^3$	1 day
Carbon monoxide	9 ppm	8 hours
	35 ppm	1 hour
Sulphur dioxide	0.20 ppm	1 hour
	0.08 ppm	1 day
	0.02 ppm	1 year

Nitrogen dioxide	0.3 ppm	1 year
Ozone	0.12 ppm	1 hour
Lead	0.50µg/m^3	1 year
Photochemical oxidants	0.10 ppm	1 hour
	0.08	4 hours

The table below shows air quality standard for motor vehicles in Malawi.

Pollutant	Maximum allowable concentration	Duration	Typical high concentration for specific pollutant, ppm
Hydrocarbons	160µg/m^3	3 hours	6
Carbon monoxide	10µg/m^3	8 hours	40
Nitrogen dioxide	100µg/m^3	1 year	0.1
Oxidants	160µg/m^3	1 hour	0.6

ppm means parts per million

µg/m^3 means concentration in micrograms per unit volume

THE OZONE LAYER

The ozone refers to a region of the earth's stratosphere that absorbs most the sun's ultraviolet radiation. It is mainly found in the lower parts of the stratosphere from approximately 20km to 30km above the earth.

IMPORTANCE OF THE OZONE LAYER

The ozone layer acts a s natural ultraviolet (UV) rays shield for the earth. It absorbs 93% to 99% of the sun's ultraviolet light. The UV light is very dangerous to life on earth.

Exposure to UV rays may cause:

- Skin cancer

- Damage to crops
- Harm to aquatic life

DEPLETION OF THE OZONE LAYER

The ozone layer is destroyed by chlorofluorocarbons (CFCs). Things that release CFCs include coolants, foams and aerosols. When CFCs get to atmosphere, the chlorine atom is released from the CFC molecule. The released chlorine atom split the ozone to oxygen atoms and oxygen molecules. The free oxygen atom then combines with ozone molecule to form another oxygen molecule.

For example

- $O_3 (g) \longrightarrow O_2 (g) + [O]$
- $O_3 (g) + [O] \longrightarrow 2O_2 (g)$

The action of the CFCs results in the depletion of the ozone layer and reduces its UV natural shielding capacity.

PROBLEMS ASSOCIATED WITH THE DEPLETION OF THE OZONE LAYER

Exposure to UV rays has several health related and environmental effects. Some of them are:

- Increased risk for developing several types of skin cancer.
- Cataract problems
- Damage to immune system.
- Aging of the skin. This results in a person looking older than what they actually are.
- Respiratory problems such as difficulty in breathing, chest pain, and throat irritation.
- Leads to the decline of some amphibians by affecting every stage of their life cycle.
- Causes disruptions in the food chain and the oceanic and carbon cycle.
- It alters physiological and developmental processes of the plants.

TOPIC 9 : WASTE MANAGEMENT

Waste refers to unwanted materials produced by human activities.

CLASSIFICATION OF WASTES

Wastes can be classified based on different criteria; physical state and degradability.

CLASSIFICATION BASED ON PHYSICAL STATE

Based on physical state, wastes can be classified as solid wastes, liquids wastes, gaseous wastes and sludge wastes

CLASSIFICATION BASED ON DEGRADABILITY

Based on degradability, wastes can be classified as:

- **Biodegradable wastes**. These are wastes that can decay relatively quickly as a result of the action of bacteria. They include animal and plant waste.
- **Non–biodegradable wastes**. These are wastes that do not decompose naturally. Examples include plastics and glass.

SOURCES OF WASTES

a. **Domestic sources:** These wastes include food leftovers, plastics, metals, untreated sewage, expired or unused medicines and their containers.
b. **Industrial sources:** These include used oil, toxic gases, metals and plastics etc.
c. **Agricultural sources:** These include plant remains excess agricultural chemicals such as pesticides, herbicides, and weedicides.
d. **Medical sources:** These are wastes from medical centres and hospitals. Examples are plastics, glassware, expired drugs, used syringes, needles, cotton and detergents.
e. **Laboratory sources:** These include expired chemicals, contaminated chemicals, broken glassware and detergents.

WAYS OF TREATING AND DISPOSING WASTE

a. Open burning

Wastes such as papers, plastics are usually burnt to ashes in open land. This method is discouraged because it causes air pollution due to emission of harmful greenhouse gases.

b. Incineration

This approach involves combustion of waste material. The refuse is burn to ashes in controlled chambers called incinerators. Incineration quickly reduces waste volume, reduces transportation costs and decreases emissions of harmful green house gases

c. Dump site

This is an open disposal site where wastes are deposited and left uncovered and uncontrolled. It is not good for the environment as it is likely to cause air and water pollution.

d. Sanitary landfills

These are land disposal sites used to bury solid waste. The solid waste is spread in thin layers compacted and covered with a fresh layer of soil such as clay. It is also characterized by the absence of surface water bodies or low water table, preventing the risk of water pollution.

e. Compositing

This is the controlled aerobic decomposition of organic waste materials by the action of small bacteria. The waste is dumped into a dug pit so that it decomposes. After it has decomposed, it can be used by farmers as manure.

f. Recycling or re-use

Recycling involves reprocessing of materials that would be considered waste. For, most plastics and paper can be recycled instead of burning them.

Advantages of recycling materials

- It reduces pollution.
- It reduces the cost of producing new products.

Re-using on the other hand involves using a material considered waste for another purpose that was not initially intended for.

For example, instead of dumping used bottles and cans, we can use them for storing water in our homes.

WAYS OF MINIMIIZNG WASTES

Waste can be minimized by three main ways. These are:

- Recycling
- Reusing
- Reduction at source

RECYCLING OF METALS AND PLASTICS

Recycling of most materials including metals and plastics proceed as follows:

a. Collection – getting the recyclable plastics from different waste disposal sites.
b. Sorting – the plastics are sorted out in terms of type, colour, even how it was made.
c. Cleaning – the goal is to remove impurities and everything that is not recyclable.
d. Resizing – it consists of shredding the material into small particles, making it easier to process, reshape.
e. Identification – the small particles are tested to determine their quality and class.
f. Compounding – it involves smashing and melting the material pellets. The pellets can then be used in the production of other products.

IMPORTANCE OF RECYCLING METALS AND PLASTICS

- It preserves natural resources. It prevents depletion of the metals and materials used to manufacture plastics.
- It reduces emissions of harmful greenhouse gases.
- It promotes economic development by saving money that could be spent on mining and processing of new metals and plastics. The money can be channeled into other development programs that can help raise the living standards of the people and the economy as a whole.
- It saves money since recycled materials are generally cheap. This means the actual cost of buying the material is reduced hence saving money.