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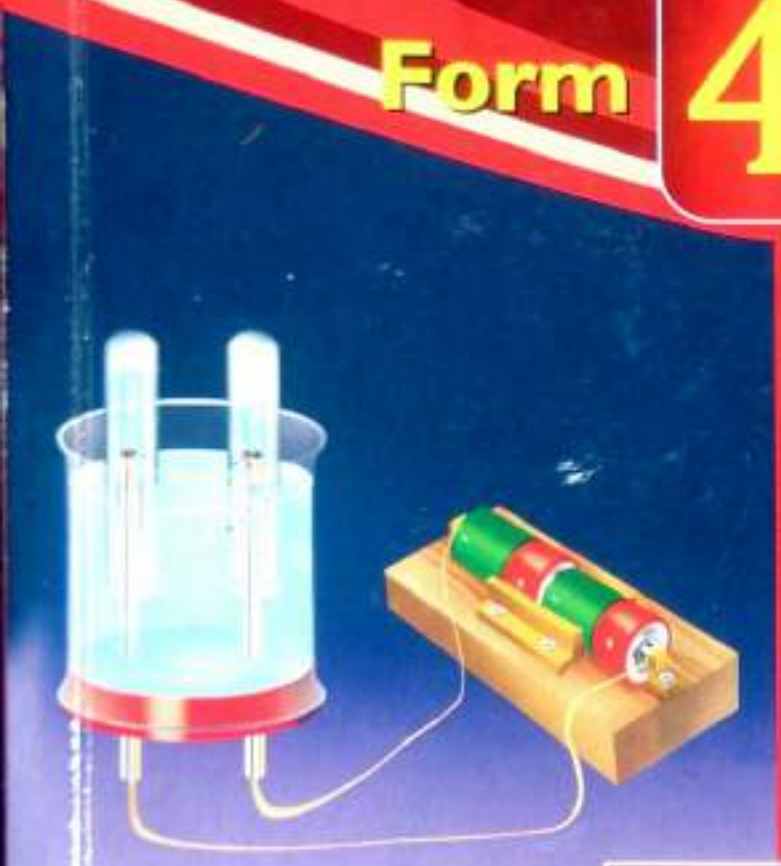
Senior Secondary

Chemistry

Student's Book

Form 4



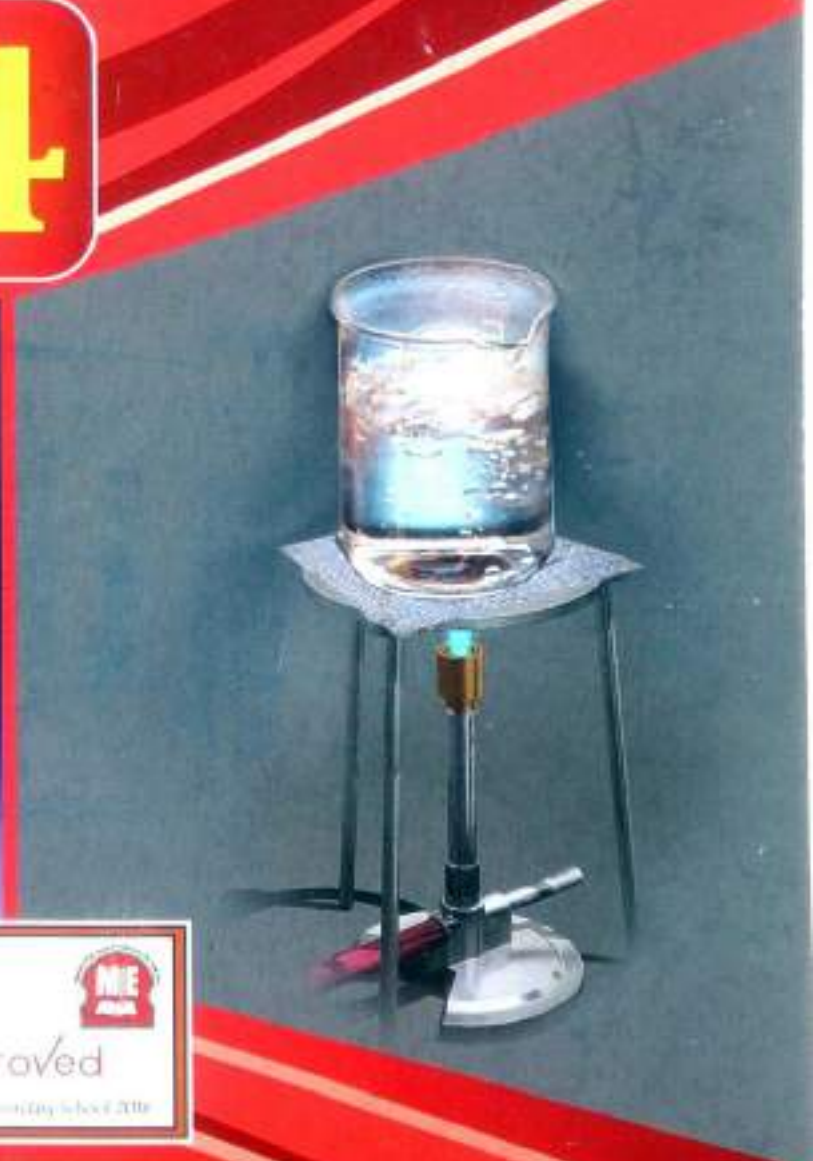


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Table of Contents

Topic 1: Rates of reactions 1
1.1: Collision theory and activation energy 1
1.2: Factors affecting rates of reactions 3
1.3: Methods of measuring rates of reactions 15
1.4: Reversible and irreversible reactions 15
1.5: Chemical equilibrium 19
1.6: Effects of different factors on the position of equilibrium 22
1.7: Uses of reversible reaction in industrial processes 28
Revision Exercise 1 32

Topic 2: Acids and Bases 34
2.1: Meaning of acid and a base 34
2.2: Conjugate acids and bases according to Bronsted-Lowry theory 36
2.3: Strengths of acids and bases 37
2.4: Ways of regulating pH in different environments 43
2.5: Classification of oxides 44
2.6: Ways of preparing salts 46
2.7: Application of precipitation reactions 64
Revision Exercise 2 65

Topic 3: Reduction and Oxidation reactions 66
3.1: Reduction and oxidation 66
3.2: Determination of oxidation numbers 68
3.3: Oxidation numbers and redox reactions 71
3.4: Writing and balancing redox equations 73
3.5: Order of reactivity of metals 75
3.6: Spontaneity of a reaction 80
3.7. Electrochemical cell 83
3.8: Corrosion and rusting 95
Revision Exercise 3 100

Topic 4: Electrolysis 102
4.1: Electrolysis process 102
4.2: Electrolysis of acqueous and molten compounds 110
4.3: Uses of electrolysis 121
4.4: Applications of electroplating 125
Revision Exercise 4 127

Topic 5: Isomerism 130
5.1: Isomerism in alkanes 131
5.2: Isomerism in alkenes 135

5.3: Isomerism in alkanols 138
5.4 (a) Isomerism in alkanals 141
5.4 (b) Isomerism in alkanones 142
5.5: Isomerism in alkanoic acids 144
Revision Exercise 5 ..... **146**

Topic 6: Polymerisation 148
6.1: Types of polymers 148
6.2: Types of polymerisaton 151
6.3: Properties of synthetic polymers 157
6.4: Thermoplastics and thermosetting polymers 158
Revision Exercise 6 ..... **159**

Topic 7: Water 160
7.1: Natural sources of water 160
7.2: Water cycle 161
7.3: Physical properties of pure water 163
7.4: Importance of water 164
7.5: Water hardness 164
7.6: Methods of removing water hardness 169
7.7: Water pollution 172
7.8: Water quality standards 174
7.9: Impact of human activities on the atmosphere 176
Revision Exercise 7 ..... **178**

Topic 8: Greenhouse gases and ozone layer 179
8.1: Natural disasters and their effects on the atmosphere 179
8.2: Greenhouse gases 180
8.3: Air quality standards 185
8.4: Global warming 186
8.5: Ozone layer 187
Revision Exercise 8 ..... **189**

Topic 9: Waste management 190
9.1: Classification of wastes 190
9.2: Source of wastes 191
9.3: Treatment and disposal of solid wastes 192
9.4: Treatment and disposal of liquid wastes 196
Revision Exercise 9 ..... **197**
Model Test Paper 198
Glossary 201
Appendices 204

Topic 1 Rates of reactions

Success criteria

After studying this topic, you should be able to:

(a) Describe factors that affect rates of reactions.
(b) Describe methods of measuring rates of reactions.
(c) Explain reversible and irreversible reactions.
(d) Describe chemical equilibrium in reversible reactions.

Introduction

A chemical reaction involves breaking of bonds in the reactants and formation of new ones in the products. As such, energy must be used for a chemical reaction to occur.

Reaction rate is a measure of how fast or slow a reaction takes place. To be able to know how fast a reaction is taking place, we measure the rate at which a reactant is consumed or a product is formed over a given period of time. i.e.

$$\text{Reaction rate} = \frac{\text{Change in amount of substance}}{\text{Time taken}}$$

Therefore we can define reaction rate as the **rate of change of an amount or concentration of a particular reactant or product per unit time**.

Different reactions occur at different rates. Some reactions occur very fast, others occur slowly while still some are very slow. For example, burning of magnesium takes place very fast. Burning of charcoal is slow while rusting of iron is very slow that you hardly see any change when you look at it!

To understand the definition of reaction rate better, let us consider the theories that explain why a reaction occurs either fast or slow.

1.1 Collision theory and activation energy

All substances are made up of particles according to **Kinetic Theory**. Chemical reactions are possible due to collisions of these particles. For a chemical reaction to take place, the particles must collide,thus what happens during chemical reactions is summarized in what is known as **collision theory**.

According to collision theory:

- Reactions take place due to collisions of reacting particles. These particles include atoms, ions and molecules.
- Collisions which produce a chemical reaction are **fruitful collisions**. Not all collisions between particles may be fruitful. The rate at which particles collide is called **collision frequency**. The greater the collision frequency, the faster the reaction rate.

- Colliding particles need minimum amount of energy before they can bring about a chemical reaction. This energy is called **activation energy**.

Reaction rate is usually the number of atoms, ions or molecules that react per unit time to form products. To increase the rate of a chemical reaction, you must increase the number of fruitful collisions between particles.

Reactions involving **solids** take place at the surface of the solid. This is because it is only here that collisions with particles of other reactants can take place. In a solid lump, most particles are locked up inside. In order to react, these particles must have fruitful collisions with the reacting reagent. If the reacting reagent is a liquid or gas, their particles collide with the particles of the solid as shown in Fig. 1.1 below.

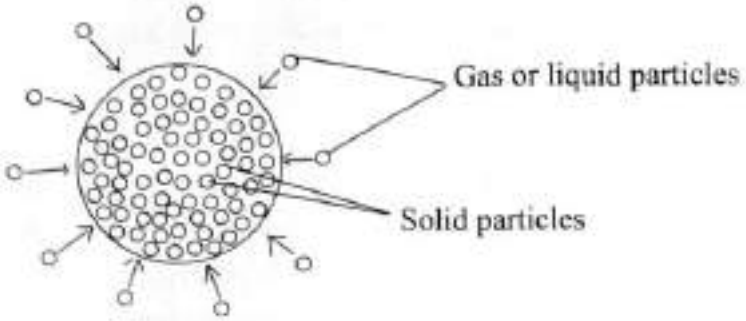

Fig. 1.1: Collision between particles of solid and that of liquids or gases

In gases or liquids, the particles are constantly in random motion. During the reaction, the particles collide with one another. If the collisions are fruitful, a reaction takes place.

Reactions do not take place between two solid particles. This is because their particles only vibrate within fixed positions. i.e. they do not move freely. They therefore cannot meet to collide effectively or fruitfully. However, when mixed with water or other solvents, solid particles have an opportunity to move freely; a chance they do not have in solid form. Under such circumstances, the particles meet, collide and reaction occurs.

Self assessment exercise 1.1

1. Define rate of reaction.
2. Explain why particles sometimes collide and products are not formed.
3. Explain the term activation energy.
4. Explain why:
   (a) One has to strike a matchstick for it to ignite.
   (b) When the matchstick starts burning, the process is self-sustaining.
5. There will be no reaction between solid sodium chloride and solid lead nitrate when they are mixed together. Explain why.

1.2 Factors affecting rates of reactions

Research Activity

Find out from textbooks in the library or the Internet factors that affect reaction rate. Write a report and share with your class members.

There are several factors which affect the rate at which a chemical reaction occurs. The factors include:

- concentration of reactants (for aqueous reactants)
- pressure (for gaseous reactants)
- temperature at which the reaction is taking place.
- surface area of reactants (if the reactants are solids).
- light (type and intensity).
- presence of catalysts.

Any of these factors or a combination of them may alter the rate of a chemical reaction. The effect on rate of reaction by any one factor can be determined in the laboratory if the other factors are kept constant. Let us now look at how each of the factors affects the rate of chemical reactions.

(a) Concentration

Experiment 1.1

Aim: *To investigate the effect of concentration on reaction rate*

Apparatus and reagents

- Three 250 cm^3 conical flasks or 100 cm^3 beaker
- Filter papers or white tile
- Stirring rod
- Stopwatch or stop clock
- Sodium thiosulphate ($Na_2S_2O_3$)
- 2M hydrochloric acid
- Black ink

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Procedure

1. Place three conical flasks on a bench.
2. Put 200 cm^3 of distilled water into each of the 250 cm^3 conical flasks. Label the flasks A, B and C.
3. Put 0.5 g, 1.0 g and 1.5 g of sodium thiosulphate in flasks A, B and C in that order. Stir using the stirring rod to dissolve.
4. Make a black ink cross on a filter paper.

5. Place flask A the filter paper or white tile with a black ink cross (X) on it. Add 5 cm^3 of 2M hydrochloric acid to the flask. Start the stop watch or stop clock and swirl the contents of the flask. Note and record the time it takes for the cross to be obscurred (when you cannot see it anymore). See Fig.1.2.

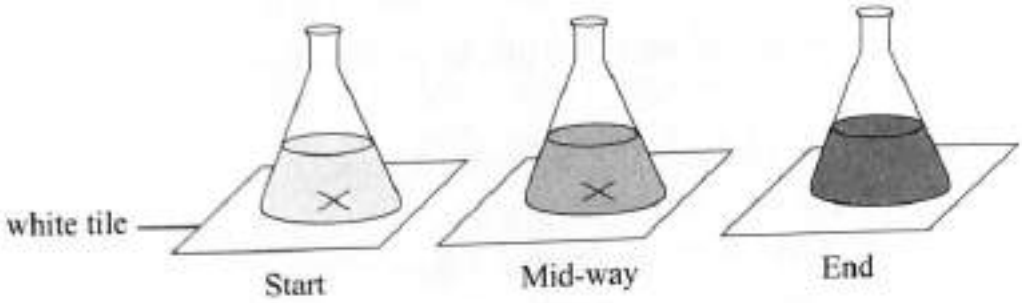

Fig. 1.2: Set-up for determining rate of reaction

6. Repeat steps 3-5 using 1.0 g and 1.5 g of sodium thiosulphate. Label the flask A and B respectively. Record the time taken in each case for the cross to disappear from your sight, use a table similar to Table 1.1.

Table 1.1: Effect of concentrations on rate of reaction

Amount of sodium thiosulphate used (g)	Time taken for the cross to dissappear (in seconds)
0.5	
1.0	
1.5	

7. Plot a graph of amount of sodium thiosulphate used against time (in seconds) taken for the cross to disappear.

Note: To ensure that consistent results are obtained, use:

- same size of the cross in each experiment.
- identical flasks in all experiments.
- cross should be viewed from the same position.
- a clean flasks.
- the same person to observe the cross disappearing.
- same amount of swirling or stirring should be done for each amount of sodium thiosulphate.

Study Questions

1. What is the name of the yellow precipitate?
2. What does using more sodium thiosulphate mean?
3. Write the chemical equation for the reaction that is taking place.
4. What do you conclude from the graph?

When sodium thiosulphate and dilute hydrochloric acid are mixed, a yellow precipitate (i.e sulphur solid) is formed.

$$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S\ (s)$$

(yellow precipitate)

The more sodium thiosulphate is dissolved in a given amount of water, the higher the concentration of this solute in the solution.

As the concentration of sodium thiosulphate increases the time taken for the cross (X) to disappear reduces. This indicates that increase in concentration increases the rate of the reaction. If we plot concentration of sodium thiosulphate against time taken; a typical graph as shown in Fig. 1.3 is obtained.

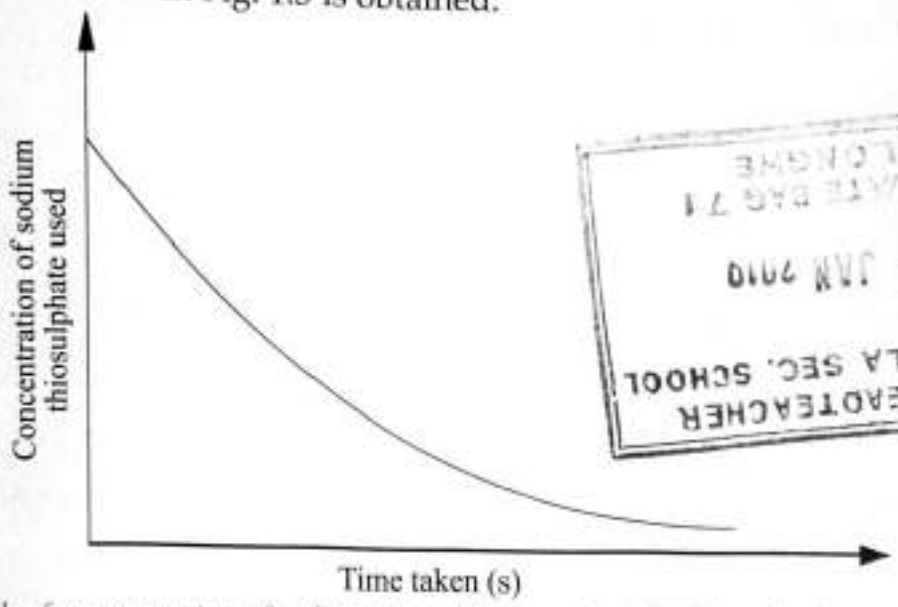

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Fig. 1.3: Graph of concentration of sodium thiosulphate used against time for the cross to disappear

From the graph, the higher the concentration of sodium thiosulphate solution, the less the time taken for the black ink cross to disappear. That is, the higher the concentration, the higher the rate of reaction. If a graph of mass of sodium thiosulphate used against the reciprocal of time i.e. is drawn, a straight line graph as shown in Fig. 1.4 below is obtained.

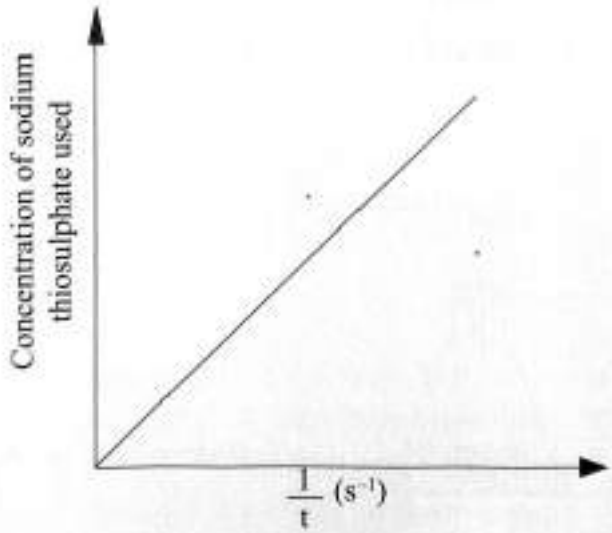

Fig. 1.3: Graph of concentration of sodium thiosulpate used against $\frac{1}{t}$ (s^{-1})

From the collision theory, increasing the concentration of reactants increases the frequency of collision between reacting particles. The greater the number of collisions, the higher the rate of reaction. An increase in mass will imply an increase in concentration of particles since in this case more solid dissolves in the solution. Hence rate of reaction is increased. The frequency of collisions between reacting particles therefore will determine the rate of reaction.

Self assessment exercise 1.2

1. How does concentration of reactants affect the rate of a reaction?
2. How do concentrations in Fig. 1.3 page 5 change with time?

(b) Temperature

Experiment 1.2

Aim: *To investigate the effect of temperature on the rate of a reaction*

Apparatus and reagents

- 250 cm^3 conical flask
- Magnesium ribbon
- A rubber bung with one hole
- Source of heat (Bunsen burner)
- Thermometer (0 – 100°C)
- Stopwatch or stop clock
- Emery paper or sand paper
- Delivery tubes
- Graduated syringe
- 0.5 M hydrochloric acid

Procedure

1. Clean about 8 cm of magnesium ribbon with an emery paper or sandpaper and coil it.
2. Put 40 cm^3 – 50 cm^3 of 0.5 M hydrochloric acid into the conical flask.
3. Arrange the apparatus as in Fig. 1.5 below.

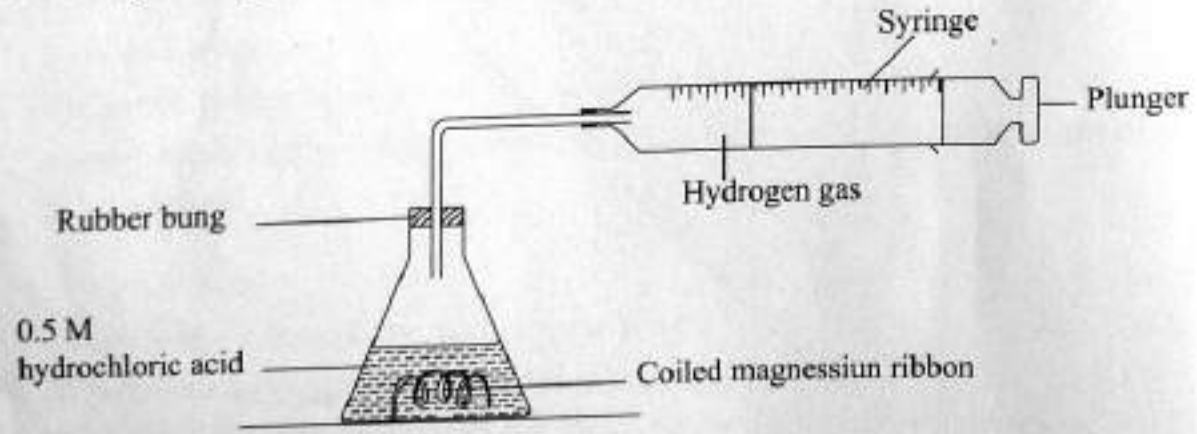

Fig. 1.5: Apparatus to investigate effect of temperature on reaction rate

4. Before the reaction begins, place the rubber bung tightly in the neck of the flask and measure the volume of air displaced on the graduated syringe. Remove the rubber bung, place the plunger at zero on the scale.
5. Measure and record the temperature of the acid.
6. Place the magnesium coil in the flask and insert the rubber bung as quickly as possible.
   - Start the stop watch or stop clock at the same time.
7. Record the volume of hydrogen gas liberated after every 30 seconds until the reaction is over.

   Note: *The reaction is over when there is no more effervescence.*

8. Draw and fill Table 1.2 in your notebook.

Table 1.2: Effect of temperature on reaction rate

Time taken (in seconds)	30	60	90	120	150	180	210
Volume of hydrogen gas produced (cm^3)							

8. Repeat the experiment using same volume and concentration of acid at 30°C, 40°C, 50°C and 60°C. For every experiment, copy Table 1.2 in your notebook and fill it.
9. Plot a graph of volume of hydrogen produced against time in each experiment on the same axes. Compare your curves with Fig. 1.6.

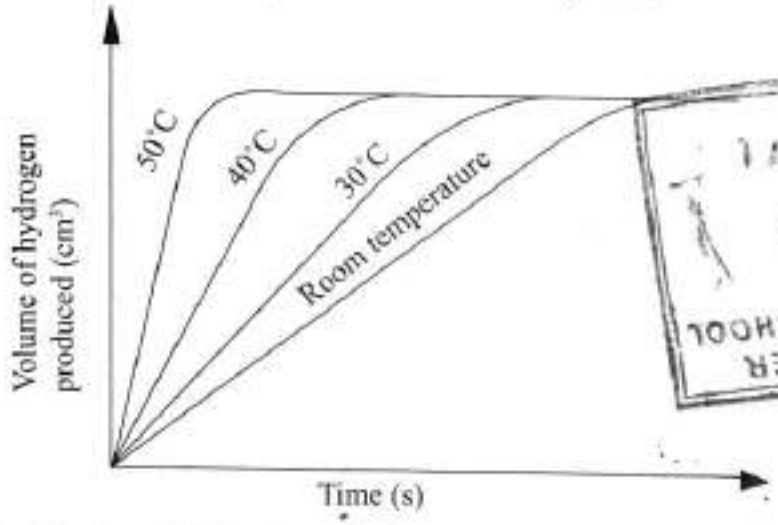

Fig. 1.6: A graph of volume of hydrogen gas produced against time

Study Questions

1. What do you conclude about the relationship between volume of hydrogen gas collected per second and the temperature of the acid?
2. How does the volume of hydrogen change with time?
3. What is your conclusion on the effect of temperature on rate of reaction?

When the temperature of the acid is increased, a higher volume of hydrogen is produced per second. The volume of hydrogen gas produced per unit time is a measure of the rate of reaction in this experiment. Therefore, the rate of reaction increases with increase in temperature. This is because the velocity of all the reacting particles increase as the temperature increases. Increase in velocity increases the frequency of collisions of reacting particles. Increase of kinetic energy provides the particles with the necessary **activation energy** required for the reaction to occur. The reaction therefore proceeds at a faster rate at higher temperatures. Therefore, the total amount of time taken for the reaction to come to an end reduces.

Self assessment exercise 1.3

1. Explain how increase in temperature increases the rate of a reaction.
2. Sketch a graph of the reaction of a metal with an acid at 30°C and 60°C.

(c) Surface area

Surface area of reactants affects rates of reactions of solids. Let us find out how.

Experiment 1.3

Aim: *To investigate the effect of surface area of solid reactants on the rate of a reaction.*

Apparatus and reagents

- Electronic balance or any other suitable balance
- Conical flask
- Cotton wool
- Watch glass
- Stop watch or stop clock
- Marble chips or Calcium Carbonate ($CaCO_3$)
  - (i) 2 g fine powder
  - (ii) 2 g small chips
  - (iii) 2 g large chips
- 2 M hydrochloric acid

Procedure

1. Measure 40 cm^3 of 2M hydrochloric acid and transfer into a conical flask.
2. Weigh 2 grams of fine marble powder on a watch glass.
3. Put the marble powder into the flask containing hydrochloric acid and immediately place the flask and its contents on the electronic balance as shown in Fig. 1.7 below. Let the initial balance reading be X.

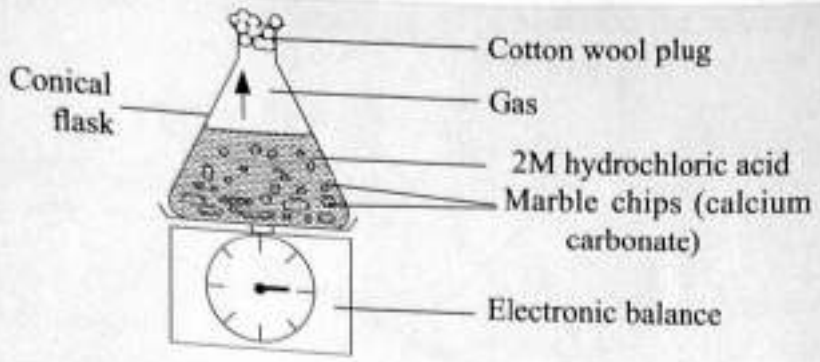

Fig. 1.7: Determining loss of mass of calcium carbonate

- What is the use of cotton wool in the set-up above?

4. Leave the set up for sometime to allow the reaction to occur.
5. Draw three tables a, b and c similar to Table 1.3 below in your notebook and record the electronic balance readings after every minute until there is no change in mass.

Table 1.3: Effect of surface area on reaction rate

Time (minutes)	1	2	3	4	5	6
Mass (y) (in g)						
Loss of mass of marble chips (x – y) g						

Calculate the loss in mass i.e. (x – y) g and complete your table.

6. Repeat the experiment with 2 g of small marble chips. Complete Table 1.3 and finally use 2 g of large marble chips and complete the table.
7. Plot a graph of loss in mass against time in each experiment on the same axes.

Study Questions

1. Write the chemical equation for the reaction taking place in the flask.
2. Why does the reaction stop?
3. What can you conclude from the graph?

Calcium carbonate reacts with dilute hydrochloric acid to form calcium chloride, water and carbon dioxide.

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

The carbon dioxide escapes through the cotton wool to the atmosphere. The cotton wool stops the acid spray from escaping during the chemical reaction. The hydrochloric acid used in this experiment is in excess. Therefore, the reaction stops only when all the calcium carbonate has reacted.

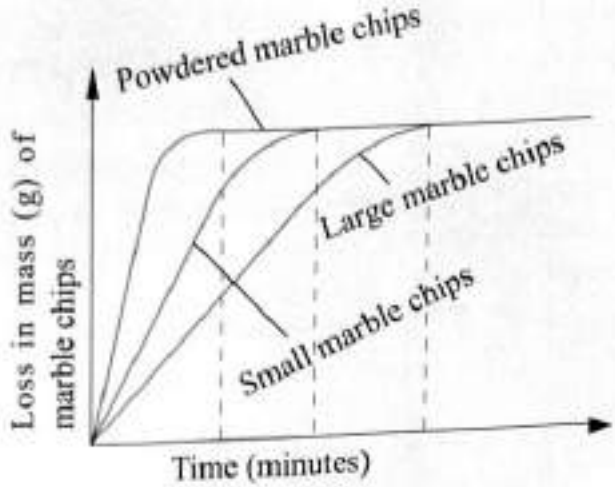

Fig. 1.8: Curves of loss of mass of marble chips against time

The loss in mass of the marble chips is as a result of production of carbon dioxide gas. When the three graphs are compared, the loss in mass is fastest for powdered marble, slower for small marble chips and slowest for large marble chips. This means that the rate of reaction is fastest for powdered marble powder, followed by small marble chips and slowest for large marble chips. The more finely divided the solid, the greater the surface area as shown in Fig. 1.9 below.

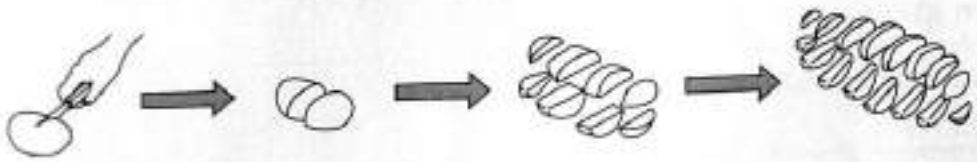

Fig. 1.9: Cutting to smaller pieces increases the total surface area of a substance

Large surface area increases the chances of successful collissions between particles and hence greater frequency of these collisions. Increased collisions leads to an increase in the rate of reaction. The rate of chemical reaction can also be represented by a graph surface area of the solid against $\frac{1}{\text{time}}$.

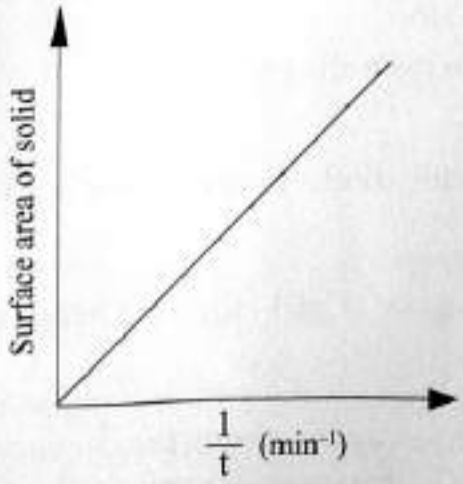

Fig. 1.10: Graph of surface area of solid reactants against reciprocal of time

This experiment, therefore, shows that the rate of reaction increases when surface area is increased.

Self assessment exercise 1.4

In an experiment, a student from Mzimba Secondary School, measured the amount of hydrogen produced when excess 1M hydrochloric acid was reacted with 0.3 g of zinc granules and obtained the following results at room temperature and pressure.

Time (in minutes)	Volume of hydrogen (cm^3)
0.0	0
0.5	20
1.0	32
2.0	52
5.0	86
7.5	103
10.0	112
12.0	118
14.0	120
16.0	120

Questions

(i) Draw a graph of volume of hydrogen produced against time.

(ii) Determine the time at which half of the original mass of zinc granules will have reacted.

(iii) Sketch the approximate curves on the same axes that would represent the reaction of excess 1M hydrochloric acid and 0.3g of zinc powder and 0.3g of zinc foil.

(d) Catalyst

Experiment 1.4

Aim: *To investigate the effect of manganese dioxide on rate of decomposition of hydrogen peroxide.*

Apparatus and reagents

- Boiling tubes
- Wooden splint
- 20 cm^3 hydrogen peroxide (H_2O_2)
- Manganese dioxide (MnO_2)
- Filter funnel

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- Filter papers
- Conical flask
- Weighing balance

Procedure

1. Dilute 20 cm^3 of hydrogen peroxide using 80 cm^3 of distilled water.
2. Put some 20 cm^3 of this diluted hydrogen peroxide in a boiling tube. Test whether oxygen is produced from this solution using a glowing splint.
   - Record your observation.
3. Put another 20 cm^3 of the diluted hydrogen peroxide in another boiling tube. Add about 0.5 g of manganese dioxide in this solution. What do you observe?
   - Test the gas evolved with a glowing splint and record your observation.
4. Filter the residue in (3) above using a filter paper of known mass. Wash the black residue (manganese dioxide) on the filter paper by pouring distilled water on it several times. Dry the filter paper. Weigh it after it has dried.

Study questions

1. What is the mass of the black solid?
2. Did you recover all the original manganese dioxide?
3. What is your conclusion from this experiment?
4. Write the chemical equation for this reaction.

Before manganese dioxide is placed into hydrogen peroxide, the solution gives off oxygen *very slowly*, not enough to re-light a glowing splint.

When manganese dioxide is added, the reaction becomes more vigorous and oxygen is obtained much *more quickly*. The oxygen is now enough to re-light a glowing splint.

The mass of manganese dioxide before and after the experiment is approximately the same. Manganese dioxide, therefore, acts as a **catalyst** in this reaction. It increases the rate of the chemical reaction but remains chemically unchanged at the end of the reaction i.e.

$$\text{Hydrogen peroxide} \xrightarrow[\text{oxide}]{\text{manganese}} \text{water} + \text{oxygen}$$

$$2H_2O_2(aq) \xrightarrow{MnO_2(s)} 2H_2O(l) + O_2(g)$$

Work to do

Find out another catalyst that can be used in the decomposition of hydrogen peroxide.

As discussed earlier, for a reaction to take place, reactants must have enough activation energy. A catalyst **lowers the activation energy** needed by the reactants to react.

Hence, more particles can get over the new activation energy barrier faster resulting in more fruitful collisions. This increases the rate of reaction. Fig. 1.11 is an analogy of how catalysts act to speed up rates of chemical reactions.

Fig. 1.11: Catalyst route is shorter than the route without a catalyst

In the analogy, the catalyst route is easier (requires less energy) than the route without a catalyst. The vehicle following this route will therefore reach its destination faster and will require less fuel because it will not climb the hill. The same applies to a reaction where a catalyst is involved. The catalyst lowers the activation energy required for the reaction to start. The reaction therefore takes place at a faster rate. Study Fig. 1.12.

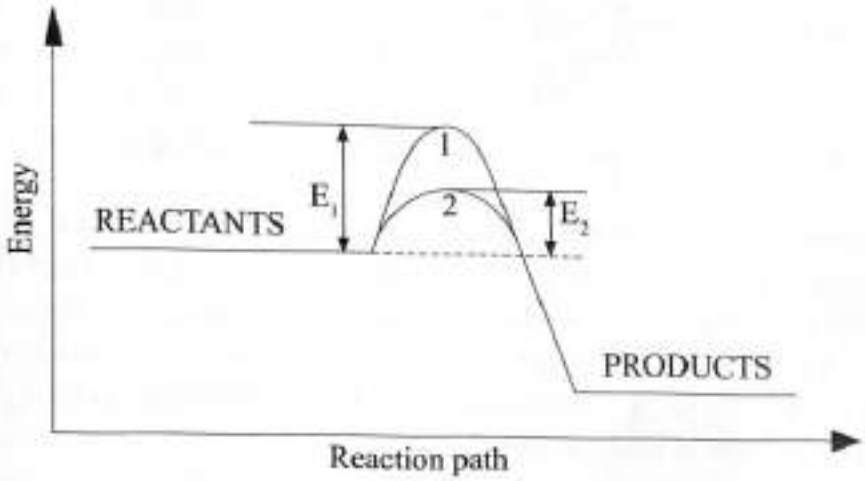

Fig. 1.12: Catalyst lowers the activation energy of the reaction

- Curve 1 represents the reaction without a catalyst. It requires activation energy E_1. Curve 2 represents the reaction with a catalyst. It requires activation energy E_2.
- According to the above graph, E_1 is greater than E_2. Therefore, a reaction in which a catalyst is involved requires less activation energy.

Note: A catalyst does not cause the reaction. It only speeds up the rate of a chemical reaction that would have otherwise taken place at a much slower rate in its absence.

Self assessment exercise 1.5

1. (i) Define the term catalyst.
   (ii) Give an example of a chemical reaction where a catalyst is used. Name the catalyst.
2. How does the catalyst in 1(ii) above increase the rate of reaction?

(e) Pressure

Look at the syringe in Fig.1.13 (a) below. If the end is sealed, how can you increase the pressure of the gases inside the syringe?

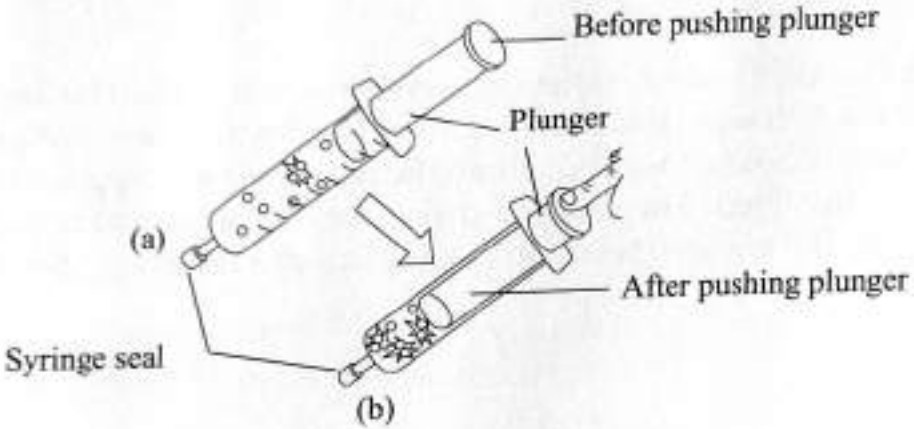

Fig. 1.13: Syringes showing how pressure can be increased (a) low pressure, (b) high pressure

To increase pressure, you push the plunger downwards as shown in fig. 1.13(b). By pressing the plunger in, you will have the same number of gas particles in a smaller volume. This reduces the space within which collisions takes place. The net effect is that more particles occupy smaller volume and therefore more collisions occur between particles. This increases the chances of fruitful collisions in a given time. As a result, the rate of reaction is increased.

Note: In reactions involving gases, increasing the pressure increases the rate of reaction. This however does not apply for liquids and solids.

Self assessment exercise 1.6

How does increase in pressure affect the rate of reaction for gaseous reactants? Explain using collision theory.

(f) Light

Light can be used to initiate (start) certain chemical reactions. It acts as an external source of activation energy for the reaction. Just like increase in temperature, increase in intensity of light increases the amount of fruitful collisions resulting in increase in the rate of a reaction. Some examples of reactions where light is involved include photosynthesis in plants and in photography.

1.3 Methods of measuring rates of reactions

The method chosen in the laboratory to measure reaction rate must consider a change that is taking place in a reasonable amount of time. The interval chosen should not be too long or too short depending on the rate of change. Also, the method should give an observable change during the reaction time.

Reaction rates can be determined by measuring one or some of the following aspects:

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

1.4 Reversible and irreversible reactions

A **reversible reaction** is a chemical reaction in which the products can be converted back to reactants under suitable conditions. Many reactions occur 'one-way only' that is, the starting materials react to form the product and the products cannot easily be turned back into the starting materials.

However in reversible reactions, products can react to form the reactants again.

Experiment 1.5

Aim: *To investigate what happens when hydrated copper sulphate is heated.*

Apparatus and reagents

- Test tubes
- Droppers
- Source of heat (Bunsen burner)
- Distilled water
- Hydrated copper (II) sulphate
- Thermometer
- Spatula
- Pair of tongs

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Procedure

1. Put two spatula endfuls of hydrated copper (II) sulphate into a test tube.
   - What is the colour of this solid?

(f) Light

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- Thermometer
- Spatula
- Pair of tongs

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Procedure

1. Put two spatula endfuls of hydrated copper (II) sulphate into a test tube.
   - What is the colour of this solid?

2. Heat the test tube gently while holding the test tube using pair of tongs as shown in Fig. 1.14 below. Stop heating the contents when the crystals have changed colour. Record this colour in your notebook.

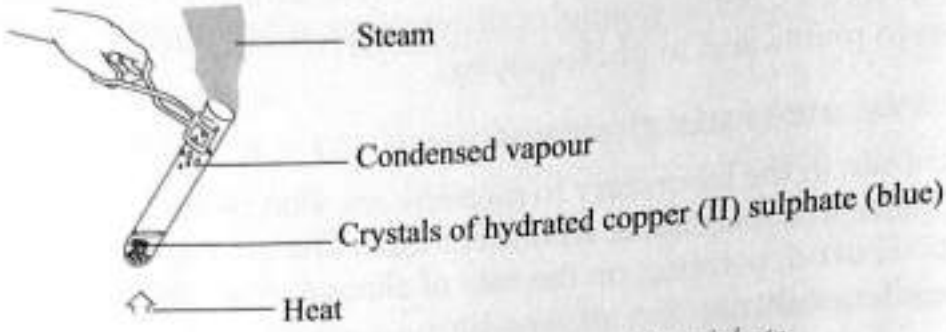

Fig. 1.14: Heating hydrated copper (II) sulphate

3. Let the test tube to cool down and note the colour of the copper (II) sulphate in the tube.
4. Now, add a few drops of distilled water from a dropper and note the colour of the solid.Measure the temperature of the resulting solution.

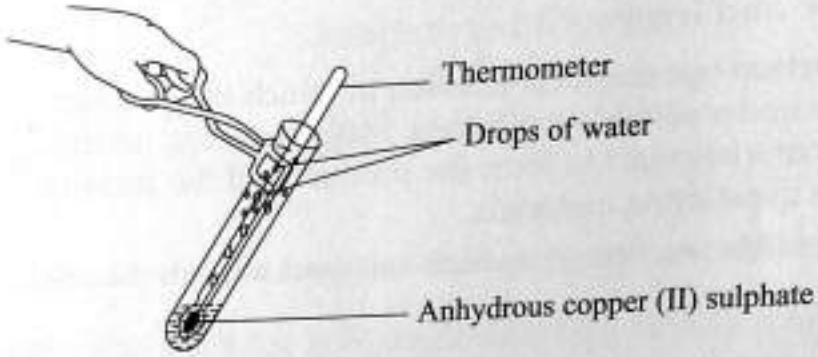

Fig. 1.15: Adding water to anhydrous copper (II) sulphate

- What is your conclusion from this experiment?

When **blue** hydrated copper (II) sulphate solid is heated, it loses its water of crystallisation. A **white** anhydrous copper (II) sulphate powder is formed. The colour persists even on cooling the tube.

Hydrated copper (II) sulphate ⟶ anhydrous copper (II) sulphate + water

$$\underset{\text{(blue crystals)}}{CuSO_4.5H_2O(s)} \longrightarrow \underset{\text{(white powder)}}{CuSO_4(s)} + 5H_2O(l)$$

When water is added to the white anhydrous copper (II) sulphate powder, its colour turns back to blue i.e. it becomes hydrated copper (II) sulphate. This is because the **water of crystallisation** is regained. This reaction is exothermic and therefore produces

a lot of heat.

A thermometer can be used to indicate a sudden rise in temperature when drops of water are added.

Anhydrous copper (II) sulphate + water $\longrightarrow$ hydrated copper (II) sulphate

$$\underset{\text{(white powder)}}{CuSO_4(s)} + 5H_2O(l) \longrightarrow \underset{\text{(blue crystals)}}{CuSO_4.5H_2O(s)} + \text{heat}$$

The overall reaction is given by:

Hydrated copper (II) sulphate $\rightleftharpoons$ anhydrous copper (II) sulphate + water

$$\underset{\text{(blue crystals)}}{CuSO_4.5H_2O(l)} \rightleftharpoons \underset{\text{(white powder)}}{CuSO_4(s)} + 5H_2O(l)$$

This is an example of a **reversible reaction**. The reactions can go in either directions. Reversible reactions are shown by the arrows pointing in both directions.

Note: The backward reaction shown above can be used as a test for presence of water.

We can also use indicators to show reversible reactions.

Experiment 1.6

Aim: *To investigate change of colour of indicator in an acid and alkali media as an example of a reversible reaction.*

Apparatus and reagents

- Beakers and droppers
- 2M hydrochloric acid
- Distilled water
- 2M sodium hydroxide
- Phenophthalein or any other suitable indicator

Procedure

1. Put some distilled water in a beaker. Add three drops of phenolphthalein indicator or any other available indicators.
   - What is the colour of the indicator in water?
2. To another clean beaker rinsed with sodium hydroxide solution, add about 10 cm^3 of 2M sodium hydroxide solution.
3. Now, pour the phenolpthalein indicator solution from the first beaker into the sodium hydroxide solution as in step 2 above.
   - What do you observe?

- What can you conclude from the observations?

4. Get a third beaker with about 15 cm^3 of 2M hydrochloric acid in it. Pour the solution in 3 above into the beaker with the acid.
   - What do you observe?
   - What can you conclude from this observation?

Phenolphthalein indicator is **colourless** in acid and **pink** in alkali. The results from the above experiment have been summarised in Fig. 1.16 below.

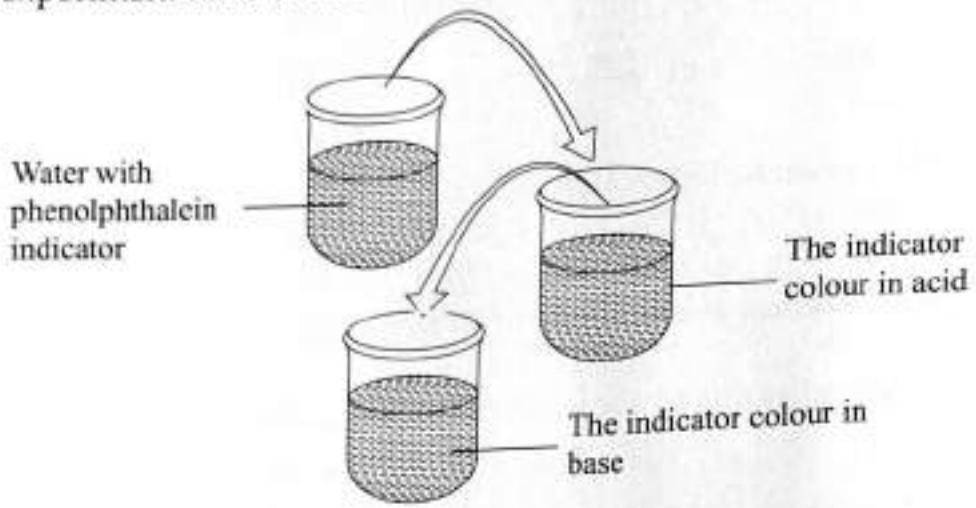

Fig. 1.16: Colour changes of indicator in acid or base

This is an example of a reversible reaction.

Irreversible reactions

There are some reactions that proceed forward only to form products. Once formed, products cannot react to form reactants. It is said once such reactions take place they cannot be reversed. They are said to be **irreversible reactions**. All combustion reactions are irreversible. Once the products are formed, the reactant cannot be re-formed. For example:

(a) When magnesium is burned in oxygen, magnesium oxide is formed. This is a permanent change and the reactants cannot be reformed.

Magnesium + oxygen ⟶ magnesium oxide

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

(b) When calcium carbonate is decomposed in an open system, the reaction cannot be reversed. This is because carbon dioxide is produced which escapes into the air.

Calcium carbonate $\xrightarrow{heat}$ calcium oxide + carbon dioxide

$CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$

However in a closed system, for instance, in a kiln the reaction can be reversed.

Self assessment exercise 1.7

1. Give two examples of irreversible reactions.
2. (a) If you placed water in a fridge, what would happen? Would you describe such a reaction as irreversible or reversible? Explain.
   (b) Suppose you heated the ice with a source of flame, would you describe the change as irreversible or reversible? Explain.

1.5 Chemical Equilibrium

What is chemical equilibrium? Let us find out by carrying out the following experiment.

Experiment 1.7

To investigate chemical equilibrium in chromate(VI) ion/dichromate(VI) ion.

Apparatus and chemicals

- 100 cm^3 glass beaker
- filter paper or a white paper
- droppers
- potassium chromate(VI) solution
- 2M sodium hydroxide
- 2M hydrochloric acid

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Procedure

1. Put 20 cm^3 potassium chromate(VI) solution in a glass beaker. Place the container on a filter paper or any other white paper.
   - What is the colour of the solution?
2. Add 2M hydrochloric acid solution drop by drop until there is a permanent colour change. Keep this solution for procedure 3.
   - What is the new colour of the solution?
3. To the same solution in procedure 2, add 2M sodium hydroxide solution drop by drop until there is no more change in colour.
   - What is the new colour of the solution?
   - What is your explanation and conclusion in this experiment?

Potassium chromate(VI) solution in water is yellow in colour. When a few drops of dilute hydrochloric acid are added to the solution, it turns orange due to formation of dichromate(VI) ions ($Cr_2O_7^{2-}$)(aq). This is because as the acid is added, the concentration of hydrogen ions (H^+) increases. An equilibrium is established with more $Cr_2O_7^{2-}$ in solution than CrO_4^-, hence orange colour is dormant.

However, on adding a few drops of 2M sodium hydroxide, the colour of the solution reverts to the original yellow colour.

Equilibrium shifts, on adding acid →

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$$

chromate (VI) ion (yellow) — dichromate (VI) ion (orange)

← *Equilibrium shifts, on adding a base*

When 2M sodium hydroxide is added, the OH^- ions react with hydrogen ions H^+ to form water. This reduces the concentration of hydrogen ions H^+ in the solution. The equilibrium shifts to the left. H^+ ions are formed to replace the hydrogen ions used to form water. The backward reaction is favoured and more chromate(VI) ions are formed turning the solution yellow.

$$Cr_2O_7^{2-}(aq) + 2OH^-(aq) \rightleftharpoons 2CrO_4^{2-}(aq) + H_2O(l)$$

(orange) (yellow)

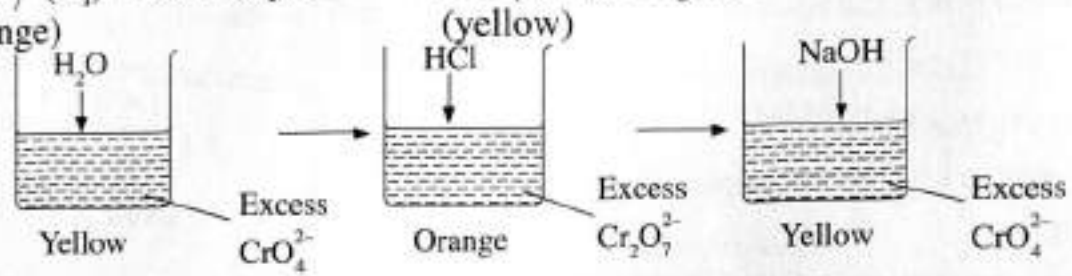


Fig. 1.18: Chemical equilibrium in chromate(VI)/dichromate(VI) mixture

Altering the concentration of any one of the components of the equilibrium mixture disturbs the equilibrium by affecting the rate of forward or reverse reaction. Hence, the reaction proceeds predominantly in that direction until a new equilibrium is established.

Note that the arrows showing reversible reaction at equilibrium are represented as $\rightleftharpoons$. They are slightly different from the ones shown earlier for reversible reactions represented as $\rightleftharpoons$.

In reversible reactions, change of reactants to products and change of the products back to reactants takes place at the same time i.e. are at equilibrium. In a reaction like the conversion of A and B to C and D predominates at the start of the reaction.

$$A + B \rightleftharpoons C + D$$

Eventually, both the forward and the backward reaction will be taking place at the same rate. As soon as C and D accumulate, the reverse reactions start building up until an equilibrium position or a state of 'balance' is reached. State of 'balance' means that the rate of forward reaction equals the rate of backward reaction i.e. at state of balance;

Rate of forward reaction = Rate of backward reaction

At state of balance; the reaction is said to have reached a **dynamic equilibrium** since the reactants and products are not simply coexisting but there is a constant interchange from products to reactants and vice versa.

The following experiment also explains

Experiment 1.8

To investigate change of colour of indicator in acid and alkali media as an example of a reversible reaction.

Apparatus and chemicals

- beakers
- 2M sodium hydroxide
- 2M hydrochloric acid
- phenophthalein or any other suitable indicator.

Procedure

1. Put some water in a beaker. Add three drops of phenolphthalein indicator to the water.
   - What is the colour of phenolphthalein indicator in water?
2. To another clean beaker, rinsed with sodium hydroxide solution, add about 10 cm^3 of 2M sodium hydroxide solution.
3. Now, pour the phenolpthalein solution from the first beaker into the sodium hydroxide solution in step 2 above.
   - What do you observe?
   - What can you conclude from the observations?
4. Get a third beaker with about 15 cm^3 of 2M hydrochloric acid in it. Pour the solution in 3 above into the beaker with the acid.
   - What do you observe?
   - What can you conclude from this observation?

Phenolphthalein indicator is colourless in acid and pink in alkali. The results from the above experiment have been summarised in Fig. 1.18 below.

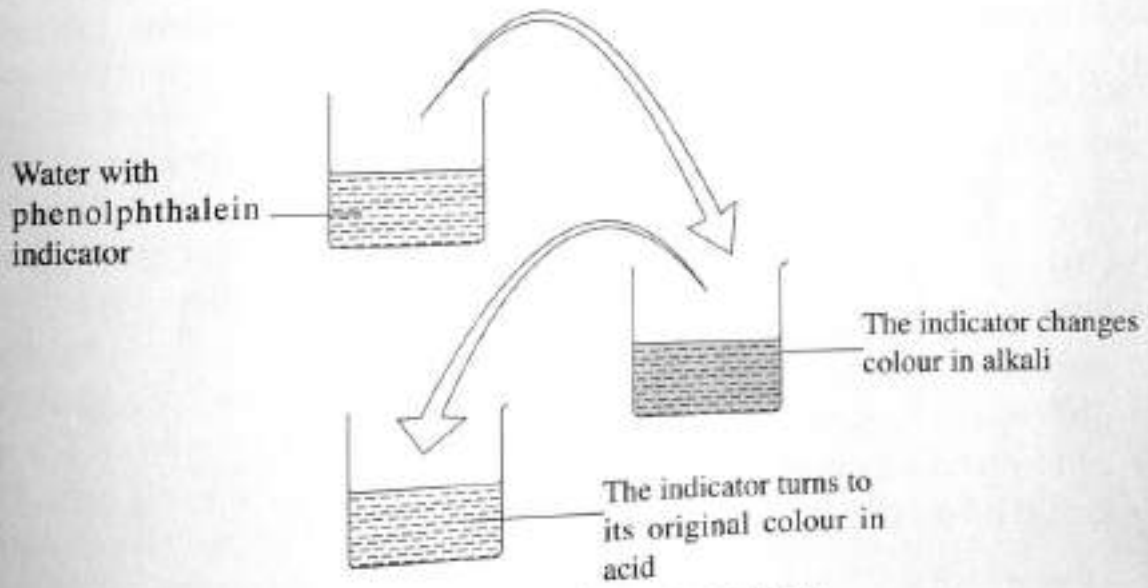


Fig. 1.18: Colour changes of indicator in acid/ base

Many reactions are 'one-way only'. The starting materials react to form the product. The products cannot easily be turned back into the starting materials.

However in reversible reaction, products can reform the reactants. In a closed system that is if you do not let any reactants or products escape, both forward and backward reactions can occur at the same time. Reactants make products, and at the same time, products make reactants. Eventually, both the forward and the backward reactions will be going on at the same rate.

When this happens, we say it has reached a state of "balance". This is called **chemical equilibrium**.

At the point of equilibrium, the rate of the forward reaction equals

the rate of backward reaction.

Self assessment exercise 1.8

1. What is a chemical equilibrium?
2. Give an example of a chemical reaction which eventually attains chemical equilibrium.

1.6 Effects of different factors on the position of equilibrium

The equilibrium position raises many interesting questions. For instance,

- What factors affect the position or 'balance' of chemical equilibrium?
- What is the change when some of the factors are altered?
- What is the composition of the mixture of the old and new equilibriums after one of the conditions is altered?

Let us consider what happens to the position of equilibrum when concentration, temperature and pressure are changed respectively.

A. Effect of changing concentration on the position of equilibrium

Experiment 1.9

To investigate how concentration affects the equilibrium of bromine and water

Apparatus and chemicals

- beaker
- filter papers or any white paper
- dropper
- 2 M sodium hydroxide
- 2 M hydrochloric acid
- Liquid bromine or 1% bromine water

Procedure

1. Prepare 1% bromine water by adding 1 cm^3 of liquid bromine to 99 cm^3 of distilled water and shaking the mixture well to dissolve.

   Caution: This should be done in a fume chamber or open space. Bromine vapours are poisonous.
2. Put 20 cm^3 of 1% bromine water in a glass beaker. Place the beaker and its contents on a filter paper or any white paper.
3. Add 2M sodium hydroxide solution drop by drop to the solution above while swirling until there is no further observable colour change. Keep this solution for procedure 4.
4. To the same solution in step 3, add 2M hydrochloric acid drop by drop while swirling until there is no further observable colour change.
   - What is the colour of 1% bromine water?
   - What is the effect of adding in turns, sodium hydroxide and hydrochloric acid to bromine water?
   - What is your explanation and conclusion in this experiment?

A. Bromine water is **yellow-orange** in colour. When the mixture is at equilibrium, it has bromine molecules, water molecules, hypobromite ions, bromide ions and hydrogen ions i.e.

$$Br_2(aq) + H_2O(l) \rightleftharpoons OBr^-(aq) + Br^-(aq) + 2H^+(aq)$$

yellow-orange colourless

Bromine molecules make the mixture yellow-orange.

When we add sodium hydroxide solution; the OH^- ions react with H^+ ions in the system to form water i.e.

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

This decreases the hydrogen ions concentration in the system. To cancel this change, the equilibrium shifts to the right to replace H^+ ions used and hence forward reaction is favoured. This produces more hydrogen ions which are added to the system. This shift converts the bromine molecules to hypobromite and bromide ions and hence the change in colour of solution from yellow-orange to colourless.

Addition of hydrochloric acid provides hydrogen ions (H^+), i.e the concentration of H^+ in the system is increased. To cancel this change, the equilibrium shifts to the left. The backward reaction favours the conversion of the additional hydrogen ions to water. In the course of this, bromine molecule are produced making the colour of the mixture to change from colourless to yellow-orange.

Changing the concentration of any one of the components of the equilibrium mixture disrupts the equilibrium by either changing the rate of forward or reverse reaction. Hence the reaction adjusts itself until the rate on both directions are equal to establish a new equilibrium.

B. Another good example that we have already discussed which demonstrates the effect of concentration on position of equilibrium is chromate(VI) ion/ dichromate(VI) ion mixture. The addition of hydrochloric acid (HCl) increases the concentration of hydrogen ions (H^+). The equilibrium shifts to the right to convert the added hydrogen ions to water. More dichromate(VI) ions are also formed making the solution orange.

When hydroxide ions (OH^-) are added through addition of 2M sodium hydroxide, they react with hydrogen ions to form water. This makes the concentration of hydrogen ions go down. The equilibrium shifts to the left to form more hydrogen ions to replace the ones converted to water. More chromate(VI) ions are formed making the solution yellow.

Self assessment exercise 1.9

$Br_2(aq) + H_2O(l) \rightleftharpoons OBr^-(aq) + Br^- + 2H^+(aq)$. What is the effect of adding potassium hydroxide when the above system is in chemical equilibrium?

B. Effect of changing temperature on the position of equilibrium

Experiment 1.10

To investigate the effect of temperature change on equilibrium.

Apparatus and chemicals

- Thistle funnel or dropper
- Flat-bottomed flask
- Delivery tube
- Boiling tube
- Copper turnings
- Concentrated nitric acid

Procedure

1. Put copper turnings in a flat-bottomed flask.
2. Arrange the apparatus as shown in Fig. 1.19.

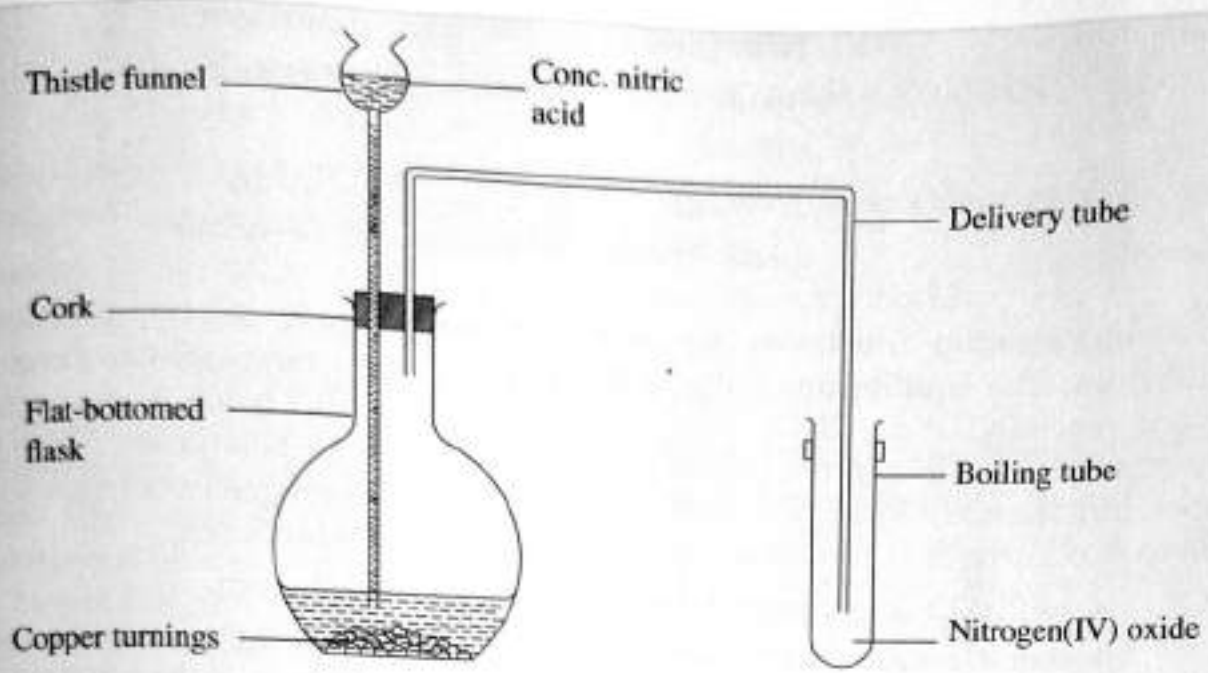


Fig. 1.19: Preparation of nitrogen(IV) oxide

NB: This should be done in a fume cupboard. NO_2 is poisonous.

3. Add concentrated nitric acid to the copper turnings in the flask.
4. Collect the gas in a boiling tube with some air and cork it tightly.
   - What is the colour of the gas in the boiling tube?
5. Slightly warm the gas in the boiling tube and note the colour. Record this in your notebook. Cool it in ice cold water as shown in Fig. 1.20 and record any colour change.

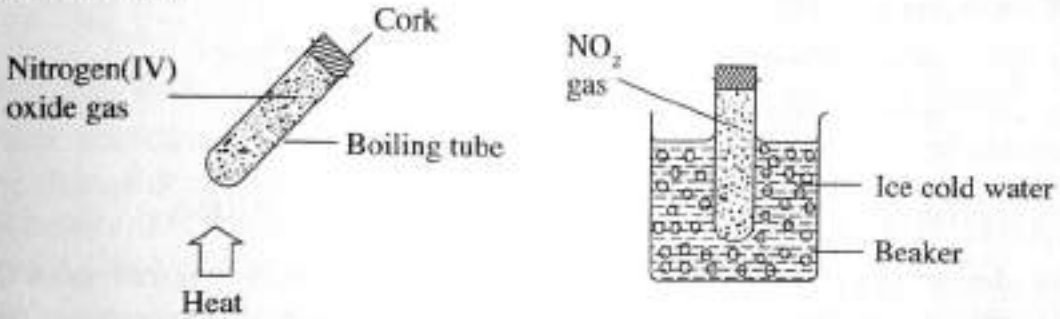


Fig. 1.20: Effect of temperature on position of equilibrium

Concentrated nitric acid reacts immediately with the copper to form red-brown fumes of nitrogen(IV) oxide.

Copper + nitric(V) acid ⟶ copper(II) nitrate + water + nitrogen(IV) oxide

$$Cu(s) + 4HNO_3(conc.) \longrightarrow Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

When air and nitrogen(IV) oxide at room temperature are heated gently then cooled in cold water, the colour changes from dark-brown to pale yellow. This is because the nitrogen(IV) oxide is converted to dinitrogen tetra-oxide. The colour of nitrogen(IV) oxide is dark brown and that of dinitrogen tetra-oxide is pale yellow.

$$\text{Nitrogen(IV) oxide} \underset{\text{increase temp.}}{\overset{\text{decrease temp.}}{\rightleftharpoons}} \text{dinitrogen tetra-oxide}$$

$$\underset{\text{(dark-brown)}}{2NO_2(g)} \underset{\text{increase temp.}}{\overset{\text{decrease temp.}}{\rightleftharpoons}} \underset{\text{(Pale-yellow)}}{N_2O_4(g)}$$

This demonstration illustrates the effect of temperature on the position of equilibrium. The equilibrium shifts to the right at lower temperatures favouring forward reaction, i.e (i.e $2NO_2 \longrightarrow N_2O_4$). Nitrogen(IV) oxide is converted to dinitrogen tetra-oxide and the colour of the gas changes to pale-yellow. At higher temperature the equilibrium will shift to the left and the dinitrogen tetra-oxide will be converted to nitrogen(IV) oxide i.e the backward reaction is favoured ($N_2O_4 \longrightarrow 2NO_2$). The colour of the gas changes to dark-brown.

C. Another example which shows that increasing the temperature causes the equilibrium of a reaction to shift in the direction that absorbs heat, is the production of sulphur(VI) oxide (SO_3) which is an exothermic reaction.

$$\text{Sulphur(IV) oxide + oxygen} \rightleftharpoons \text{Sulphur(VI) oxide + heat}$$

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + \text{heat}$$

Heating the reaction mixture at equilibrium shifts the equilibrium to the left. Cooling shifts the equilibrium to the right.

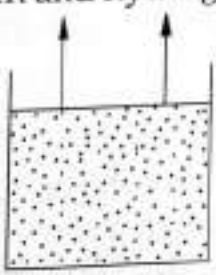
Self assessment exercise 1.10

1. What is the effect of increasing the temperature in
   (i) an endothermic reaction which is in equilibrium?
   (ii) an exothermic reaction in equilibrium?

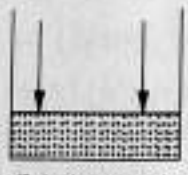
C. Effect of changing pressure on the position of equilibrium

Effects of changing pressure on the equilibrium applies only to substances that are in gaseous form and not those in liquid or solid forms. However, the number of moles of gases on either side of the equation must be different. If you increase the pressure as shown in Fig. 1.21 (b), the equilibrium shifts so as to try and reduce it.

We are going to illustrate the effects of pressure on the equilibrium by considering the reaction of nitrogen and hydrogen to form ammonia.



(a) decreasing



(b) increasing

Fig. 1.21: Changing pressure

Nitrogen + hydrogen $\rightleftharpoons$ ammonia + heat

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92kJ/mol$$

1 mole 3 moles 2 moles

The reactants are 4 moles (2.4×10^{24} molecules) whereas the product is two moles (1.2×10^{24} molecules). The more the molecules the higher the pressure.

(Remember 1 mole = 6.023×10^{23} molecules, i.e. Avogadro's number L)

Reactant(s) $\rightleftharpoons$ product

(high pressure) (low pressure)

In this reaction, there is a change of pressure when the product is being formed. Changing the pressure in a reaction like this one; which is in equilibrium is applying a stress to the system. This imposed stress has to be reduced. Increase in pressure will, therefore, cause the equilibrium to shift in the direction which will lower the increased pressure. This means that forward reaction will be favoured and hence more ammonia gas will be formed. The equilibrium therefore shifts to the right.

D. Other examples where the equilibriums shift from the left to the right with increase in pressure include:

(1) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

(2) $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$

However, in other reactions i.e where the number of moles on the right side of the equation is higher than the left side, the equilibrium will shift from right to left when the pressure is decreased. For example.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

1 mole 2 moles

In this reactions, more product can be obtained by decreasing the pressure.

E. In other reactions where the number of moles on either side of the reaction is the same, change in pressure will have no effect on the composition of the mixture at equilibrium. The position of equilibrium is therefore not altered. The following are examples of such reactions:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

$$CO(g) + NO_2(g) \rightleftharpoons CO_2(g) + NO(g)$$

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

Le Chatelier's Principle

From previous examples, you noticed that the addition of one or more of the substances which appear on one side of the equation always displaces the position of equilibrium away from that side. The temperature change also affects the equilibrium mixture. If the reaction involves gases, changing the pressure might affect the equilibrium as well.

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$$\text{Nitrogen + hydrogen} \rightleftharpoons \text{ammonia + heat}$$

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92kJ/mol$$

1 mole 3 moles 2 moles

The reactants are 4 moles (2.4×10^{24} molecules) whereas the product is two moles (1.2×10^{24} molecules). The more the molecules the higher the pressure.

(Remember 1 mole = 6.023×10^{23} molecules, i.e. Avogadro's number L)

$$\text{Reactant(s)} \rightleftharpoons \text{product}$$

(high pressure) (low pressure)

In this reaction, there is a change of pressure when the product is being formed. Changing the pressure in a reaction like this one; which is in equilibrium is applying a stress to the system. This imposed stress has to be reduced. Increase in pressure will, therefore, cause the equilibrium to shift in the direction which will lower the increased pressure. This means that forward reaction will be favoured and hence more ammonia gas will be formed. The equilibrium therefore shifts to the right.

D. Other examples where the equilibriums shift from the left to the right with increase in pressure include:

(1) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

(2) $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$

However, in other reactions i.e where the number of moles on the right side of the equation is higher than the left side, the equilibrium will shift from right to left when the pressure is decreased. For example.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

1 mole 2 moles

In this reactions, more product can be obtained by decreasing the pressure.

E. In other reactions where the number of moles on either side of the reaction is the same, change in pressure will have no effect on the composition of the mixture at equilibrium. The position of equilibrium is therefore not altered. The following are examples of such reactions:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

$$CO(g) + NO_2(g) \rightleftharpoons CO_2(g) + NO(g)$$

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

Le Chatelier's Principle

From previous examples, you noticed that the addition of one or more of the substances which appear on one side of the equation always displaces the position of equilibrium away from that side. The temperature change also affects the equilibrium mixture. If the reaction involves gases, changing the pressure might affect the equilibrium as well.

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Fig. 1.22: Henri Le Chatelier 1850 – 1936

Such factors which affect the position of equilibrium in a reaction were first investigated by a French Chemist, Henri Louis Le Chatelier. He expressed his results in a generalised way which later came to be known as Le Chatelier's Principle.

Le Chatelier's Principle

"It states that if a system in equilibrium is subjected to a stress, the equilibrium will shift in the direction which tends to relieve that stress"

This principle can be expressed in the following different ways without necessarily changing its meaning.

1. If a system in equilibrium is subjected to a change, the position of equilibrium will move in such a way as to tend to eliminate the change.
2. If a system in equilibrium is subjected to a change, processes will occur which tend to counter the change imposed etc.

Le Chatelier's Principle can be used to predict the effect of changing concentration, temperature or pressure in any chemical system in equilibrium. This principle has been applied in reversible reactions in industrial processes. Let us now look at examples of some industrial applications.

1.7 Uses of reversible reactions in industrial processes

Many industrial processes involve **reversible reactions** and since cost reduction is an extremely important consideration in industry, a great deal of care is taken when determining the conditions that affect these processes. It is important to maximise the concentration of the desired products and minimise the 'left over' reactants. A set of rules can be used to predict the best reaction conditions to give the highest possible yield of the products. The rules are based on the conditions that affect the reactions involved as discussed below.

Rule 1: Based on temperature

(a) If the forward reaction is **endothermic**, raising the temperature favours increased yield of the product. Lowering the temperature decreases the yield of the product.

(b) If the forward reaction is **exothermic**, lowering the temperature favours increased yield of the product. Raising the temperature decreases the yield.

Note that rule 1 applies to any reaction.

Rule 2: Based on pressure

(a) Increasing the pressure favours the side of the equilibrium with the least number of gaseous molecules.

(b) Decreasing the pressure favours the side of the equilibrium with the most number of gaseous molecules.

> Rule 2 applies to reactions involving one or more gaseous reactants or products only.

Rule 3: Based on the concentration of reactants

(a) If the concentration of any reactant on the left is increased then some of it must be used up to form products on the right to maintain a balanced equilibrium position.

(b) If the concentration of any reactant on the left is decreased then some of the products on the right must change back to the reactants to maintain a balanced equilibrium

> Rule 3 applies to any reaction.

Rule 4: Based on catalyst

A catalyst does not affect the position of an equilibrium. It only makes the forward and reverse reaction take place at a faster rate. In this way, it helps the equilibrium to be established faster. But once the equilibrium has been established, it cannot influence the final composition of the equilibrium mixture.

The following are some of the ways in which reversible reactions are applied in industry.

(a) Manufacture of concentrated sulphuric(VI) acid (Contact process)

- In the manufacture of sulphuric(VI) acid, the main reaction that converts sulphur(IV) oxide (SO_2) to sulphur(VI) oxide(SO_3) is reversible.

  Sulphur(IV) oxide + oxygen $\rightleftharpoons$ sulphur(VI) oxide + heat energy

  $$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g); \Delta H = -197kJ/mol$$

  - The reaction of sulphur(IV) oxide and oxygen to form sulphur(VI) oxide is an exothermic reaction.
  - 3 moles of reacting gases form 2 moles of the product, so there is net decrease in the number of moles on forming sulphur(VI) oxide.
  - So applying the equilibrium rules, the formation of sulphur(VI) oxide is favoured by:

1. **Increasing pressure** because there is a decrease of 3 moles of reacting gases to 2 moles of the products formed.
2. **Lowering temperature** because the reaction is exothermic.

Since the reaction is exothermic, once the process has started, no more heating is required i.e the reaction is self-sustaining.

According to the above equation, increase in pressure will give an increased yield of sulphur(VI) oxide (favour forward reaction). The effect is small and the yield of sulphur(VI) oxide is good even at low pressures. Therefore 1-2 atmospheres pressure is used in practice.

- Can you predict the optimum temperature for a high yield of sulphur(VI) oxide?
- Should the system be allowed to attain equilibrium at low or high temperature?

Application of Le Chatelier's Principle suggests that the lower the temperature the more the equilibrium state will favour the production of sulphur(VI) oxide. But, what about rate or time required to produce a given quantity? At low temperatures, reaction rates are slow. A compromise temperature of about 450 – 500°C is usually used to give desirable yield of SO_3 and a satisfactory rate of reaction.

You recall that during contact process, sulphur(VI) oxide is reacted with concentrated sulphuric acid to form oleum. This reduces the concentration of SO_3 in the system. This removal of sulphur(VI) oxide creates stress and shifts the equilibrium to the right.

A catalyst of vanadium(V) oxide is also used to increase the rate of reaction. It does this by reducing the activation energy required to initiate the reaction. A yield of about 98% is obtained when all these conditions are applied.

(b) Manufacture of ammonia (Haber process)

Ammonia gas is synthesised in the industry by reacting nitrogen gas and hydrogen gas through a process called Haber process. The reaction is reversible as shown below.

Nitrogen + hydrogen $\rightleftharpoons$ ammonia + heat

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92kJ/mol$$

1 mole 3 moles 2 moles

This reaction is an exothermic reaction. The four moles of 'reactants' (1 mole of N_2 and 3 moles of H_2) form two moles of 'product' gas (2 moles of NH_3), so there is a net decrease in gas molecules on forming ammonia.

Applying the earlier discussed equilibrium rules, the formation of ammonia is favoured by:

1. **High pressure.** Because we are going from 4 to 2 moles. High pressure increases the number of collisions of gaseous particles hence higher rate of foward reaction to form ammonia
2. **Low temperatures.** Because the reaction is an exothermic reaction, high temperatures would decompose ammonia, the product we want.

Unfortunately, at low temperatures and high pressure, the rate at which equilibrium is reached is very slow. Also, if very high pressures are used, an expensive industrial plant will be required. In practice, 'optimum conditions are applied'.

Theoretically, a low temperature would give a high yield of ammonia but nitrogen is a very stable molecule and therefore not very reactive especially at low temperatures. A compromise temperature of 400 – 450°C is used. To speed up the reaction, finely divided iron catalyst is used. Using the catalyst in this process does not affect the position of equilibrium but the reaction reaches equilibrium faster.

Table 1.4 Summary of effects of changing some factors on chemical equilibrium

Factor	Change		Effect
Concentration	Increase — for reactants or (decrease for products)		• Shifts from left to right (⟶) • Favours forward reaction (⟶) • More products formed
	Decrease — for reactants or (increase for products)		• Shifts from right to left (⟵) • Favours backward reaction (⟵) • More reactants are formed
Pressure	Increase	Increase in number of moles on formation of products	• Shifts from right to left (⟵) • Favours reverse reaction (⟵) • More reactants formed in the reverse reaction
		Decrease in number of moles on formation of products	• Shifts from left to right (⟶) • Favours forward reaction (⟶) • More products formed from the reactants in forward reaction
	Decrease	Increase in number of moles on formation of products	• Shifts from left to right (⟶) • Favours forward reaction (⟶) • More products formed
		Decrease in number of moles on formation of products	• Shifts from right to left (⟵) • Favours reverse reaction (⟵) • More reactants are formed
Temperature	Increase	Exothermic reaction	• Shifts from right to left (⟵) • Favours reverse reaction (⟵) • More reactants are formed
		Endothermic reaction	• Shifts from left to right (⟶) • Favours forward reaction (⟶) • More products formed.
	Decrease	Exothermic reaction	• Shifts from left to right (⟶) • Favours forward reaction (⟶) • More products are formed
		Endothermic reaction	• Shifts from right to left (⟵) • Favours reverse reaction (⟵) • More reactants are formed.

Self assessment exercise 1.7

1. State three factors which determine the composition of equilibrium mixture.
2. Why does the catalyst have no effect on the composition of equilibrium mixture.
3. Explain how the following will affect the yield of ammonia.
   (i) Increase in pressure
   (ii) Increase in temperature
4. Use Le Chatelier's Principle to explain how removal of a little of D in the following reaction favours a higher yield of D.

$$A + B \rightleftharpoons D$$

Revision Exercise 1

1. (a) Define the following terms.
   i) Rate of reaction
   (ii) Activation energy
   (iii) Reversible reaction

   (b) 25.0 cm^3 portions of 0.2M sodium thiosulphate were put into conical flasks. The temperature of the solutions was recorded. A cross was drawn on a piece of white paper underneath the flask. An excess of 2M hydrochloric acid was added. The time taken for the cross to disappear was recorded. The experiment was then repeated with the solutions at different temperatures. The results obtained were as shown in the table below.

Temperature of sodium thiosulphate (°C)	19	29	40	49	60
Time (in seconds)	340	250	155	90	55

   (i) Plot a graph of time (vertical axis) against temperature (horizontal axis).
   (ii) From your graph, estimate the reaction time at
   I. 25°C
   II. 54°C

(iii) From the graph, estimate the temperature at which the reaction time is

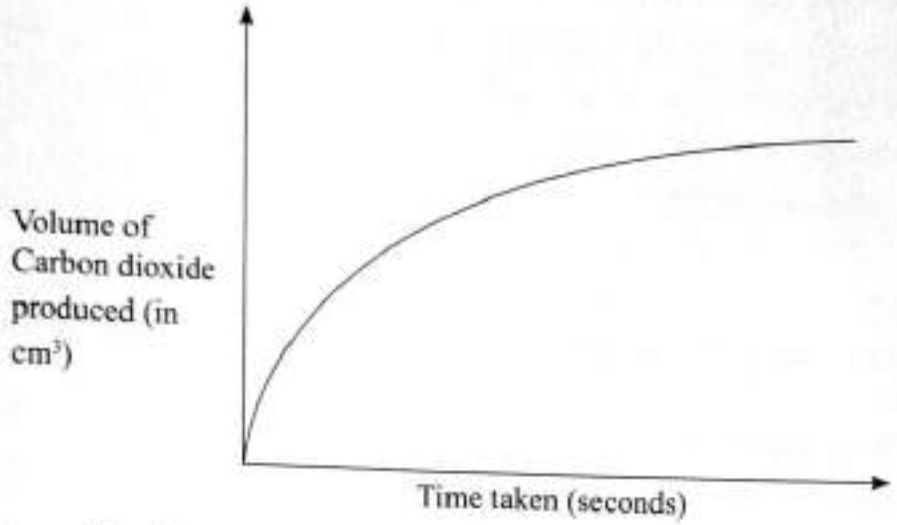
I. 110 seconds

II. 65 seconds

(iii) What effect does an increase in temperature have on the rate of the reaction? Explain your answer.

2. (a) Excess dilute nitric acid reacts more rapidly with powdered marble than with lumps of marble. In both reactions, the rates of reaction decrease with time. Explain these two observations.

(b) The graph below shows the volume of carbon dioxide produced and time taken when lumps of marble are reacted with excess 2M nitric acid.



(i) Sketch on the same axis, the curve that would be obtained when the same mass of marble powder and excess 2M nitric acid are reacted.

(ii) Write the chemical equation for the reaction.

3. What is a catalyst?

4. State the factors that can affect the rate of a chemical reaction.

5. (a) State Le Chatelier's Principle.

(b) Name three factors which may disturb a chemical equilibrium.

(c) What is the use of a catalyst in chemical reactions?

6. Explain briefly the effect of addition of aqueous potassium hydroxide on equilibrium of the following reactions.

$$2CrO^{2-}_4(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$$

What would happen when a few drops of dilute sulphuric acid are added to the mixture at equilibrium?

Topic 2 Acids and Bases

Success criteria

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

(a) Define acid and base.
(b) Identify conjugate acid-base pairs from given equations.
(c) Explain the difference between strength and concentration of an acid or a base.
(d) Explain ways of regulating pH in different environments.
(e) Classify oxides as acidic, basic or amphoteric.
(f) Describe different ways of preparing salts.
(g) Design an experiment to prepare and purify a salt.
(h) State applications of precipitation.

Introduction

In Form 1, you learnt about properties of acids and bases. Some of the properties of acids include:

- Have sour taste.
- Turn blue litmus red.
- Are neutralised by alkalis to form salt and water only.
- React with certain common metals to form salt of the metal and hydrogen gas.
- React with carbonate or hydrogen carbonates to form salt, carbon dioxide and water.

On the other hand, bases have properties which are the opposite to those of acids. They include:

- Turn red litmus blue and other indicators to their respective colour.
- Have a soapy feel.
- Are neutralised by acids to form salt and water only.

In this class, you will learn more about strengths of acids and bases and how to prepare salts.

2.1 Meaning of acid and a base

The two common definitions of the word acid are:

(a) A substance that reacts with a base to form salt and water only. For example:

Hydrochloric acid + zinc oxide $\longrightarrow$ zinc chloride + water

$$2HCl\,(aq) + ZnO(s) \longrightarrow ZnCl_2(aq) + H_2O(l)$$

Sulphuric acid + sodium hydroxide $\longrightarrow$ sodium sulphate + water

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

(b) A substance that dissolves in water to produce hydrogen ions (H^+) as the only positively charged ion, for example:

Hydrochloric acid + water $\longrightarrow$ hydrogen ion + chloride ion

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

Nitric acid + water $\longrightarrow$ hydrogen ion + nitrate ion

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

Sulphuric acid + water $\longrightarrow$ hydrogen ion + sulphate ion

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

Note: Acids such as hydrochloric acid which dissociate in water to give only one hydrogen ion are known as **monobasic acids**. Those that dissociate to give two hydrogen ions i.e. $2H^+$ like sulphuric acid are called **dibasic acids**. Others such as phosphoric acid (H_3PO_4) are **tribasic** meaning they dissociate to give three hydrogen ions, $3H^+$.

On the other hand, a base can be defined as:

(a) A substance which reacts with an acid to form salt and water only. For example:

Nitric acid + magnesium hydroxide $\longrightarrow$ magnesium nitrate + water

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

(b) A substance which dissociates in water to give hydroxide ions (OH^-) as the only negatively charged ions. For example:

Potassium hydroxide + water $\longrightarrow$ potassium ion + hydroxide ion

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

Bronsted - Lowry Acid and Base

Let us consider ionisation of hydrochloric acid as shown below.

$HCl(l) + H_2O(l) \longrightarrow Cl^-(aq) + H_3O^+(aq)$

(Hydroxonium ion)

During ionisation, an acid like hydrochloric acid provides a proton (H^+) or hydrogen ion to a water molecule, H_2O thus forming hydroxonium ion (H_3O^+) i.e.

$H_2O(l) + H^+(aq) \longrightarrow H_3O^+(aq)$

Because of this, Bronsted and Lowry in 1923 came up with a better definition of an acid as follows:

An acid is a substance which donates protons (H^+). As such, the definition of an acid according to Bronsted and Lowry is "**Proton donor**". On the other hand, a base is defined as a "**Proton acceptor**" because it accepts a proton from an acid.

2.2 Conjugate acids and bases according to Bronsted-Lowry Theory

Consider the following relationship between an acid and a base:

$$HB(aq) \rightleftharpoons H^+(aq) + B^-(aq)$$

(Acid) (proton-donor) (Base) (proton-acceptor)

HB has lost a proton (H^+) to form B^-. Therefore, according to Bronsted and Lowry theory, HB has donated a proton whereas B^- has accepted a proton. The substance that is produced after an acid has donated its proton i.e. B^- is called a **conjugate base** while the substance formed when a base accepts a proton, i.e. HB is a **conjugate acid**.

Other examples of conjugates are as follows:

(a) $HSO_4^-(aq) + OH^-(aq) \rightleftharpoons SO_4^{2-}(aq) + H_2O(l)$

(b) $NH_4^+(aq) + OH^-(aq) \rightleftharpoons NH_3(g) + H_2O(g)$

(c) $2H_3O^+(aq) + S^{2-}(aq) \rightleftharpoons 2H_2O(l) + H_2S\ g)$

(d) $H_3O^+(aq) + NH_3(g) \rightleftharpoons H_2O(l) + NH_4^+(aq)$

(e) $H_2O(l) + H_2O(l) \rightleftharpoons OH^-(aq) + H_3O^+(aq)$

Identify the conjugate acid and the conjugate base in the equations above.

Note: The conjugate acid can donate a proton to the conjugate base to reform the original conjugate reactants in the reverse reactions.

If aqueous solution is considered, the relationship between conjugate pairs is given by:

$$HB(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + B^-(aq)$$

In the above example, water is accepting a proton from acid (HB), thus acting as a base, while H_3O^+ ion is donating a proton to base B^- ions. Therefore B^- ions acts as a base. Hence H_3O^+ is a **conjugate acid** while B^- is a **conjugate base**.

Further examples are:

(a) $HAc(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Ac^-(aq)$

(b) $H_2SO_4(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$

(c) $NH_4^+(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NH_3(g)$

(d) $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

Identify the conjugate acids and bases in the above pairs.

Self assessment exercise 2.1

1. Giving a reason in each case, identify an acid and a base in the equation below.

   $NH_3(g) + H_2O(l) \longrightarrow NH^+_4(aq) + OH^-(aq)$

   (a) acid

   reason _____

   (b) base

   reason _____

2. Identify Bronsted-Lowry acid and base in the chemical reaction below.

   $HBr\ (aq) + H_2O(l) \longrightarrow H_3O^+\ (aq) + Br^-(aq)$

3. Considering the equation below, identify the reactant that acts as an acid and explain.

   $NH_4^+\ (aq) + H_2O\ (l) \longrightarrow NH_3\ (aq) + H_3O^+(l)$

2.3 Strength of acids and bases

(a) Strong and weak acids

An acid is described as strong or weak depending on its ability to give hydrogen ions when dissolved in water i.e. strength of an acid is based on the extent to which they dissociate into ions in aqueous solutions.

A strong acid is therefore one which dissociates completely in water to give all its hydrogen ions.

Three examples of strong acids are hydrochloric, nitric and sulphuric acids.

(i) $HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$

(ii) $HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$

(iii) $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}\ (aq)$

On the other hand, a weak acid is one which dissociate partially in aqueous solution. That is, it gives fewer hydrogen ions when dissolved in water than the total number of hydrogen atoms present in the molecule. Examples of weak acids are ethanoic and carbonic acid.

(i) $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

(ii) $H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$

The double half arrows show that the acid is **weak** and the process is **reversible**. The above reactions show that a weak acid contains more hydrogen atoms than the ions that are released.

Self assessment exercise 2.2

1. Write an equation for dissociation of each of the following acids:
   (a) Hydrochloric acid
   (b) Ethanoic acid
   (c) Sulphuric acid

(b) Strong and weak bases

Like acids, a base is described as strong or weak depending on its ability to yield hydroxide ions (OH^-) when dissolved in water.

A strong base is one which dissociates completely to produce all hydroxide ions when dissolved in water.

Examples of strong bases are sodium hydroxide and potassium hydroxide.

(i) $NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$

(ii) $KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$

A weak base dissociates partially to give fewer hydroxide ions when dissolved in water. An example is ammonium hydroxide.

(i) $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

(ii) $NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

Self assessment exercise 2.3

1. Write dissociation equations for each of the following bases:
   (a) Potassium hydroxide
   (b) Ammonium hydroxide
   (c) Lithium hydroxide

Experiment 2.1

Aim: *To determine the pH of weak and strong acids and bases*

Apparatus and reagents

- Beakers
- Droppers
- 2M hydrochloric acid

- 2M ethanoic acid
- 2M sodium hydroxide
- 2M ammonium hydroxide
- Universal indicator solution
- pH chart

Procedure

1. Place about 20 cm^3 of dilute hydrochloric acid, ethanoic acid, sodium hydroxide solution and ammonia solution in different beakers.
2. Add 2 - 3 drops of universal indicator solution. Note the colour of the solution.
3. Compare the colours of the solutions with colours given on the standard universal indicator colour chart to determine their pH values.
4. Record the pH values in your notebook as shown in Table 2.1 below.

Table 2.1: pH values for weak and strong acids and bases

Substance being tested	Colour obtained on adding universal indicator	pH value	Indicate whether they are weak or strong acid or base
2M Hydrochloric acid			
2M Ethanoic acid			
2M Sodium hydroxide			
2M Ammonia solution			

5. Classify the solutions into strong or weak acids and bases.

Note: Universal indicator paper can also be used. Dip the paper into the solution to be tested and then compare the colour change on the paper with the standard colour chart.

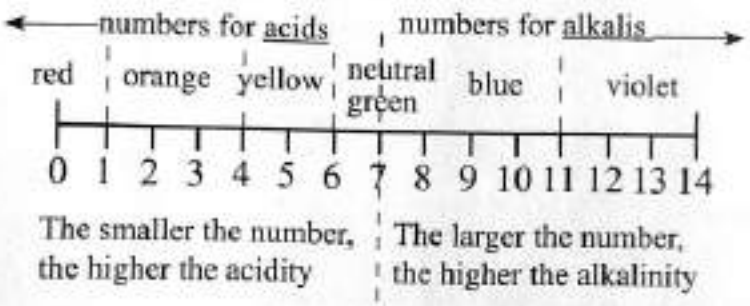


Fig. 2.1: pH scale

Strong acids have greater amounts of hydrogen ions (H^+) and their pH range from 1.0-3.5. Weak acids have fewer hydrogen ions and their pH range from 3.5-6.5. The pH of pure water is 7.0 and it is said to be ***neutral.***

Weak bases have fewer hydroxide ions and their pH range from 7.5 - 10.5. Strong bases have greater amounts of hydroxide ions and their pH range from 10.5 - 14.0. From experiment 2.1 the pH value of dilute hydrochloric acid shows that it is a strong acid while that of ethanoic acid shows that it is a weak acid. The pH value of dilute sodium hydroxide solution shows that it is a strong base while that of dilute ammonia solution shows that it is a weak base.

Note: The term *strong* applies to acids and bases when they are *completely dissociated in aqueous solutions*. It does not indicate the concentration of the acid or base.

Self assessment exercise 2.4

1. Define the following terms:
   (a) Strong acid
   (b) Weak base
2. Ethanoic acid turns blue litmus paper faintly red, hydrochloric acid turns it deeply red. Explain why this happens.
3. Is the pH value of a strong acid low or high?

Electrical conductivity of weak and strong acids and bases

Experiment 2.2

Aim: *To determine the electrical conductivity of weak and strong acids and bases*

Apparatus and reagents

- 250 cm^3 beaker
- Two 1.5 V dry cells
- 2.5 V bulb
- Carbon rods
- Connecting wires
- Switch
- 2M Hydrochloric acid
- 2M Ethanoic acid
- 2M Sodium hydroxide
- 2M Ammonium hydroxide

Caution: NEVER taste nor touch acids and alkalis with bare hands. They are corrosive!

Procedure

1. Put 50 cm^3 of 2M hydrochloric acid in the beaker.
2. Make a circuit using carbon rods, wire and switch as shown in Fig. 2.2.

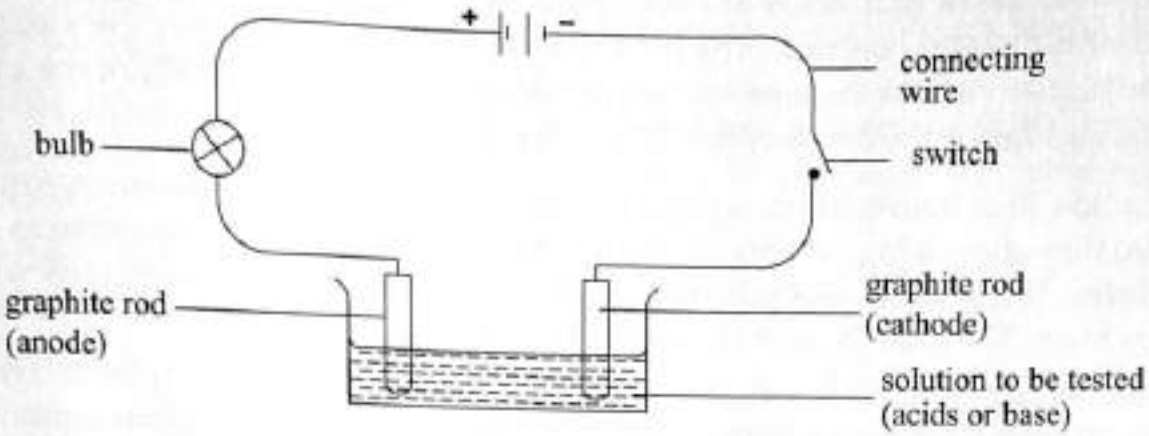


Fig. 2.2: Investigating electrical conductivity in acids and bases

3. Switch on the current.
   - Does the bulb glow?
   - Note the intensity of the light given out?
   - What do you observe on the carbon rods? What does this indicate?
4. Repeat the same experiment with 2M ethanoic acid, 2M sodium hydroxide solution and 2M ammonia solution. Use 50 cm^3 of each solution.

   Note: Rinse the beaker and the rods thoroughly with distilled water before adding the next solution.
5. Note the changes occurring around the carbon rods, in the contents of the beaker and in the brightness of the bulb.
6. Copy Table 2.2 in your notebook and fill in your observations and conclusions.

Table 2.2: Electrical conductivity of acids and bases

Solution	Brightness of the bulb (bright or faint)	Conclusion
2M Hydrochloric acid		
2M Ethanoic acid		
2M Sodium hydroxide		
2M Ammonia solution		

Ionic substances that conduct electric current in their aqueous solutions are called ***electrolytes***. Electrolytes can be acids or soluble bases. Soluble bases are known as ***alkalis***.

Electrolytes contain charged particles called ***ions*** which conduct electric current as they move in the solution.

Electrical conductivity depends on the strength of the acid or alkali used as an electrolyte. Strong acids and bases have higher conductivity compared to weak acids and alkalis. This is because strong acids and alkalis dissociate fully releasing all the ions which then conduct electric current. Weak acids and alkalis do not dissociate fully into ions and therefore fewer ions are available to conduct electric current.

When soluble bases and acids are dissolved in water, they dissociate to form ions which conduct electric current. The more the dissociation, the higher the yield of the ions. The higher yield of the ions enables greater electrical conductivity of the aqueous solutions and hence the brighter the light from the bulb.

Strong acids like hydrochloric acid, yield many ions when in aqueous form and therefore they show a high electrical conductivity. They are hence referred to as ***strong electrolytes***. Weak acids, like ethanoic acid (or acetic acid) yield fewer ions when in aqueous form. They are thus called ***weak electrolytes***.

Likewise, strong soluble bases like sodium hydroxide yield many ions when in aqueous solutions forming strong electrolytes. Weak alkalis like aqueous ammonia on the other hand give fewer ions and they form weak electrolytes.

Note: All electrolytes are decomposed during electrolysis to form various substances around the **electrodes.**

Difference between strength and concentration of acids and alkalis

Concentration of an acid or base refers to the mass of the acid or base (strong or weak) that is dissolved in a definite volume of solution (usually 1000 cm^3 to make up a total volume of 1 litre or 1dm^3). Strength of an acid or a base on the other hand is based on their ability to dissociate in water.

Consider these molar solutions:

(a) 1M HCl contains 36.5g/litre

(b) 1M HNO_3 contains 63g/litre

(c) 1M CH_3COOH contains 60g/litre

(d) 1M NaOH contains 40g/litre

(e) 1M KOH contains 56g/litre

The acids have equal molar concentration (i.e. molarity = 1M) but HNO_3 and HCl are strong acids while CH_3COOH (ethanoic acid or acetic acid) is a weak acid.

If in acidic solution, the proportion of acid is very high and that of water is very low, then it is referred to as a **concentrated acid**. On the other hand, if in an acid solution, the proportion of water is very high and that of acid is very low, it is said to be a **dilute acid.** Likewise, the molarity of sodium hydroxide might be the same, but sodium hydroxide is a **strong base** and ammonium hydroxide is a **weak base**.

In concentrated acids, the proportion of acid is very high and that of water is very low. In dilute acids, the proportion of water is very high and that of acid is very low. An acid or base can have high concentration but is still considered to be weak, for

example acetic acid and ammonia solution. An acid or base can be dilute but still said to be a **strong acid** or base, for example sulphuric acid and sodium hydroxide solution respectively.

2.4 Ways of regulating pH in different environments

In our everyday lives, we encounter various substances and environments that have different pH values.

The pH of different substances and environments has to be made stable for them to function efficiently. The level of pH hence has to be regulated. The environmental activities or cases that involve pH regulation include:

- stomach acid and indigestion.
- soil and water acidity.
- insect sting.

(a) Stomach acid and indigestion

The stomach contains hydrochloric acid. The acid provides acid medium required for protein digestion. Sometimes, excess hydrochloric acid is produced causing stomach ache. An "antacid" may be taken to treat the stomachache. The antacid contains a mild base i.e. milk for magnesia (or magnesium hydroxide solution) which helps to neutralise the excess acid. This helps in relieving the pain.

$$2HCl(aq) + Mg(OH)_2(aq) \longrightarrow MgCl_2(aq) + 2H_2O(l)$$

Stomach acid	milk of magnesia	salt	water

(b) Soil and water acidity

When fertilisers are used for a long period of time, the amount of acids in soils tend to increase. The acids are also washed into underground water sources thereby increasing the acid levels of the water. Also, when acidic gases dissolve in rain water, they make water more acidic. This is dangerous for the health of plants and animals.

To increase the soil pH, **lime** (calcium oxide) is added to the acidic soil. However, sometimes, soil pH also may increase above optimum levels. To lower soil pH, acid fertiliser such as ammonium sulphate is added to the soil to lower the pH. On the other hand, high acidity in water can be removed by adding equal amounts of a base e.g. sodium hydrogen carbonate.

(c) Insects sting

Ants, bees, and nettles are known to sting insects and animals. When they sting, they inject their victims with **methanoic acid**. This is a painful experience. A very dilute alkali can be applied on the affected area on the skin to neutralise the acid in the sting.

Self assessment exercise 2.5

1. State three applications of acid-base neutralisation.

2.5 Classification of oxides

Research Activity

Find out from textbooks in the library or the internet the various categories of oxides. Write a report and present to the rest of the class.

An **oxide** is a binary compound formed when oxygen reacts with another element. Such elements can be non metals like carbon, sulphur and phosphorus or metals such as sodium, magnesium or calcium. The three types of oxides are:

- acidic oxides
- basic oxides
- amphoteric oxides

(a) Acidic oxides

When non-metals are burnt in oxygen, the resultant oxide when dissolved in water forms a solution that turns blue litmus paper red. The solution is therefore **acidic**. Therefore, oxides of non-metals are acidic oxides. Examples are carbon dioxide (CO_2), sulphur dioxide (SO_2), phosphorous pentoxide (P_2O_5) among others. They form carbonic acid, sulphurous acid and phosphoric acids respectively.

Basic oxides

Basic oxides are formed when metals combine with oxygen. Such metals include magnesium, iron, sodium, zinc, calcium, among others. Some of these oxides dissolve in water to form alkalis. Others are insoluble in water. Examples of alkalis are sodium hydroxide, potassium hydroxide, magnesium hydroxide and calcium hydroxide.

Amphoteric oxides

Some oxides have both **basic** and **acidic** properties. They react with both bases and acids to form **salt** and **water** among other characteristics. Such oxides are said to be amphoteric oxides. Examples include aluminium oxide (Al_2O_3), zinc oxide (ZnO) and lead (II) oxide (PbO).

Experiment 2.3

Aim: *Reaction of metal oxides with acids and bases*

Apparatus and reagents

- Lead oxide
- Zinc oxide

Self assessment exercise 2.5

1. State three applications of acid-base neutralisation.

2.5 Classification of oxides

Research Activity

Find out from textbooks in the library or the internet the various categories of oxides. Write a report and present to the rest of the class.

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Experiment 2.3

Aim: *Reaction of metal oxides with acids and bases*

Apparatus and reagents

- Lead oxide
- Zinc oxide

- Aluminium oxide
- Dilute nitric acid
- Sodium hydroxide
- Boiling tubes
- Source of heat
- Spatula

Procedure

1. Scoop a small amount of lead (II) oxide (1g) using the spatula and place in a boiling tube followed by a few drops of nitric acid.
2. Warm the contents of the tube gently. Observe what happens. Draw a table similar to the one below in your note book and fill it as you carry out the experiment.

   Record your observations.
3. Repeat steps 1 and 2 using zinc oxide and aluminium oxide and record your observations.

Table 2.3: Reaction of oxides with acids and bases.

Oxide	Observation	
	2M nitric acid	2M sodium hydroxide
Lead (II) oxide		
Zinc oxide		
Aluminium oxide		

You may have noticed that these oxides react with both nitric acid and sodium hydroxide solutions. With nitric acid, they react to form a colourless solution of their respective salts i.e:

- Lead (II) nitrate
- Zinc nitrate
- Aluminium nitrate

With sodium hydroxide solution, they react to form colourless complex salt solution. These salt solutions are:

- Sodium plumbate
- Sodium zincate
- Sodium aluminate

The following table summarises the reactions.

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Table 2.4: Summary of reactions of amphoteric oxides

Metal oxide	Simple salt from nitric acid	Complex salt from sodium hydroxide
Zinc oxide (ZnO)	Zinc nitrate $Zn(NO_3)_2$	Sodium zincate $Na_2Zn(OH)_4$
Lead (II) oxide (PbO)	Lead nitrate $Pb(NO_3)_2$	Sodium plumbate $Na_2Pb(OH)_4$
Aluminium oxide (Al_2O_3)	Aluminium Nitrate $Al(NO_3)_3$	Sodium aluminate $NaAl(OH)_4$

Self assessment exercise 2.6

1. Name the three amphoteric oxides.
2. Mention the salt formed when excess sodium hydroxide solution is mixed with zinc nitrate solution.

2.6 Ways of preparing salts

A **salt** is the product formed when the hydrogen ion of an acid is replaced directly or indirectly by a metal (or ammonium radical). Names of salts are derived from the metal or ammonium radical and the acid radical from which they are formed.

When naming salts, the name starts with the name of the metal or ammonium radical in the salt followed by the respective acid radical. For example,

Name	Formula
Sodium chloride	$NaCl$
Ammonium chloride	NH_4Cl
Potassium sulphate	K_2SO_4
Zinc nitrate	$Zn(NO_3)_2$

From the name of the salt, you can tell the acid from which the salt was formed because:

- Hydrochloric acid always gives chlorides.
- Nitric acid always gives nitrates.
- Sulphuric acid always gives sulphates.
- Carbonic acid always gives carbonates.
- Ethanoic acid always gives ethanoates.
- Phosphoric acid always gives phosphates.

Salts that dissolve in water are said to be **soluble salts** while those that do not dissolve in water are known as **insoluble salts**. The method chosen for preparing a particular salt depends on the solubility of the salt in water. There are several methods of preparing salts. Soluble salts can be prepared by any of the following methods.

1. Action of an acid on a metal.
2. Action of an acid on an insoluble base (metal oxide or hydroxide).
3. Action of an acid on an insoluble carbonate.
4. Action of an acid on an alkali.
5. Direct combination of elements (direct synthesis).

We are now going to prepare soluble salts by doing the following experiments.

Preparation of soluble salts

1. Reaction of a dilute acid with a metal

Note: It is not safe to react a very reactive metal such as sodium or calcium with an acid.

Experiment 2.4

Aim: To prepare zinc sulphate

Apparatus and reagents

- Beaker (100 cm^3)
- Evaporating dish
- Copper (II) sulphate solution
- Zinc
- Filter funnel
- Conical flask
- Glass rod
- Filter papers
- Bunsen burner
- Dilute sulphuric acid

Procedure

1. Pour 50 cm^3 dilute sulphuric acid into a beaker.
2. Add zinc granules. What do you observe?
3. If the reaction is too slow, add a little copper (II) sulphate solution and warm gently, but do not boil. What do you observe? Can you give a reason why copper (II) sulphate was added?
4. When all zinc has reacted, add more and wait until it has all reacted. Add more again until some remain unreacted. Give a reason why we must keep on adding zinc until no more can react.
5. Filter off the excess zinc and solid impurities and collect the filtrate in a conical flask. What is the colour of the filtrate?
6. Evaporate the filtrate to dryness as illustrated in Fig 2.3 What do you observe in the evaporating dish?
7. Record your observations in your notebook

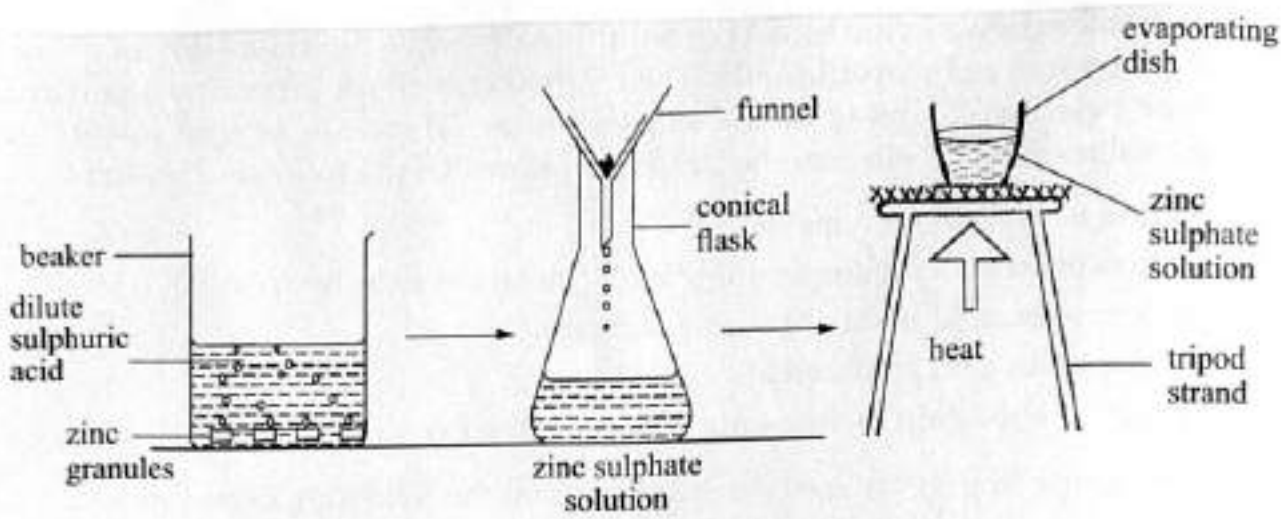


Fig. 2.3: Preparation of zinc sulphate salt

8. What do you conclude? Is it possible to prepare copper (II) sulphate crystals using copper metal and sulphuric acid? Explain.

Obtaining crystals

1. To obtain crystals of zinc sulphate, boil the filtrate to evaporate excess water.
2. Test the boiling filtrate from time to time to find out if the solution is saturated. You can do this by:

(a) Dipping a glass rod into the hot solution to remove a little of the liquid. Allow the solution to cool on the glass rod. If crystals form on the glass rod, stop heating and let the solution in the dish cool slowly. Cover the solution with a filter paper to keep off dust particles. Leave it until the next lesson.

(b) Pour a little of the hot solution into a test tube and cool with cold water. If small crystals form, then let it cool slowly until the next lesson.

3. When good crystals form, pour off the liquid. The liquid we decant is called "mother liquor".
4. Rinse the crystals with a little distilled water.
5. Dry the crystals between two filter papers or in sunshine.
6. Observe the crystals. Why are the crystals not dried by heating?

When the acid comes into contact with zinc, bubbles are observed. This shows that a gas is produced.

When copper sulphate solution is added to the reacting mixture it speeds up the reaction. When zinc is added hydrogen gas is evolved. If more zinc is added and some remain unreacted, it means all the acid has reacted and a colourless solution of zinc sulphate is formed.

When excess water is evaporated and we let the saturated solution to cool slowly, big crystals form. Crystals incorporate water as they form. This water is known as **water of crystallisation** and gives shape to the crystals. Crystals of zinc sulphate or $ZnSO_4 \cdot 7H_2O$ are white and needle-like. All hydrated salts contain water of crystallisation. But the number of water molecules differ in different crystals.

The reactions that occur are as follows:

(i) Formation of the salt.

Zinc + sulphuric acid ⟶ zinc sulphate + hydrogen

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

(ii) Formation of crystals from the salt

Zinc sulphate + water ⟶ zinc sulphate crystals

$ZnSO_4(aq) + 7H_2O(l) \longrightarrow ZnSO_4{\cdot}7H_2O(s)$

Iron sulphate and zinc nitrate are soluble salts which can also be prepared by this method. The equations for the reactions that will take place are as follows.

Iron + sulphuric acid ⟶ iron sulphate + hydrogen

$Fe(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + H_2(g)$

Iron sulphate + water ⟶ iron sulphate crystals

$FeSO_4(aq) + 7H_2O(l) \longrightarrow FeSO_4{\cdot}7H_2O(s)$

Zinc nitrate can also be prepared by a reaction of a dilute nitric acid and zinc metal.

Zinc + Nitric acid ⟶ zinc nitrate + water

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

Metals, such as sodium, potassium and calcium react vigorously with acids. They **should not** be used to prepare salts by the method described above. Other metals such as copper, silver and mercury do not react with acids and hence cannot be used to prepare salts by this method. Lead metal when reacted with dilute hydrochloric acid or dilute sulphuric acid form insoluble salts i.e. lead chloride or lead sulphate which coat the metal preventing further reaction. Therefore, lead metal is not suitable for preparing salts by this method.

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

Experiment 2.5

Aim: *To prepare copper (II) sulphate from copper oxide.*

Apparatus and reagents

- Conical flask
- Spatula
- Wire gauze
- Glass rod stirrer
- Filter funnel
- Evaporating dish
- Watch glass or tile.
- Beaker
- Tripod stand
- Bunsen burner
- Filter papers
- Lighter
- Dilute sulphuric acid
- Copper oxide

Procedure

1. Pour 50 cm^3 of dilute sulphuric acid into a beaker. Set up the apparatus as shown below.

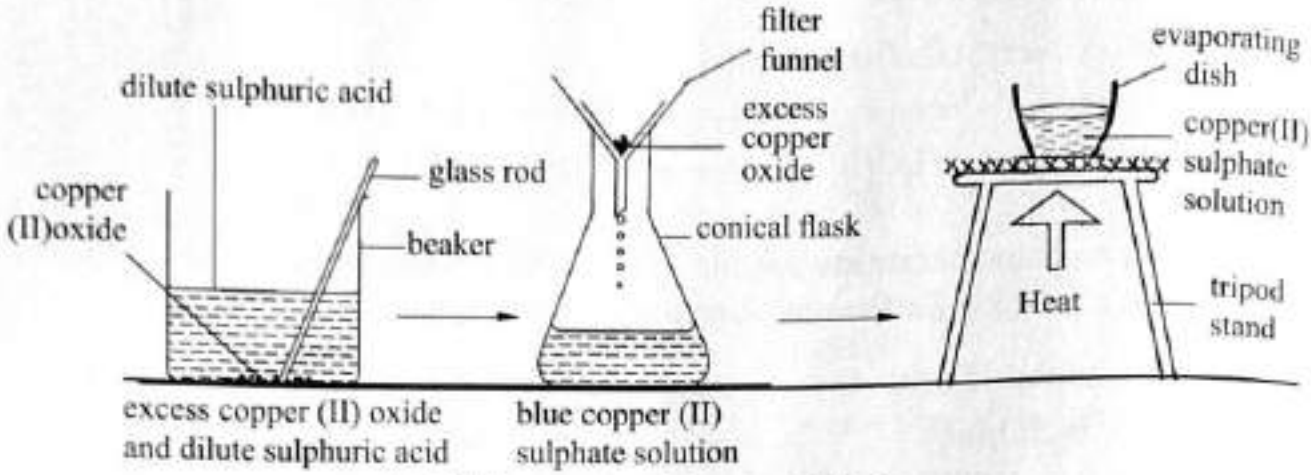


Fig 2.4 Preparation of copper sulphate

2. Place the beaker on the tripod stand and warm gently but do not boil.
3. Add copper oxide to the warm acid a little at a time. Stir and observe what happens. Test the solution with blue litmus paper.
4. Continue to add the oxide a little at a time stirring until no more oxide reacts with the acid. Give a reason why excess oxide is added. Test the solution with litmus paper.
5. Filter off the excess oxide and collect the filtrate in a conical flask.
6. Evaporate a little of the blue filtrate in an evaporating dish to dryness and note what happens.
7. Heat the remaining filtrate to evaporate excess water and test now and then for saturation point. At saturation point, allow the filtrate to cool slowly in the atmosphere.
8. When good crystals form, pour off the mother liquor.
9. Dry the crystals between two filter papers or under the sun.
10. Observe the crystals and note the difference with the salt obtained when the salt was evaporated to dryness in 6 above.
11. Put a few crystals in a test tube and heat strongly. Observe what happens. Place the residue on watch glass or tile and add a few drops of water. What do you observe?
12. What do you conclude from this experiment? Why is it not possible to prepare copper sulphate by reacting copper metal with dilute sulphuric acid? Explain.

When copper oxide is added to the warm acid, a blue solution is formed. Litmus paper turns red because there is a lot of acid. As more copper oxide reacts the intensity of the blue colour increases. Excess copper oxide is added to make sure that all the acid has reacted. The solution does not change the colour of litmus paper red at this point because there is no acid present in the solution.

When the filtrate is evaporated to dryness, a white powder remains. This is copper sulphate salt without water of crystallisation. It is known as **anhydrous copper sulphate**. But when excess water is evaporated and the filtrate cooled slowly, large blue salt crystals form. This is known as **hydrated copper sulphate.** Such salts have water of crystallisation and are known as **hydrated salts**. Hydrated copper sulphate has 5 water molecules.

When the hydrated copper sulphate is heated it loses water of crystallisation to form **white anhydrous copper sulphate** which is a white powder. When water is added to white anhydrous copper sulphate it turns blue.

The reactions that occurred are as follows:

(i) When the oxide reacted with acid

Copper oxide + sulphuric acid $\longrightarrow$ copper sulphate + water

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

(ii) During crystal formation

Copper sulphate + water $\longrightarrow$ hydrated copper sulphate

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

(iii) On heating the blue hydrated copper sulphate

Hydrated copper sulphate $\xrightarrow{\text{heat}}$ anhydrous copper sulphate + water

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

(iv) On adding water to anhydrous copper sulphate

Anhydrous copper sulphate + water $\longrightarrow$ hydrated copper sulphate

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

This is a reversible chemical reaction and is written as follows:

$$\underset{\text{(blue)}}{CuSO_4.5H_2O(s)} \rightleftharpoons \underset{\text{[(white)]}}{CuSO_4(s)} + 5H_2O(l)$$

The following salts can be prepared in the same way by adding an excess of the metal oxide or hydroxide to the appropriate acid.

(a) Magnesium sulphate from magnesium hydroxide

Magnesium hydroxide + sulphuric acid $\longrightarrow$ magnesium sulphate + water

$$Mg(OH)_2(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + 2H_2O(l)$$

(b) Lead nitrate from lead oxide

Lead oxide + nitric acid ⟶ lead nitrate + water

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

Examples of some other hydrated salts are:

Name	Formula
Magnesium sulphate -7-water	$MgSO_4.7H_2O$
Sodium carbonate -10-water	$Na_2CO_3.10H_2O$
Calcium sulphate - 2- water	$CaSO_4.2H_2O$
Cobalt chloride -6-water	$COCl_2.6H_2O$

3. Reaction of a dilute acid with metal carbonate or metal hydrogen carbonate

Experiment 2.6

Aim: *To prepare lead nitrate from lead carbonate.*

Apparatus and reagents

- Conical flask
- Beaker
- Spatula
- Tripod stand
- Wire gauze
- Bunsen burner
- Glass rod
- Filter paper
- Funnel
- Lighter
- Lead carbonate
- 2M dilute nitric acid

Procedure

1. Pour 50 cm^3 of dilute nitric acid into a beaker.
2. Add lead carbonate, a little at a time and stir. What do you observe?
3. Test the solution with blue litmus paper to find out whether all the acid has reacted. Filter off excess carbonate. Proceed to crystallise as in Experiment 2.5.

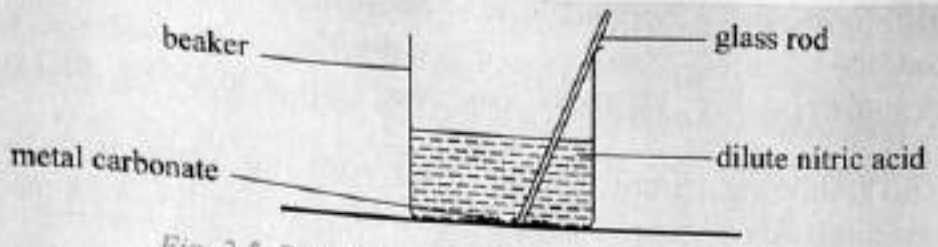


Fig. 2.5 Reacting a metal carbonate with an acid

4. Record your observations in your notebook.
   - Why is it not appropriate to use sulphuric acid or hydrochloric acid when preparing lead sulphate or chloride from lead carbonate?

When the metal carbonate is added to the acid, effervescence occurs showing that a gas is given off. The gas is carbon dioxide. All metal carbonates react with dilute acids to produce carbon dioxide and a solution of the metal salt so long as the salt formed is soluble i.e.

Carbonate + acid $\longrightarrow$ salt + carbon dioxide + water

Example:

Lead carbonate + nitric acid $\longrightarrow$ lead nitrate + carbon dioxide + water

$$PbCO_3(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

The excess carbonate is added to ensure that all the nitric acid has reacted.

Other soluble salts that can be prepared in a similar way from the carbonates or hydrogen carbonates include:

(a) Carbonates

(i) Copper sulphate, copper nitrate, magnesium sulphate and zinc sulphate.

Copper carbonate + sulphuric acid $\rightarrow$ copper sulphate + water + carbon dioxide

$$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$$

(ii) Copper carbonate + nitric acid $\rightarrow$ copper nitrate + water + carbon dioxide

$$CuCO_3(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2 + H_2O(l) + CO_2(g)$$

(iii) Zinc carbonate + sulphuric acid $\rightarrow$ zinc sulphate + water + carbon dioxide

$$ZnCO_3(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l) + CO_2(g)$$

(iv) Ammonium carbonate + sulphuric acid $\rightarrow$ ammonium sulphate + water + carbon dioxide

$$(NH_4)_2CO_3(s) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + H_2O(l) + CO_2(g)$$

(b) Hydrogen carbonates can also be used as shown:

(i) Potassium hydrogen carbonate + nitric acid $\rightarrow$ potassium nitrate + water + carbon dioxide

$$KHCO_3(s) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l) + CO_2(g)$$

(ii) Sodium hydrogen carbonate + hydrochloric acid → sodium chloride + water + carbon dioxide

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

However, if the following carbonates are used with the dilute acids as paired, the reaction will stop quickly before all the carbonate has reacted. Give a reason why?

- $CaCO_3(s)$ and $H_2SO_4(aq)$
- $PbCO_3(s)$ and $H_2SO_4(aq)$
- $PbCO_3(s)$ and $HCl(aq)$

The equations of the reactions are:

(i) Calcium carbonate + sulphuric acid → calcium sulphate + water + carbon oxide

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$$

(ii) Lead carbonate + sulphuric acid → lead sulphate + water + carbon dioxide

$$PbCO_3(s) + H_2SO_4(aq) \rightarrow PbSO_4(s) + H_2O(l) + CO_2(g)$$

(iii) Lead carbonate + hydrochloric acid → lead chloride + water + carbon dioxide

$$PbCO_3(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2O(l) + CO_2(g)$$

The salts formed are insoluble and will form a coating on the carbonate preventing further reaction. Calcium sulphate is only slightly soluble. Lead(II) chloride is soluble in hot water.

4. Reaction of a dilute acid with an alkali

Experiment 2.7

Aim: *To prepare sodium chloride from sodium hydroxide (a soluble base).*

Apparatus and reagents

- Burette
- Conical flask
- Dropper
- Evaporating dish
- Dilute sodium hydroxide
- Phenolphthalein indicator
- Pipette and a pipette filler
- Beaker
- Funnel
- Burner
- Dilute hydrochloric acid

Procedure

1. Using a pipette, transfer 25 cm^3 of the alkali into a conical flask. Then add 2-3 drops of phenolphthalein indicator as shown in Fig. 2.6a. The indicator turns pink. Why is it advisable to use a pipette filler to fill the pipette rather than the mouth?

2. Add the acid from a burette, a little at a time. Then swirl the flask in a controlled manner, to allow the acid and alkali to mix. See Fig. 2.6(b).
3. When all the alkali has been used up, the indicator suddenly turns **colourless**, showing that the solution is neutral. This is the end point of the reaction. There is no need of adding more acid. Look at Fig. 2.6(c).
4. You can tell how much acid was added, using the scale on the burette. Look at Fig. 2.6(d). This is the amount of acid that is needed to neutralise 25 cm^3 of the alkali.
5. Repeat steps 1 to 4, but this time there is no need for an indicator. 25 cm^3 of alkali is put in the flask and the correct amount of acid added. See Fig. 2.6(e).
6. The solution from the flask is then heated, to let the water evaporate. See Fig. 2.6(f). You will find that dry crystals of sodium chloride are left behind.
7. Record your observations in your note book. What do you conclude?

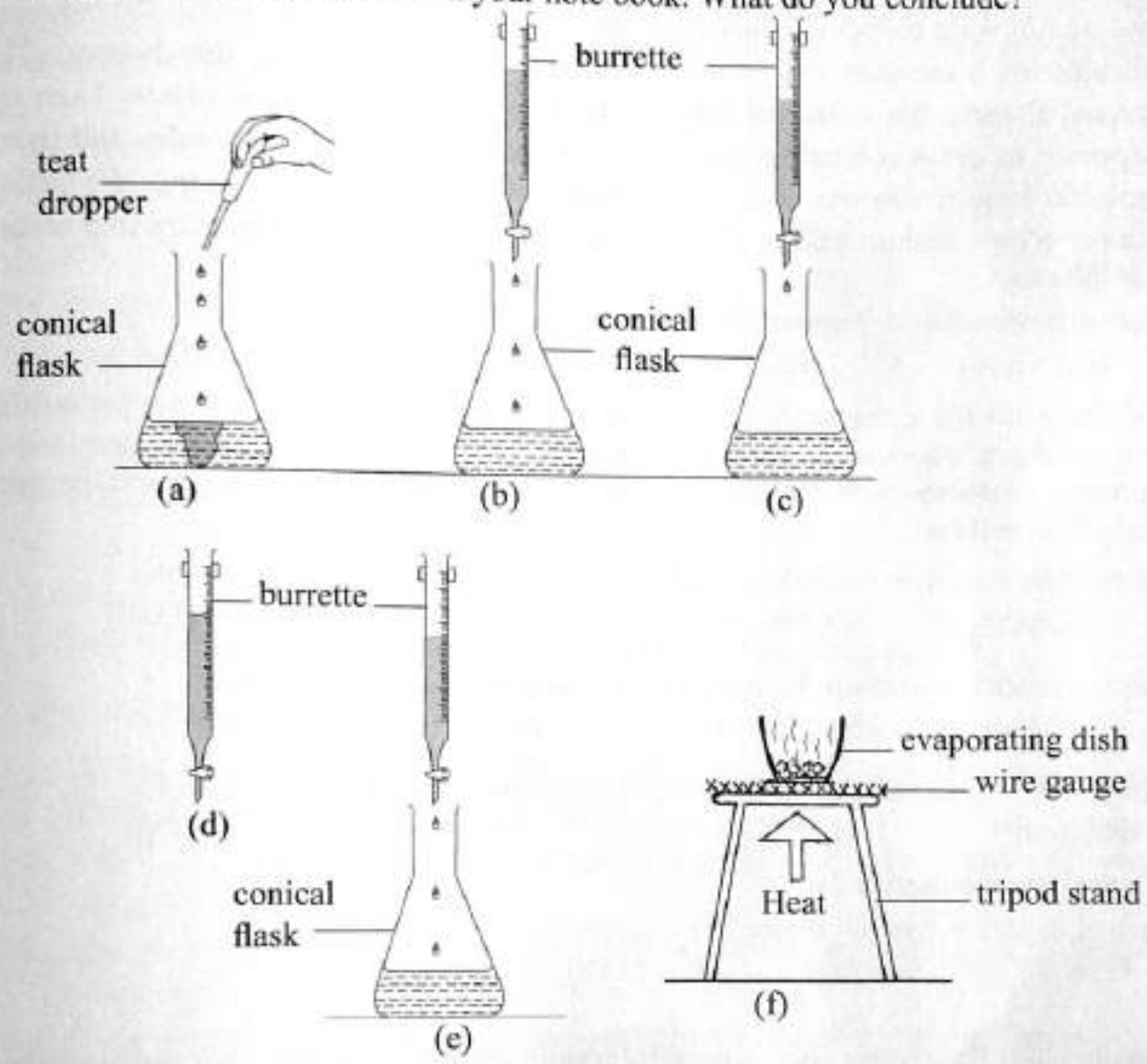


Fig 2.6 Preparation of sodium chloride

Study questions

1. What do you think would happen if we evaporated the solution obtained in step 3 to dryness?
2. Why was the whole process repeated i.e. steps 1 to 4?
3. Using the same procedure, describe how you can prepare potassium sulphate. Choose the appropriate reagents.

When we want to fill a pipette, it is advisable to use a pipette filler to avoid sucking the liquid into the mouth accidentally. It is necessary to use an indicator because both the reactants and products are colourless and it would be difficult to know the **neutralisation point or end point of the reaction** of the reaction. If we evaporate the solution with the indicator, our salt would be coloured.

If we do not want to run the burette twice, we can boil the solution with the coloured indicator for 5 minutes with animal charcoal and then filter off the charcoal. The charcoal absorbs the colour of the indicator and the filtrate is colourless. Then we evaporate to get a colourless salt. This method of preparing a soluble salt using a soluble base is known as **neutralisation**. **Neutralisation** is a chemical reaction between a base and an acid to produce **salt** and **water** only. The reaction that occurs is as follows:

Sodium hydroxide + hydrochloric acid → sodium chloride + water

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

The three alkalis commonly used in this way are sodium hydroxide, potassium hydroxide and aqueous ammonia (ammonium hydroxide). The following equations represent reactions between the mineral acids and the alkalis to form soluble salts using this method.

Hydrochloric acid + potassium hydroxide → potassium chloride + water

$$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$$

Sulphuric acid + sodium hydroxide → sodium sulphate + water

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

Nitric acid + ammonium hydroxide → ammonium nitrate + water

$$HNO_3(aq) + NH_4OH(aq) \rightarrow NH_4NO_3(aq) + 2H_2O(l)$$

General ionic equation is:

Hydrogen ion + hydroxide ion → water

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

Note (i) Potassium and sodium hydroxide exist in solid form but ammonium hydroxide exists in solution form. For example,

(ii) Using the mineral acids we can also form sodium, potassium and ammonium salts from their carbonates.

Sodium carbonate + hydrochloric acid → sodium chloride + water + carbon dioxide

$$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Ammonium carbonate + nitric acid → sodium chloride + water + carbon dioxide

$$(NH_4)_2CO_3(aq) + 2HNO_3(aq) \rightarrow 2NH_4NO_3(aq) + H_2O(l) + CO_2(g)$$

General Ionic equation

Carbonate ion + hydrogen ion → water + carbon dioxide

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

Preparation of insoluble salts by precipitation

The methods we have discussed so far are used to prepare soluble salts only. However, some salts are insoluble in water; for example lead sulphate, silver chloride and barium sulphate among others. Such salts are prepared by a method known as **precipitation or double decomposition.**

Experiment 2.8

Aim: *To prepare lead sulphate*

Apparatus and reagents

- Test tube
- Teat pipette
- Funnel
- Conical flask
- Filter papers
- Lead nitrate solution
- Bunsen burner
- Sulphuric acid (or any soluble sulphate e.g. sodium sulphate)

Procedure

1. Fill one test tube to about a third with lead nitrate solution and the other with dilute sulphuric acid (or a soluble sulphate).
2. Add dilute sulphuric acid or the soluble sulphate to lead nitrate. Observe what happens?
3. Remove the colourless solution by filtration into a conical flask or allow the salt to settle and decant the solution.
4. Wash the salt with hot water from a teat pipette to remove impurities.
5. Allow the precipitate to dry on the filter paper to become a white powder.
6. Record observations in your notebook. What do you conclude?

7. Suggest another substance which you can react with lead nitrate solution to get lead sulphate in place of sulphuric acid. Write an equation for the reaction.

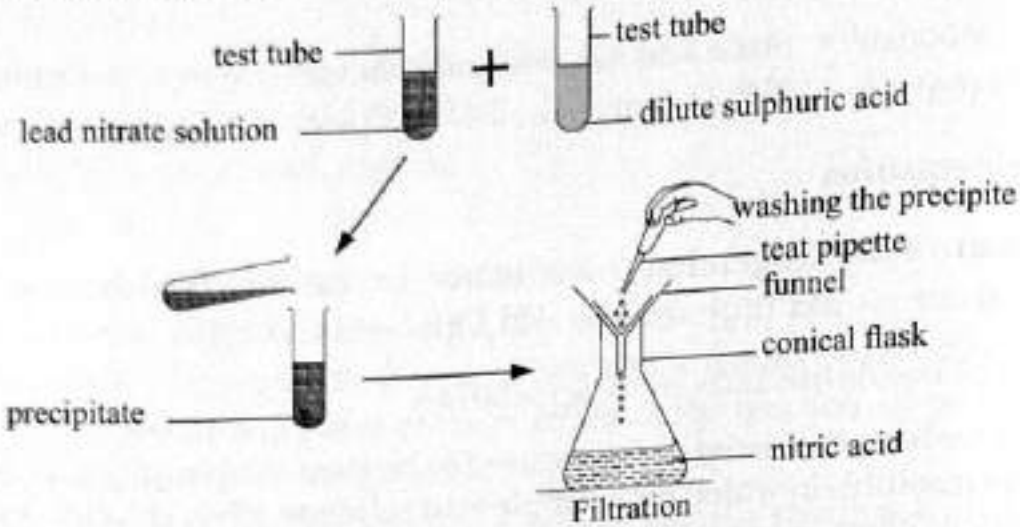


Fig. 2.7: Preparation of lead sulphate

When dilute sulphuric acid or any other soluble sulphate is added to the lead nitrate solution, a white insoluble salt is formed. This insoluble product is called **precipitate**. This method of preparing salts is called **precipitation** or **double decomposition**. In this method, we mix solutions of two substances. When we mix the two solutions, they exchange ions as shown in the following equation.

Lead nitrate (solution) + hydrogen sulphate (solution) → lead sulphate (precipitate) + hydrogen nitrate (remains in solution)

$$Pb(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow PbSO_4(s) + 2HNO_3(aq)$$

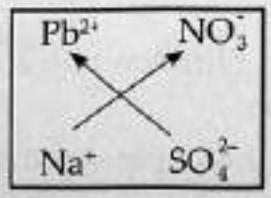
Note: Sulphuric acid is called hydrogen sulphate and nitric acid hydrogen nitrate in this case. This is meant to show how hydrogen ions are exchanged. A solution of any sulphate could have been used instead of sulphuric acid for this preparation; for example sodium sulphate.

Lead nitrate + sodium sulphate → lead sulphate + sodium nitrate

$$Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$$

The ions present in the solutions are

- From $Pb(NO_3)_2(aq)$ →
- From $Na_2SO_4(aq)$ →



ions exchanged to give products $PbSO_4$ and $NaNO_3$

Ionic equation

When we write an ionic equation, we select the ions that form what we see; in this case the precipitate. We leave out the ions that remain in solution. These are known as **spectator ions**. Spectator ions do not undergo change during the reaction.

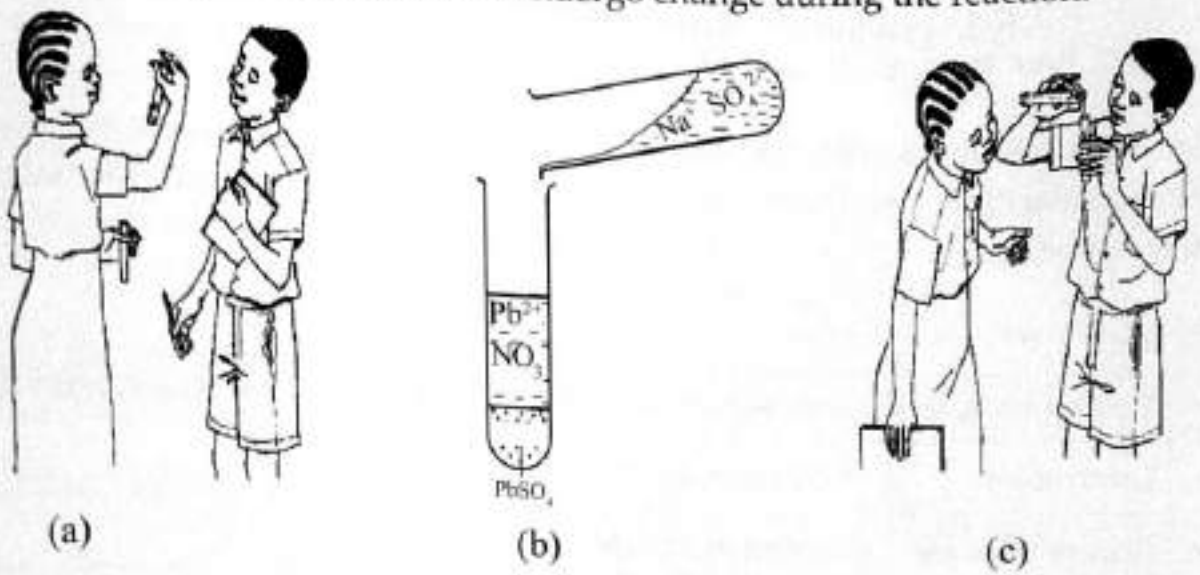


Fig. 2.8 Mixing two solutions

For instance, in the reaction above, Fig 2.8, the precipitate is lead sulphate. The ionic equation for this reaction would be written as:

Lead(II) ions + sulphate ions → lead sulphate

$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

Other insoluble salts which can be prepared by double decomposition are: barium sulphate and lead chloride.

Note: Lead chloride is a precipitate that is soluble in hot water.

Barium chloride + sodium sulphate → barium sulphate + sodium chloride

$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

Ionic equation:

Barium ion + sulphate ion → barium sulphate

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Lead nitrate + sodium chloride → lead chloride + sodium nitrate

$Pb(NO_3)_2(aq) + 2NaCl(aq \rightarrow PbCl_2(s) + 2NaNO_3(aq)$

Ionic equation:

Lead ion + chloride ion → lead chloride

$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

soluble in hot water

Note: In neutralisation reactions when an acid is mixed with an alkali the solution resembles water. Therefore, when we write an ionic equation we choose the ions that form water as follows.

$$\text{Hydrogen ion} + \text{hydroxide ion} \rightarrow \text{water}$$
$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Repeat Experiment 2.8 with the following pairs of aqueous solutions in Table 2.5. Note the colour of the precipitates and write balanced equations in each case followed by ionic equations.

Table 2.5: Formation of some coloured precipitates

	Solution of salt X	Solution of substance Y	Record colour of precipitate if any
(a)	Lead nitrate	Potassium iodide	
(b)	Copper sulphate	Sodium hydroxide	
(c)	Iron sulphate	Sodium hydroxide	
(d)	Iron chloride	Sodium hydroxide	
(e)	Sodium chloride	Silver nitrate	
(f)	Lead nitrate	Ammonium chloride	
(g)	Ammonia solution	Zinc nitrate	

Note: Precipitation reactions are used in analysis of salts to identify the cations and anions in the salt. We will do this later in the course.

Writing ionic equations

Ionic compounds are made of a lattice of ions that are rigidly fixed in solid form. When ionic compounds are dissolved in water, the ions are set free to move and react. When a chemical reaction takes place between dissolved ionic compounds, it may happen that only certain ions undergo changes while others do not. Reactions that take place between ionic compounds dissolved in water can be recorded in a much simpler manner.

Solubility Rules.

The following are some of the rules to remember when choosing the method to use when preparing salts.

- All potassium (K), sodium (Na) and ammonium (NH_4) salts 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 carbonate salts are insoluble except potassium, sodium and ammonium carbonates.

Example

When sodium carbonate solution is mixed with barium chloride solution, a reaction takes place and a white precipitate is formed. Fig 2.8.

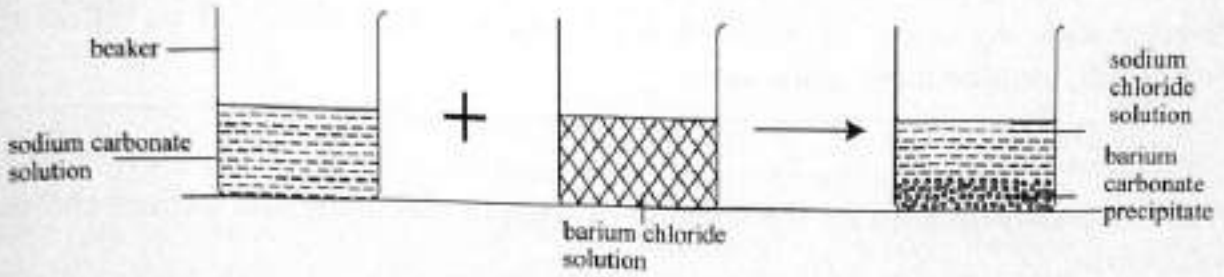


Fig. 2.9: Precipitation reaction

The balanced chemical equation for this reaction is as follows:

Sodium carbonate + barium chloride → sodium chloride + barium carbonate

$$Na_2CO_3(aq) + BaCl_2(aq) \rightarrow 2\,NaCl(aq) + BaCO_3(s)$$

Before the solutions were mixed, the ions present were:

- From sodium carbonate solution ($Na_2CO_3(aq)$) – $2Na^+$ and CO_3^{2-}
- From barium chloride solution ($BaCl_2(aq)$) – Ba^{2+} and $2Cl^-$

When the solutions are mixed, Ba^{2+} ions unite with the CO_3^{2-} ions to form $BaCO_3$ which is an insoluble white precipitate. Na^+ and Cl^- ions undergo no changes.

An ionic equation shows only those ions that change in some way during a reaction. From the above example, the ions that do not undergo any change are omitted from the ionic equation. Thus the ionic equation for the above is:

There are three possibilities during ionic changes:

$$Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$$

Na^+ and Cl^- ions are known as spectator ions. They remain unchanged at the end of the reaction.

Table 2.6 shows some common substances that contain free ions and those that do not for your guidance.

Table 2.6 Substances containing free and no free ions

Free ions	No free ions
Metal salts in solution	Insoluble solids including precipitates
Ammonium salts in solution	Covalent compounds (HCl, NH_3, N_2, O_2)
Gases in solution ($HCl(aq)$)	Uncombined elements (O, N, P, S)
Dilute acids (HCl, HNO_3, H_2SO_4)	Gases (O_2, N_2, CO_2, SO_2)
Soluble metal alkalis ($NaOH$, KOH,)	

Knowledge of solubility of salts in water is very crucial when writing ionic equations. This assists very much in identifying correct symbols.

Further an ionic equation can be written following certain rules. Let us follow these rules to write various ionic equations.

Example 1

Write an ionic equation for the reaction of sodium carbonate and barium chloride solutions.

Step 1

Write down well balanced stoichiometric or the normal chemical equation for the reaction.

Sodium carbonate + barium chloride → sodium chloride + barium carbonate

$$Na_2CO_3(aq) + BaCl_2(aq) \rightarrow 2NaCl(aq) + BaCO_3(s)$$

Step 2

Decide which substances are ionic and whose ions are free in solution.

Omit the spectator ions by crossing the ion that appears on both sides of the equation.

Note: Remember that spectator ions are those ions that remain unchanged during the chemical reaction. They appear in the same state on both sides of the equation.

$$Ba^{2+}(aq) + \cancel{2Cl^-}(aq) + \cancel{2Na^+}(aq) + CO_3(aq) \rightarrow \cancel{2Na^+}(aq) + \cancel{2Cl^-} + BaCO_3(s)$$

Step 3

Re arrange the ions and rewrite the equation in ionic form.

$$Ba^{2+}(aq) + CO^{2-}_3(aq) \rightarrow BaCO_3(s)$$

Example 2

Write the ionic equation for the reaction of iron chloride with sodium hydroxide.

Step 1

Write down balanced stoichiometric or the normal chemical equation for the reaction.

Iron chloride + sodium hydroxide → sodium chloride + iron hydroxide

$FeCl_3(aq) + 3NaOH(aq) \rightarrow 3NaCl(aq) + Fe(OH)_3(s)$

Step 2

Decide which substances are ionic and whose ions are free in solution.

Omit the spectator ions by crossing the ions that appear on both sides of the equation.

$Fe^{3+}(aq) + 3Cl^-(aq) + 3Na^+(aq) + 3OH^-(aq) \rightarrow 3Na^+(aq) + 3Cl^-(aq) + Fe(OH)_3(s)$

Step 3

Re-arrange the ions and rewrite the equation in ionic form.

$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

Preparation of salts by direct combination of elements

Some salts can be made by directly reacting together the two elements that make them. This process is called **direct combination** or **direct synthesis.** The two elements combine to produce one substance. Usually, energy in form of heat is required to make the two elements combine chemically.

Experiment 2.9

Aim: *Preparation of sodium chloride by direct combination of sodium and chlorine*

Apparatus and reagents

- Deflagrating spoon
- Gas jar full of dry chlorine
- Knife/scalpel
- Tongs
- Bunsen burner
- Sodium metal

Procedure

1. Remove sodium from the storage container using a pair of tongs.
2. Cut a small piece of sodium metal and place it in a deflagrating spoon.
3. Heat the piece of metal over the Bunsen flame.
4. When the metal catches fire, lower the burning sodium into a gas jar containing dry chlorine gas.
5. Record your observations.

In the experiment 2.9 above a white powder of sodium chloride is formed i.e.

Sodium + chlorine → sodium chloride

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

Some other metals also combine with chlorine to form anhydrous salts. For example, when chlorine is passed over heated iron wire, iron combines with chlorine to form iron chloride.

Iron + chlorine → iron chloride

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

Also, Iron can combine with sulphur to your iron sulphide

Iron + sulphur $\xrightarrow{\text{heat}}$ Iron sulphide

$$Fe + S \xrightarrow{\text{heat}} FeS$$

Self assessment exercise 2.7

1. Define the term salt.
2. Given these methods of preparing salts:
   (a) Direct combination
   (b) Precipation
   (c) Neutralisation

   Select a suitable method for the preparation of the following salts:
   (i) sodium chloride
   (ii) silver chloride
   (iii) iron chloride

2.7 Application of precipitation reactions

Research Activity

Find out how precipitation reactions are useful in our society. Write a report and share with other class members.

Precipitation reactions involve two different salt solutions in which two different metal ions exchange acid radicals forming an insoluble salt and a soluble one.

Such reactions are useful in the removal of some toxic substances e.g. lead and mercury from water for domestic use and in treating industrial effluents before releasing them into water bodies.

(a) Treatment of domestic water

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

$$Pb^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow PbCO_3(s)$$

$$Hg^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow HgCO_3(s)$$

Also, sodium carbonate (Na_2CO_3) removes water hardness i.e. it softens water by precipitating ions that cause water hardness i.e. calcium (Ca^{2+}) ions and Magnesium ions (Mg^{2+}) ions.

(b) Treatment of industrial effluents

In treating industrial effluents, precipitation process is carried out to remove toxic ions in the water solution.

It involves reacting the ions (to be removed) with other ions to produce insoluble solids that can be removed by sedimentation or filtration.

Note: In the treatment of domestic water and effluents, the liquid water is first neutralised.

Remember!

The solid sludge formed from most methods of industrial effluent treatment is often toxic and hence requires proper disposal.

Revision exercise 2

1. Differentiate between an acid and a base.
2. Given the following equations; identify two acid-base conjugate pairs

   (a) $H_3O^+(aq) + NH_3(aq) \rightleftharpoons H_2O(l) + NH_4^+(aq)$

   (b) $HAc(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Ac^-(aq)$
3. Name any two examples of amphoteric oxides.
4. (a) The pH of a sample of soil was found to be 5.1. A soil chemist recommended the addition of calcium oxide in the soil. State two functions of the oxide in the soil.

   (b) When a student was stung by a nettle plant, a laboratory assistant applied ammonia solution to the affected area of the skin and the student was relieved of pain. Explain.
5. Define the term salt
6. Briefly explain how you would prepare zinc sulphate crystals.

Topic 3 Reduction and Oxidation reactions

Sucess criteria

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

(a) Define the terms oxidation and reduction.
(b) Assign an oxidation number to an atom in an element, ion or compound.
(c) Identify reducing and oxidising agents.
(d) Write half and overall redox equations.
(e) Deduce the order of reactivity of metals.
(f) Predict spontaneity of a reaction.
(g) Describe corrosion.
(h) State conditions necessary for corrosion and methods of prevention.

Introduction

You are already familiar with various types of reactions which involve use of oxygen. Examples include burning of substances in air (oxygen) and respiration (reaction of substances with oxygen) in the body cells, all of which are different forms of combustion. These chemical reactions can be defined using the terms **oxidation** and **reduction**. The other common types of reactions such as precipitation, decomposition, displacement reactions and reactions of metals with acids among many others can equally be described in terms of oxidation and reduction as we shall see shortly.

3.1 Reduction and oxidation

Oxidation can be defined in three broad ways:

- Addition of oxygen to a substance.
- Removal of hydrogen from a substance.
- Loss of electrons from a substance.

Reduction on the other hand can be described as:

- Removal of oxygen from a substance.
- Addition of hydrogen to a substance.
- Gain of electrons by a substance.

When one substance is oxidised in a reaction, another substance must be reduced. Therefore, oxidation and reduction occur together and are called **redox reactions**.

The term **redox** is an abbreviation for **reduction** and **oxidation**. The first three letters i.e. 'red' are borrowed from the word reduction and the next two i.e. 'ox' from the word oxidation. All redox reactions therefore involve oxidation and reduction processes.

1. **(a) Oxidation in terms of addition of oxygen to a substance**

During burning of elements such as magnesium and hydrogen in air, oxygen is added to the elements. When a substance gains oxygen in a reaction it is said to be oxidised. The process is known as **oxidation**. For example;

(i) Magnesium + oxygen $\longrightarrow$ Magnesium oxide

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

(ii) Hydrogen + Oxygen $\longrightarrow$ water

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

(b) Reduction by removal of oxygen from a substance

We have reactions where some substances remove oxygen from compounds. For example, where a stream of hydrogen or carbon monoxide gas is passed over some heated metal oxides, the oxygen atom in the oxide is removed. Removal of oxygen from compounds is known as **reduction**. The substance which is removing the oxygen is the **reducing agent**. The metal oxide adding oxygen atom to the hydrogen as in equation (i) and to carbon monoxide as in equation (ii) is called the **oxidising agent**. For example;

(i) Copper oxide + hydrogen $\xrightarrow{\text{heat}}$ copper + water

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

(ii) Copper oxide + carbon monoxide $\xrightarrow{\text{heat}}$ copper + carbon dioxide

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

(iii) Lead oxide + carbon $\xrightarrow{\text{heat}}$ lead + carbon monoxide

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

In the above reactions, oxygen is being removed from one substance and added to another substance. Hence both reduction and oxidation is taking place simultaneously.

2. **(a) Oxidation by removal of hydrogen from a substance to another substance**

Also, oxidation reaction takes place when a substance loses hydrogen. For example:

Hydrogen sulphide + chlorine $\longrightarrow$ sulphur + hydrogen chloride

$H_2S(g) + Cl_2(g) \longrightarrow S(s) + 2HCl(g)$

In the above reaction, there is no oxygen in any of the products or reactants. The sulphide loses its hydrogen to chlorine and the reaction is said to be an **oxidation reaction** irrespective of absence of oxygen.

This is because chlorine removes hydrogen from hydrogen sulphide, (H_2S) gas. Chlorine is hence the **oxidising agent** while hydrogen sulphide is the **reducing agent**. This is because it gives its hydrogen to chlorine.

(b) Reduction by addition of hydrogen to a substance

When one substance undergoes oxidation due to loss of its hydrogen, the substance that gains the hydrogen is said to have undergone **reduction.**

For example, in the equation in 2(a), chlorine undergoes reduction. This is because it gains hydrogen. Hydrogen sulphide in this case undergoes **oxidation** – due to loss of hydrogen.

Another example of reduction reaction due to addition of hydrogen is:

Chlorine + Hydrogen $\longrightarrow$ hydrogen chloride

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

3. **(a) Oxidation in terms of loss of electrons and increase in oxidation number**

A substance that loses electrons in a reaction is said to have undergone oxidation, and its oxidation number also increases. For example, when metal atoms react with non-metal atoms, electrons are transferred from the metal atoms to the non-metal atoms, hence oxidation number of the metal atoms also increase, for example,

Magnesium + chlorine $\longrightarrow$ magnesium chloride

$$2Mg(s) + Cl_2(g) \longrightarrow 2MgCl(s)$$

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

The oxidation number of magnesium has increased from zero (0) to positive two (+2)

(b) Reduction in terms of gain of electrons and decrease in oxidation number

In the equation in 3(a), chlorine atom undergoes reduction and in the process it gains an electron. This causes its oxidation number to decrease i.e. moves from zero (0) to negative one (–1).

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$$

From examples 3(a) and 3(b), magnesium is said to be the **reducing agent** while chlorine is the **oxidising agent.** Therefore an oxidizing agent is the reactant which gains electrons and whose oxidation number decreases, while a reducing agent is the reactant that loses electrons and whose oxidation number increases.

3.2 Determination of oxidation numbers

Oxidation number, also called **oxidation state** is determined by the electrons gained, lost or shared by the element in a compound during bonding with other elements present. The charge exhibited by an atom of an element determines their oxidation number. See Table 3.1 on page 69. This implies that a neutral atom has an oxidation number of zero.

Table 3.1: Oxidation number of some common ions

Ion	Oxidation number	Ion	oxidation number
Na^+	+1	K^+	+1
Mg^{2+}	+2	Ca^{2+}	+2
Al^{3+}	+3	Li^+	+1
Cl^-	–1	S^{2-}	–2
O^{2-}	–2	N^{3-}	–3
P^{3-}	–3	H^{+1}	+1

To find the oxidation number of an element in a molecule or a complex ion, the following rules are applied:

1. The oxidation number for an atom of any free (uncombined) element is **zero**, thus the atoms; K, Na and Fe have zero (0) oxidation number. This rule also applies to neutral molecules.
2. The oxidation number for any simple one-atom ion is equal to its charge, thus the oxidation number of $K^+ = +1$; $Ca^{2+} = +2$ and of $Cl^- = -1$.
3. The oxidation number of hydrogen in non-ionic compound is +1. This applies to hydrogen compounds such as H_2O, HCl and CH_4 among others. For the ionic metal hydrides such as sodium hydride, NaH, the oxidation number of hydrogen is -1.
4. The oxidation number of oxygen is –2 in all compounds for example H_2O, NO, CO_2 except in peroxides, for example in H_2O_2 where it is –1. Another exception is OF_2 in which oxygen has +2 as its oxidation number.
5. The algebraic sum of the oxidation numbers of all atoms in the formula of a neutral compound must be zero. For example, in H_2O since the oxidation number of H = +1 and O = –2, it follows that the oxidation state of water (H_2O) is zero (0) given by $(+1 \times 2) + (-2) = 0$.
6. The algebraic sum of the oxidation numbers of all the atoms in a complex ion is equal to the charge on the ion. Thus in NH_4^+ ion, if we add the oxidation number of N and those of the 4H atoms, the answer must be equal to +1. In SO_4^{2-}, the sum of all oxidation numbers must be equal to –2 and so on.

Example 1

What is the oxidation number of Na, Mg, O_2 and N_2?

Solution

Zero (0) because Na and Mg are uncombined elements while O_2 and N_2 are neutral molecules.

Example 2

What is the oxidation number of chlorine in

(a) Cl_2?

(b) HCl?

(c) HOCl?

(d) ClO_3^-?

Solution

(a) Zero (0)

- In (b) to (d), we will assign oxidation number of chlorine a value x, and write down an algebraic equation and solve for x.
- In (b) and (c), the algebraic sum of oxidation numbers must be equal to zero.
- From the rules above, the oxidation number of H = +1 and O = –2.

Thus

(b) HCl = 0

$+1 + x = 0$

$x = 0 - 1$

$= -1$

The oxidation number of chlorine in HCl is –1

(c) The oxidation number of H = +1 , O = –2

HOCl = 0

$+1 + (-2) + x = 0$

$x = +1$

The oxidation number of chlorine in HOCl is +1

(d) Rule number 6 – sum of oxidation number of all atoms in a complex ion is equal to the charge on the ion i.e. –1 hence.

$x + 3(-2) = -1$

$x + (-6) = -1$

$x = -1 + 6$

$= +5$

The oxidation number of Cl atom in ClO_3^- is +5.

Self assessment exercise 3.1

1. What are the oxidation states of nitrogen in the following substances:
   (a) N_2O (b) NO_2 (c) NO_3^-
2. Find the oxidation state of manganese in the following compounds:
   (a) $KMnO_4$ (b) MnO_4^- (c) MnO_2
3. (a) Define oxidation in terms of:
   (i) Oxygen (ii) Hydrogen
   (b) Define reduction in terms of:
   (i) Hydrogen (ii) Oxygen
4. (a) What do you understand by the term oxidation number?
   (b) What is the oxidation number of?
   (i) I_2 (ii) N_2 (iii) O_2 (iv) Fe
   Explain your answer.

3.3 Oxidation numbers and redox reactions

When an atom or an ion is oxidised, its oxidation number also increases while when it is reduced, its oxidation number decreases.

Consider the following equations:

1. Reaction of iron (II) ions, Fe^{2+} with acidified hydrogen peroxide

When iron (II) ions react with acidified hydrogen peroxide, iron (II) ions are oxidised to iron(III) ions as shown in the following half equations.

(i) $2Fe^{2+}(aq) \longrightarrow 2Fe^{3+}(aq) + 2e^-$

The oxidation number for iron has increased by +1 from +2 in Fe^{2+} ions to +3 in Fe^{3+} ions i.e. Fe^{2+} ion is oxidised to Fe^{3+} ion. Reaction (i) is therefore an **oxidation reaction.**

(ii) $2H^+(aq) + H_2O_2(aq) + 2e^- \longrightarrow 2H_2O(l)$
(from acid)

The oxidation state of oxygen has decreased from –1 in hydrogen peroxide (H_2O_2) to –2 in water (H_2O) i.e. oxygen is reduced. Equation (ii) is therefore a **reduction reaction**.

When the two half equations are combined, an overall redox equation is obtained as shown below.

$$2Fe^{2+}(aq) \longrightarrow 2Fe^{3+}(aq) + \cancel{2e^-}$$

$$2H^+(aq) + H_2O_2(aq) + \cancel{2e^-} \longrightarrow 2H_2O(l)$$

$$2Fe^{2+}(aq) + 2H^+ + H_2O_2(aq) \longrightarrow 2Fe^{3+}(aq) + 2H_2O(l)$$

2. *Reaction of Iron (II) ions (Fe^{2+}) with acidified potassium permanganate*

When Iron sulphate solution is reacted with acidified potassium permanganate the green colour of the solution turns brown. This is because the Fe^{2+} ions (green) are oxidised to Fe^{3+} ions (brown). This is another example of a redox reaction.

reduction

$$5Fe^{2+}(aq) + 8H^{+}(aq) + MnO_4^{-} \longrightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$$

oxidation

The oxidation number of Iron (II) ions, Fe^{2+} is +2. The oxidation number of Iron (III) ion, Fe^{3+} is +3. Because there is an increase of oxidation number, it is an **oxidation reaction**. However, the oxidation number of Manganese (Mn) decrease from +7 to +2. This is a **reduction reaction.**

3. *Reaction of hydrogen sulphide with sulphur dioxide gas*

Sulphur in hydrogen sulphide is moved from an oxidation state of -2 to zero. On the other hand sulphur in sulphur dioxide is reduced from oxidation number of +4 to zero. Because there is an increase in oxidation number, this is an **oxidation reaction.** Therefore, this is a reduction reaction.

oxidation

$$2H_2S(g) + SO_2(g) \longrightarrow 3S(s) + 2H_2O(l)$$

reduction

Oxidising and reducing agents

Go back to the reactions 1-3 above. For each reaction, can you say:

(a) Which species is oxidised

(b) Which species is reduced and by what?

For example:

In Reaction 1 – Iron (II) ions Fe^{2+} are oxidised by hydrogen perioxide (H_2O_2) while hydrogen perioxide is reduced by Fe^{2+}.

In Reaction 2 – Iron (II) ions Fe^{2+} are oxidised by manganese (MnO_4^-) ions are reduced by Fe^{2+} ions.

In Reaction 3 – Hydrogen sulphide (H_2S) is oxidised by sulphur dioxide (SO_2) while sulphur dioxide is reduced by hydrogen sulphide.

In the three cases, the species that causes oxidation is the **reducing agent** while the species that causes reduction is the **oxidising agent**. Therefore, can you identify oxidising and reducing agents in the reactions above.

Self assessment exercise 3.2

1. Distinguish between oxidising and reducing agent.
2. Identify the oxidised and reduced substances in each of the following reactions.

   (a) $Cr_2O_7^{2-}(aq) + 6I^-(aq) + 14H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 3I_2(s) + 7H_2O(l)$

   (b) $5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \longrightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$
3. State the species being oxidised, giving a reason for your answer.

   (a) $2I^-(aq) + Br_2(g) \longrightarrow I_2(s) + 2Br^-(aq)$

   (b) $2Fe^{3+}(aq) + H_2S(g) \longrightarrow 2Fe^{2+}(aq) + 2H^+(aq) + S(s)$

3.4 Writing and balancing redox equations

When balancing redox equations, the equation is broken into two parts:

- the oxidation half equation.
- the reduction half equation.

For a redox equation to be said to be balanced, the charges on both sides of the equations must be equal i.e. the number of electrons lost by one chemical species must be equal to the number of electrons gained by another chemical species in a reaction.

Rules for balancing redox equations

1. Identify the oxidised and the reduced species then write the reduction and oxidation half equations.
2. Balance the half-equations separately.
3. Ensure the number of electrons lost in oxidation half equation are equal to those gained in reduction half equations.
4. Combine the half equations and cancel any identical species that appear on both sides of the equation.

Example 1

Balance the following equation:

$Al(s) + Ag^+(aq) \longrightarrow Al^{3+}(aq) + Ag(s)$

Aluminium is oxidised and silver is reduced

Oxidation half equation

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

Reduction half equation

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

Multiply the reduction half equation by three to get

$3Ag^+(aq) + 3e^- \longrightarrow 3Ag(s)$

Add both equations then cancel any identical species that appear on both sides of the equation to get a balanced redox equation.

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

$3Ag^+(aq) + 3e^- \longrightarrow 3Ag(s)$

$Al(s) + 3Ag^+(aq) + \cancel{3e^-} \longrightarrow Al^{3+}(aq) + \cancel{3e^-} + 3Ag(s)$

$Al(s) + 3Ag^+(aq) \longrightarrow Al^{3+}(aq) + 3Ag(s)$

Example 2

Balance the equation below:

$Zn(s) + Fe^{3+}(aq) \longrightarrow Zn^{2+}(aq) + Fe(s)$

Zinc is oxidised and iron is reduced.

Oxidation half equation

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

Reduction half equation

$F^{3+}(aq) + 3e^- \longrightarrow Fe(s)$

Multiply the oxidation half equation by 3 and the reduction half equation by 2 to get six electrons in each equation. Combine the two half equations then simplify to get a balanced redox equation. Thus:

$3Zn(s) \longrightarrow 3Zn^{2+}(aq) + 6e^-$

$2Fe^{3+}(aq) + 6e^- \longrightarrow 2Fe(s)$

$3Zn(s) + 2Fe^{3+}(aq) + \cancel{6e^-} \longrightarrow 3Zn^{2+}(aq) + \cancel{6e^-} + 2Fe(s)$

$3Zn(s) + 2Fe^{3+}(aq) \longrightarrow 3Zn^{2+}(aq) + 2Fe(s)$

Example 3

Balance the equation below.

$H_2(g) + I^-(aq) \longrightarrow H^+(aq) + I_2(s)$

Hydrogen is oxidised and iodine is reduced.

Oxidation half equation

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

Reduction half equation

$2I^-(aq) + 2e^- \longrightarrow I_2(s)$

Combine the two half equations then simplify to get balanced redox equation.

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

$2I^-(aq) + 2e^- \longrightarrow I_2(s)$

$H_2(g) + 2I^-(aq) + \cancel{2e^-} \longrightarrow 2H^+(aq) + \cancel{2e^-} + I_2(s)$

$H_2(g) + 2I^-(aq) \longrightarrow 2H^+(aq) + I_2(s)$

Self assessment exercise 3.3

1. Balance the following redox equations.
   (a) $Al(s) + Fe^{2+}(aq) \longrightarrow Al^{3+}(aq) + Fe(s)$
   (b) $Na(s) + Mg^{2+}(aq) \longrightarrow Na^+(aq) + Mg(s)$
   (c) $Li(s) + Cr^{3+}(aq) \longrightarrow Li^+(aq) + Cr(s)$
2. Balance the following half equaiton.
   $O_2(g) + 2H^+(aq) \longrightarrow H_2O_2(aq)$

3.5 Order of reactivity of metals

Some metals react with cold water and with dilute acids. Others do not react with either water or acids. Based on the rate of reaction of metals with water and acids, we can arrange metals in order of their reactivity.

Reactions of metals with cold water

Experiment 3.1

Aim: *To investigate how metals react with water.*

Apparatus and reagents

- Pair of tongs
- Ceramic tiles
- Filter paper
- Knife/ scalpel

- Glass trough
- Universal indicator or any suitable indicator
- Litmus papers
- Water
- Sodium in a bottle of oil, magnesium ribbon, calcium, aluminium foil, zinc granules and copper.

Procedure

1. Half-fill the trough with cold water.
2. Add 2-3 drops of the indicator. What colour change do you observe? Indicate the pH of the solution as neutral, alkaline or acidic.
3. Record the observations in your note book.
4. Using a pair of tongs, remove a small piece of sodium from the bottle, place it on a ceramic tile covered with a filter paper to absorb excess oil.
5. Cut a very small piece of sodium (pea size) and return the rest into the bottle.
6. Drop the small piece of sodium using the pair of tongs into the glass trough with water. What do you observe?
7. Repeat the experiment using other metals such as calcium, cleaned magnesium ribbon, aluminium foil, zinc granules and copper.
8. Test the resultant solutions using red and blue litmus papers, note any colour change. Record your observation and arrange the metals in order of their reactivity with water.

Caution: Do not expose your eyes directly to the reaction point. Make observations from a reasonable distance.

Sodium

The water changes colour depending on the indicator solution used. When the sodium is placed on top of a floating piece of filter paper and it is dropped in water, it moves around the surface of the water and finally it catches fire.

Sodium reacts **vigorously** with water than other metals to form hydrogen and sodium hydroxide which is an alkaline solution. Sodium placed on the filter paper ignites because the heat produced is sufficient to ignite the hydrogen gas liberated.

Sodium + water $\longrightarrow$ sodium hydroxide + hydrogen

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

This is also a **redox** reaction. First, water undergoes slight ionisation to form hydrogen ions, (H^+) i.e

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

Then sodium atoms lose electrons and are thus oxidised to form sodium ions, (Na^+).

Oxidation half equation

Sodium $\longrightarrow$ sodium ions + electrons

(i) $Na(s) \longrightarrow Na^+(aq) + e^-$

The electrons are then picked up by H^+ ions reducing them to hydrogen gas i.e.

Reduction half equation

Hydrogen ions + electrons $\longrightarrow$ hydrogen gas

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

Since two H^+ ion need 2e- to form H_2 (g), the first equation (i) is multiplied by two i.e. the number of electrons lost and those gained must be equal.

(iii) $2Na(s) \longrightarrow Na^+(aq) + e^-$

$2Na(s) \longrightarrow 2Na^+(aq) + 2e^-$

Adding the equations (iii) and (ii) after balancing, we obtain the following overall redox reaction equation.

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

$2H^+(aq) + \cancel{2e^-} \longrightarrow H_2(g)$

$2Na(s) + 2H^+(aq) \longrightarrow 2Na^+(aq) + H_2(g)$

Calcium

Calcium reacts **less vigorously** with cold water to give calcium hydroxide and hydrogen gas-which burns with a pop sound i.e.

Calcium + water $\longrightarrow$ calcium hydroxide + hydrogen

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

Magnesium

Very clean magnesium reacts **slowly** with cold water to form magnesium hydroxide (which turns red litmus blue) and hydrogen gas, (which burns with a pop sound). The reaction stops after sometime because the magnesium hydroxide formed is almost insoluble in water and forms a coat on the magnesium preventing further reaction.

Magnesium + water $\longrightarrow$ magnesium hydroxide + hydrogen

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

Aluminium, zinc and copper do not react with water.

Caution: Reaction of potassium with water should never be tried in the laboratory. Potassium reacts **explosively** with water.

However, just like sodium, potassium reacts with water to form potassium hydroxide solution and hydrogen gas.

Potassium + Water ⟶ Potassium hydroxide + Hydrogen

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

Self assessment exercise 3.4

1. Calcium metal reacts with cold water. This is a redox reaction.
   (i) Write an ionic equation for the reaction.
   (ii) Which are the oxidising and reducing agents in this reaction.
2. A metal X was burned in air. It burnt with a yellow flame.
   (a) Is this a reduction or oxidation reaction? Explain.
   (b) Write an equation for the reaction.
3. In the following reactions, suggest the oxidising and reducing agents.
   (a) $PbO(s) + C(s) \longrightarrow Pb(s) + CO(g)$
   (b) $2Y(s) + O_2(g) \longrightarrow 2YO(s)$
   (c) $H_2S(g) + Cl_2(g) \longrightarrow S(s) + 2HCl(g)$

Reaction of metals with acids

Experiment 3.2

Aim: *To investigate the reaction of metals with dilute hydrochloric acid.*

Apparatus and reagents

- Boiling tubes
- Bunsen burner
- Wooden splint and a match box
- 2M hydrochloric acid
- Clean piece of magnesium ribbon, calcium, aluminium foil, zinc granules and copper.

Procedure

1. Put 5 cm^3 of dilute hydrochloric acid provided in a boiling tube.
2. Add a piece of magnesium ribbon . What do you observe?

3. Test the gas evolved with a burning splint. What do you observe?
4. Repeat the experiment with pieces of calcium, aluminium foil, zinc granules and copper.
5. Arrange the metals in order of reactivity starting with the most reactive.

When the piece of magnesium is placed into the acid, effervescence occurs, indicating that a gas is evolved. When the gas is tested with a burning splint, it is put off with a pop sound.

Most metals react with dilute acids liberating hydrogen gas and a forming salt solution. When a metal reacts with an acid the reaction that occurs is a **redox** reaction. The metal atoms lose electrons i.e. the atoms are oxidised. The electrons lost are gained by aqueous ions from the acid i.e. H^+ ions are reduced.

Consider the reaction of magnesium and hydrochloric acid.

Magnesium + hydrochloric acid $\longrightarrow$ magnesium chloride + hydrogen

$$Mg\,(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Half equations

Oxidation half equation

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

Reduction half equation

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

When the two half equations are added we get the overall redox ionic equation as:

Overall ionic equation: $Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}\,(aq) + H_2(g)$

Very dilute hydrochloric acid reacts vigorously with calcium to form calcium chloride and hydrogen gas. Aluminium and zinc react slowly with dilute hydrochloric acid to form corresponding salts and hydrogen gas. Some metals react more vigorously with the acid than others. Not all metals react with dilute acids e.g. copper does not react with dilute acids.

From the results of the experiments, it is possible to arrange metals in order of their reactivity starting with the most reactive. This order of reactivity is referred as **reactivity series**.

Below is an example of a reactivity series:

Potassium
Sodium
Calcium
Magnesium
Aluminium
Zinc
Iron
Lead
Copper

As mentioned earlier metals react by losing electrons. The ease with which the metals lose electrons determines their position in the reactivity series. A metal which loses electrons easily is positioned at the top while the others follow in that order as shown above.

Self assessment exercise 3.5

1. (a) Write an ionic equation to show the redox reaction between:
   (i) zinc and dilute sulphuric acid.
   (ii) iron and dilute hydrochloric acid.
   (b) Name the oxidising and reducing agents in (a) i and (a) ii above.
2. Which of the following reactions are oxidation reactions, reduction reactions both oxidation and reduction (redox) reactions and which are none? Explain.
   (i) $K^+(aq) + e^- \longrightarrow K(s)$
   (ii) $Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$
   (iii) $Pb^{2+}(aq) + 2Cl^-(aq) \longrightarrow PbCl_2(s)$
   (iv) $Cu^+(aq) \longrightarrow Cu^{2+}(aq) + e^-$
   (v) $Ag^+(aq) + Cu^+(aq) \longrightarrow Ag(s) + Cu^{2+}(aq)$
   (vi) $2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$
   (vii) $Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$

3.6 Spontaneity of a reaction

Can we be able to predict accurately which way reaction will proceed? For example, if we mixed two solutions or a salt solution with a solid or a metal, can we accurately say which reagent will be reduced or oxidised?

Reactive metals displace less reactive metals from their solutions. This kind of reaction is known as **redox** reaction since the electrons lost are gained by the metal ions of the less reactive metal in the solution. Therefore, the metal donating the electrons is **oxidised**, while the metal ions of the less reactive metal in solution gaining electrons are **reduced** as we have already learnt. The metal atom losing electrons is the reducing agent and the substance gaining electrons is the oxidising agent. Such reactions involved displacement of ions from their solutions.

The general reaction that takes place during displacement reactions occur in two half reactions. Suppose we represent a metal with letter M:

Oxidation half equation

(i) $M(s) \longrightarrow M^{n+}(aq) + ne^-$ where n = 1, 2, 3

Reduction half equation

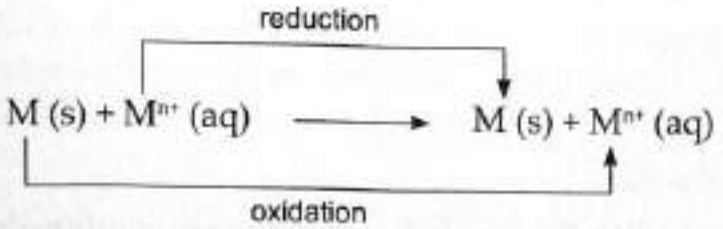
(ii) $M^{n+}(aq) + ne^- \longrightarrow M(s)$

In equation (i) above, the more reactive metal atoms lose electrons i.e. the atoms are oxidised.

In equation (ii) the less reactive metal ions in the solution gain electrons lost by the more reactive metal atoms i.e the ions are reduced.

Adding equations (i) and (ii) gives us the overall equation:

Redox reaction



Consider the following displacement reactions:

(a) Reaction of zinc and copper sulphate

In a reaction where zinc displaces copper from copper sulphate solution;
zinc dissolves and in the process, it loses two electrons. This is an **oxidation** reaction.

Oxidation reaction

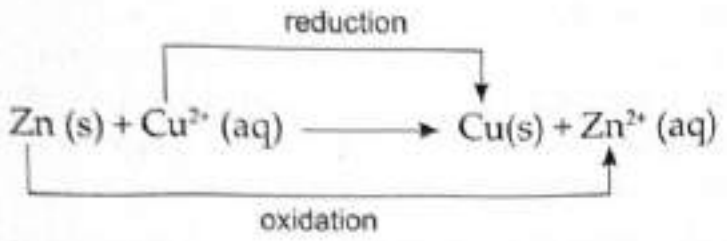
$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

Copper ions, (Cu^{2+}) in solution gain the electrons and copper metal is deposited. In the process the blue colour of the solution **fades** or **disappears**. This is a **reduction** reaction.

Reduction reaction

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

Adding the two half equations, the overall ionic equation is obtained:



Zinc is oxidised by losing two electrons. Its oxidation state also increases from 0 to +2 but copper ions are reduced (gain $2e^-$) and its oxidation number decreases from +2 to 0 (zero).

In the reaction considered above, copper ion Cu^{2+} is referred to as **oxidising agent**. This is because it removes electrons from zinc to form Cu atom. Zinc metal is the **reducing agent** as it loses electrons to copper ions.

(b) Reaction of iron and copper sulphate

If Iron metal is put in a beaker of copper sulphate solution, a reaction takes place.

Iron + copper sulphate ⟶ iron sulphate + copper

$$Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$$

From the overall equation, we can get two half equations; oxidation half equation and reduction half equation as follows:

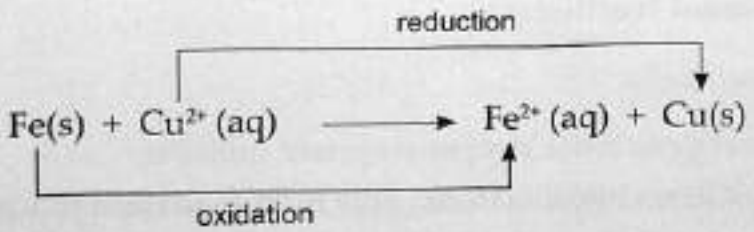
Oxidation half equation

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

Reduction half equation

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

Overall redox equation is obtained when the two half equations are combined:



We know that zinc and iron are more reactive than copper. They can therefore displace copper from copper sulphate solution.

Note: The order of reducing power is the same as the order of reactivity i.e the most reactive metal in the reactivity series has the highest reducing power while the least reactive has the lowest reducing power. This can be used to predict the direction in which a reaction will proceed.

Displacement reactions involving halogens

Halogens also displace each other from their respective salt solutions according to their position in the reactivity series. This happens by oxidation of **halide ions** in the solution. For example, chlorine is more reactive than iodine. Therefore, if we bubble chlorine through a solution of potassium iodide, iodine is displaced. Iodide ions, (I^-) in the solution are oxidised (lose electrons). The electrons are gained by chlorine atoms (Cl) and become chloride ions (Cl^-) as shown below.

Oxidation half equation

$$2I^-(aq) \longrightarrow I_2(s) + 2e^-$$

Reduction half equation

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$$

Overall redox equation

reduction

$$2I^-(aq) + Cl_2(g) \longrightarrow I_2(s) + 2Cl^-(aq)$$

oxidation

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Note that chlorine is also more reactive than bromine. Chlorine can also displace bromine from potassium bromide solution. We therefore say that chlorine is a more powerful oxidising agent than both bromine and iodine. The order of oxidising power is the same as the order of reactivity. This information can also be used to predict the direction of reactions involving halogens.

Self assessment exercise 3.6

1. What is a redox reaction in terms of electron transfer.
2. Write the ionic equation for the reaction between:
   (i) Solid iron and copper sulphate solution.
   (ii) Chlorine gas and potassium iodide solution

3.7 Electrochemical cell

Earlier, we mentioned that a chemical reaction can generate an electric current. This is what happens in an electrochemical cell.

We will now see how an electric current is generated by the chemical reaction by studying an electrochemical half-cell below, which is a metal electrode dipped into an electrolyte.

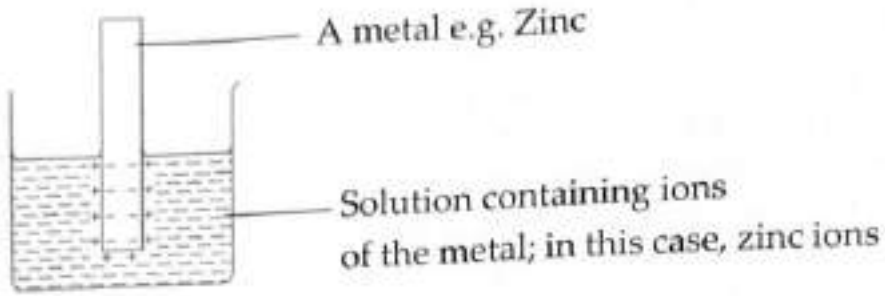


Fig 3.1: An electrochemical half-cell

A metal dipped in an electrolyte as shown in Fig. 3.1 above tends to form ions. If a metal, M, is placed in a solution containing M^{n+} ions, of the same metal, the metal atoms may lose n electrons and go into solution as positive ions i.e.

$$M(s) \longrightarrow M^{n+}(aq) + ne^- \text{ where; } n = 1,2,3$$

The electrons, which are released accumulate on the metal surface, which becomes negatively charged. The metal ions in the solution increase until some start to gain electrons on the metal surface and form atoms.

$$M^{n+}(aq) + ne^- \longrightarrow M(s)$$

The two reactions above take place at the same time and at the same rate. This is represented as shown in the following equation.

$$M(s) \rightleftharpoons M^{n+}(aq) + ne^-$$

If metal M is zinc, the equation would be written as follows:

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

The ease with which metals lose electrons depend on their position in the reactivity series. A metal, which loses electrons easily, will acquire a much bigger negative charge on its surface, whereas a metal which loses electrons less easily will have a much smaller negative charge on its surface as shown in Fig. 3.2.

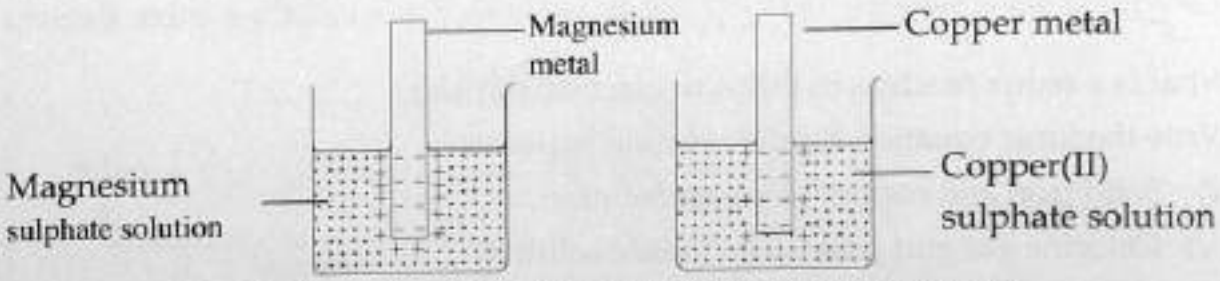


Fig 3.2: Comparison of the number of electrons on metal surfaces

The electrons which are released, accumulate on the metal surface. The metal becomes negatively charged and attracts a layer of positive ions. Two layers of opposite charges are formed. This results in a potential difference between the metal and a solution of the metal ions. This potential difference is reffered to as the **electrode potential** of the metal. The potential difference depends on the tendency of metals to form ions. The greater the tendency for the metal to ionise, the greater the potential difference. For example, in Fig. 3.2, magnesium forms ions more easily than copper. The magnesium half-cell therefore, exhibit a higher potential difference. An arrangement like we have

in Fig. 3.2 where a metal strip is dipped in its salt solution is called a **half-cell**. It is not possible to measure the electrode potential of one electrode but we can measure the difference in potential of two electrodes. This can be done by connecting the two electrodes by a metal wire and connecting the two solutions by a salt bridge in such a way that ions can flow between them as shown in Fig. 3.3.

Each beaker with the metal rod and salt solution is the half-cell. Two half-cells connected through the two electrodes using a metal wire and the two solutions with a salt bridge containing a salt solution as shown in Fig. 3.3 above is known as an **electrochemical or voltaic cell**.

Now, let us see how the chemical reactions generate electrical energy in zinc–copper electrochemical cell.

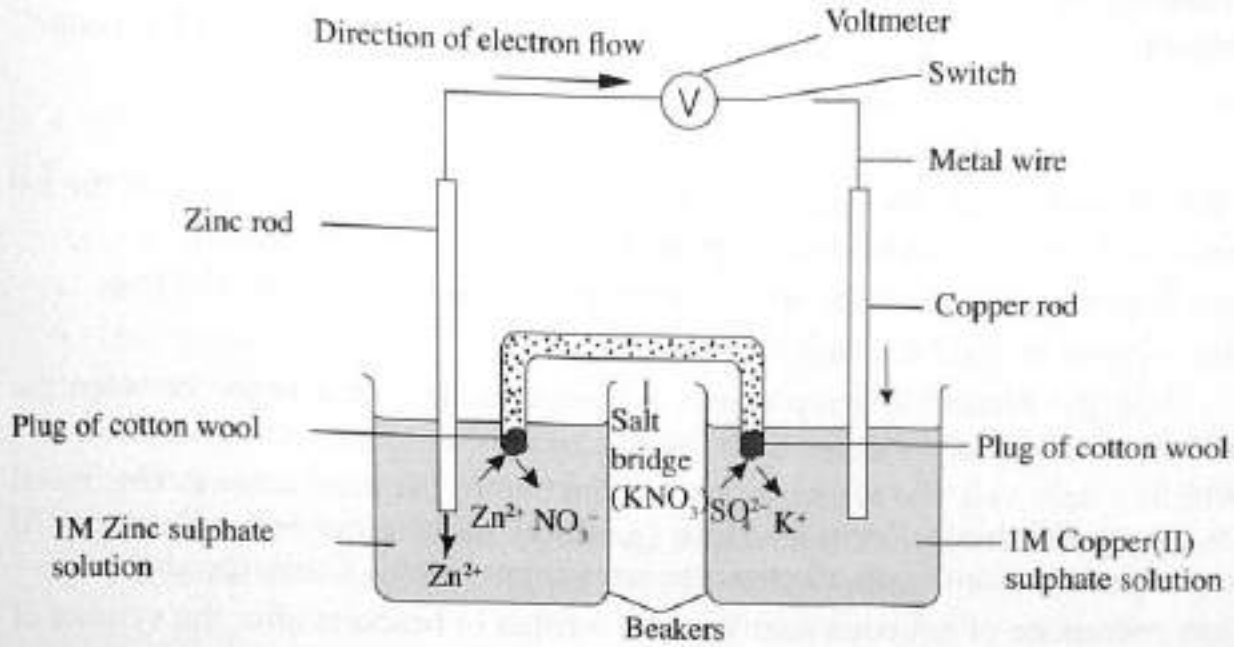


Fig. 3.3: An electrochemical cell

In the left-hand side beaker as shown in Fig. 3.3, the zinc rod dissolves by losing two electrons to form zinc ions, $Zn^{2+}(aq)$.

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

The electrons then flow through the external wire to the right-hand side beaker. The copper(II) ions, $Cu^{2+}(aq)$, take these electrons and copper metal is deposited.

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

The overall reaction is as follows:

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

The overall reaction is an oxidation-reduction reaction. Oxidation always occurs at the **anode** and reduction always occurs at the **cathode** of an electrochemical cell.

The $Zn(s)|Zn^{2+}(aq)$ half-cell is therefore the anode, and the $Cu^{2+}(aq)|Cu(s)$ half-cell is the cathode.

The electrons are transported by the ions (anions) present in the solution up to the electrode and through the wire to the other electrode. Look at Fig 3.3 and see how the $Zn^{2+}(aq)$ and SO_4^{2-} ions move into the salt bridge, while K^+ and NO^-_3 ions diffuse out of

the salt bridge into the solution. Since electricity is a flow of electrons, the arrangement as in Fig. 3.3 is called an **electrochemical cell**. The potential difference between the two metal electrodes is measured by a voltmeter. It gives the voltage of the cell, also known as the ***electromotive force*** (e.m.f). It is measured in volts and given the symbol, E. The e.m.f of zinc - copper cell is 1.10V. The e.m.f varies with temperature, concentration of the solution and state of the metal used.

The electrochemical cells as shown in Fig. 3.3 can be described by a line notation as illustrated below.

$Zn(s) \mid Zn^{2+}(aq)\ (1M) \mid\mid Cu^{2+}(aq),\ (1M) \mid Cu\ (s)$

metal conducting in left-hand beaker.	ions in contact with metal in left-hand beaker.	salt bridge	ions in contact with metal in right-hand beaker.	metal conducting in right-hand beaker.

Note: By convention, (i) the half-cell of the more reactive metal is placed on the left hand side. (ii) Hydrogen electrode, if part of the cell is placed on the left.

It is very important to follow the above shorthand notation of the cell. The line notation is based on the following conventions:

- | – A single vertical line represents a change in state or a phase between the metals and their ions e.g. $Zn(s) \mid Zn^{2+}(aq)$.
- Within a half- cell, the reactants are written before the products e.g. zinc metal (reactant) dissolves to form zinc ions (product), in the other half-cell, copper(II) ions, Cu^{2+}, (reactant) gain, electrons to form copper metal, Cu(s), (product).
- Concentrations of aqueous solutions are written in brackets after the symbol of the ion (sometimes the concentrations are omitted).
- || – A double vertical line or a double dotted vertical line indicates the salt bridge or a porous barrier connecting the half-cells.
- By convention, the more electropositive electrode (one that does not ionise easily) is put on the right-hand side.

The e.m.f. of the cell is given by:

$$E_{cell} = E\ (\text{right-hand side electrode}) - E\ (\text{left-hand side electrode})$$

Or simply,

$$E_{cell} = E_{RHS} - E_{LHS}.$$

The salt bridge can be a U-tube filled with electrolyte e.g.KNO_3, $NaNO_3$ or a filter paper soaked in the electrolyte. It serves two main functions.

I. Completes the circuit by allowing ions carrying charge to move from one half-cell to the other.

II. It provides positive ions (cations) and negative ions (anions) to replace those used up at the electrodes and balance the charges on ions formed at the electrodes. For example, at the cathode, Cu^{2+} ions gain electrons to form copper metal Cu(s)

at the cathode. This leaves many negative sulphate, SO_4^{2-} ions. If $CuSO_4$ solution was used in the second half-cell, potassium, K^+ ions from the salt bridge flow to balance the increased negative charge at the cathode.

What happens when the electrochemical cell is used to do work?

- The zinc electrode becomes thinner and thinner as zinc atoms are oxidised to Zinc ions (Zn^{2+}), which go into solution.
- The copper electrode becomes thicker and thicker as copper(II) ions, Cu^{2+} in the solution are reduced to copper metal which get deposited on the electrode.
- The concentration of Zn^{2+} ions at the anode increases and the concentration of the Cu^{2+} ions at the cathode decreases.

Over a period of time, all the zinc dissolve and the cell eventually has to be replaced.

Experiment 3.3

To measure the voltage, (e.m.f.) of an electrochemical cell using different electrodes.

Apparatus and chemicals

- beakers
- filter papers
- connecting wires
- voltmeter capable of measuring positive and negative voltage 0-3V with zero – centre.
- metal electrodes - copper, zinc, iron, magnesium, aluminium, silver.
- the respective metal salt solution of the above metals

Procedure

1. Assemble the apparatus as in Fig. 3.4

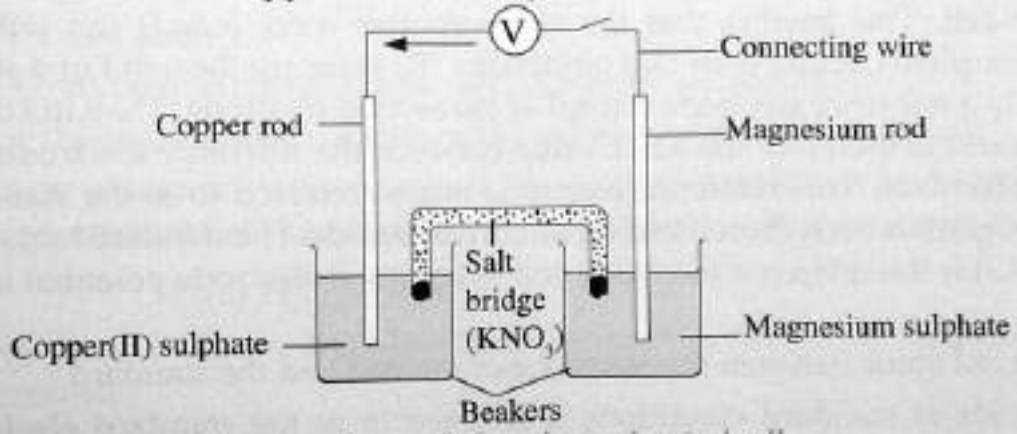


Fig. 3.4: Measuring the e.m.f of an electrochemical cell

2. Use copper electrodes throughout the experiment in the left-hand beaker.
3. In the beaker on the right-hand side, use an electrode of a different metal, which must be inserted in its own electrolyte.
4. Note the voltage shown by the deflection in the voltmeter. Note whether the deflection is –ve or +ve.

5. Record the voltage in your notebook. Indicate the voltage with (–) or (+) sign.
6. Arrange the list of the metals plus their voltage in your notebook, starting with most negative (smallest value) and end with most positive (largest value).
7. Repeat procedure 4 and 5 using the different metals in their salt solutions.
8. What do you conclude?
   - From your knowledge of reactivity series, which is the most reactive metal?

The list you have obtained is called *electrochemical series*. The most reactive metal has the largest negative (-ve) voltage. This is the metal that loses electrons most easily. We can conclude that the voltage difference is a measure of how easily metals lose electrons forming ions. It also shows the strongest reducing agent. We are going to discuss the electrochemical series in detail later in this unit.

Self assessment exercise 3.7

1. State briefly what you understand by the following terms.
   (i) electrode potential.
   (ii) half-cell.
   (iii) electromotive force.
2. (i) How is a salt bridge represented?
   (ii) Write the line notation of cell made of the following metals
   (a) $Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$ – 0.44V
   (b) $Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$ – 0.13V

Standard electrode potentials

It has already been mentioned that it is impossible to measure the electrode potential of a single half-cell. This implies that the electromotive force (e.m.f) can only be measured for complete circuits with two electrodes. To measure the e.m.f of a single electrode, usually a reference electrode is used as the second electrode. The e.m.f of the electrode of interest is therefore the e.m.f value between the reference electrode and that particular electrode. This reference electrode is also referred to as the **standard electrode.** Hydrogen has been chosen as the standard electrode. The standard hydrogen electrode (Fig. 3.5) is the reference electrode and is assigned electrode potential value of zero.

The measured e.m.f value between the electrode of interest and the standard hydrogen electrode at standard conditions is referred to as the **standard electrode potential** denoted by E^{θ}.

Note: The other half cells are usually compared to this standard half-cell (Fig. 3.5).

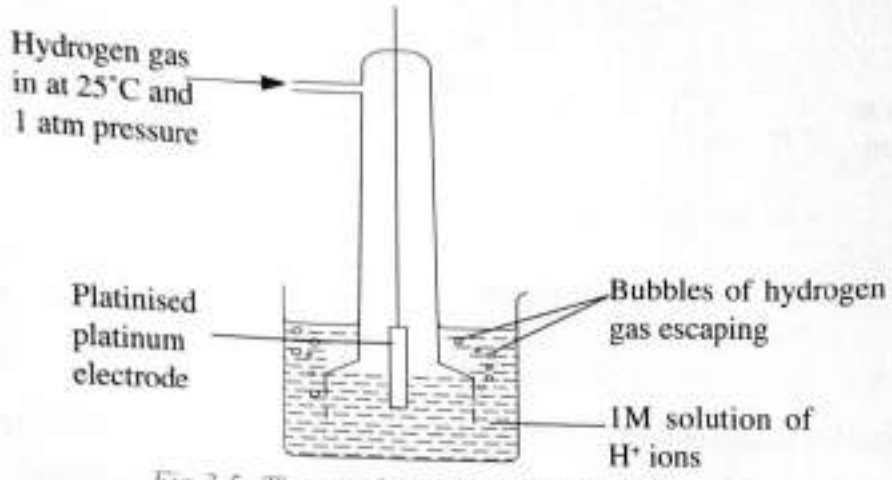


Fig 3.5: The standard hydrogen half-cell

The hydrogen half-cell

The standard hydrogen half-cell consists of the following:

- Hydrogen gas at 1 atmosphere pressure and 25°C temperature.
- Platinised-platinum electrode immersed in 1M H^+ ions solution.

The platinised-platinum electrode is coated with finely divided platinum (i.e. it is platinised). This increases the surface area so that a lot of hydrogen gas can be adsorbed on the platinum and an equilibrium is established between the adsorbed layer of H_2 (g) and H^+ ions in the solution as shown in the equation below.

$$\frac{1}{2} H_2(g) \rightleftharpoons H^+(aq) + e^- \quad E^\theta = 0.00 \text{ volt}$$

The standard electrode potential, E^θ, for the reference electrode is taken as zero. The superscript θ denotes standard conditions. Where standard conditions are:

- 25°C temperature
- 1 atm pressure
- hydrogen electrode immersed in 1 M H^+ ion solution
- 1M concentration of ions in the solution of the other half-cell.

The inert platinum electrode acts as the connection between hydrogen gas H_2 (g) and hydrogen ions, H^+ in the solution. The hydrogen electrode is represented as follows:

$$Pt(s)\ H_2(g) \mid H^+(aq)$$

Let us now see how the standard electrode potential of various metals can be determined.

The standard electrode potential of zinc

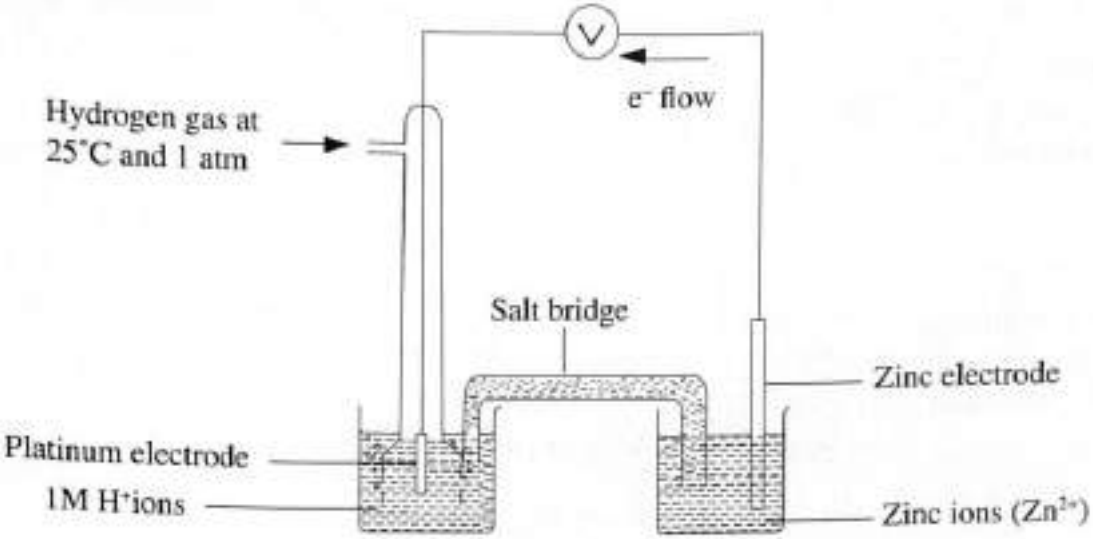


Fig. 3.6: Measuring the standard electrode potential of Zn(s) | Zn^{2+}(aq) half-cell

In Fig. 3.6, zinc is placed in 1M Zn^{2+} ion solution and connected to standard hydrogen half-cell. The cell diagram is represented as

$$Pt(s), H_2(g) \mid H^+(aq) \parallel Zn^{2+}(aq) \mid Zn(s) \qquad E^\theta = -0.76 \text{ volts}$$

Thus the standard electrode potential, E^θ, for zinc is – 0.76V.

Zinc loses electrons more readily than hydrogen. Therefore, zinc is the negative electrode. Zinc reduces H^+ions in aqueous solution to hydrogen gas (H_2) by donating electrons i.e.

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^- \quad E^\theta = +0.76V$$

$$2H^+(aq) + 2e^- \longrightarrow H_2(g) \quad E^\theta = 0.00V$$

The overall reaction of the cell is obtained by adding the two equations above i.e

$$Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g) \qquad E^\theta = +0.76V$$

Note: The two electrons cancel out as they occur on both sides of the equation.

The standard electrode potential for the half- cell $Zn^{2+}(aq) \mid Zn(s)$ is – 0.76V

(b) The standard electrode potential of copper

When the hydrogen electrode is connected to a copper electrode, the electrons flow in the opposite direction as shown in Fig. 3.7.

The electron flow is from left half-cell to the right half-cell in the external circuit i.e. from the hydrogen electrode to the copper electrode. This is because copper does not lose electrons as readily as hydrogen i.e copper is lower than hydrogen in the reactivity series. In this case, copper is the positive electrode. The standard electrode potential, E^θ of copper is +0.34V. The cell can be represented as:

$$Pt(s)\ H_2(g) \mid H^+(aq) \parallel Cu^{2+}(aq) \mid Cu(s) \qquad E^\theta = +0.34V$$

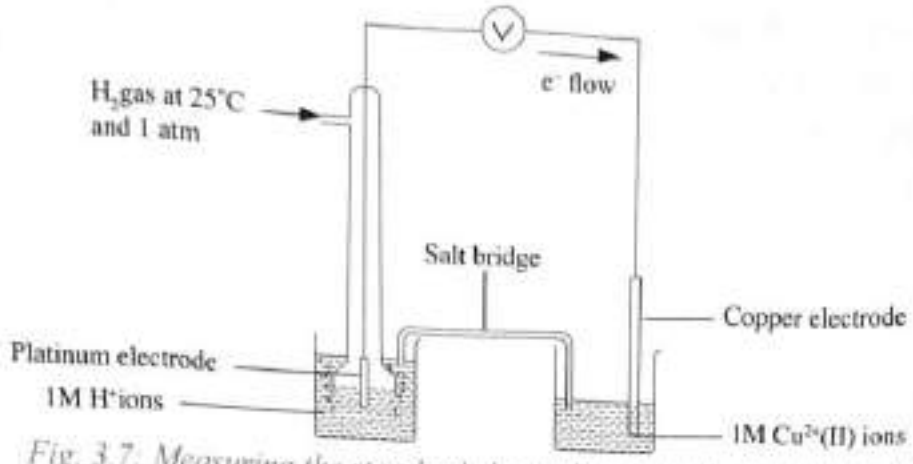


Fig. 3.7. Measuring the standard electrode potential of Cu(s) | Cu2+(aq) cell

Note that the positive e.m.f. of the cell tells us that the right-hand side electrode is positive. We can therefore define the standard electrode potential as follows:

The standard electrode potential, E^0, is the potential of that half-cell relative to a standard hydrogen half-cell under standard conditions.

Standard electrode potentials, E^0, provide a measure of the relative oxidising and reducing power of different species. Electrode systems are written showing particles gaining electrons. Gain of electrons is reduction; hence sometimes they are referred to as **reduction potentials**. When all the redox systems are arranged in order of the magnitude of their standard electrode potentials starting with the highest to the lowest, the arrangement is called, the **electrochemical series**. Table 3.4 is an example of electrochemical series. Notice that the relative value of E^0 gives a measure of the strengths of the oxidising and reducing agents.

Table 3.4 Electrochemical series

Reaction			E^0, volts	
$K^+(aq) + e^-$	⇌	$K(s)$	– 2.92	*Strongest reducing agent*
$Ca^{2+}(aq) + 2e^-$	⇌	$Ca(s)$	– 2.87	
$Na^+(aq) + e^-$	⇌	$Na(s)$	– 2.71	
$Mg^{2+}(aq) + 2e^-$	⇌	$Mg(s)$	– 2.36	
$Al^{3+}(aq) + 3e^-$	⇌	$Al(s)$	– 1.66	
$Zn^{2+}(aq) + 2e^-$	⇌	$Zn(s)$	– 0.76	*Reducing power increases*
$Fe^{2+}(aq) + 2e^-$	⇌	$Fe(s)$	– 0.44	
$Ni^{2+}(aq) + 2e^-$	⇌	$Ni(s)$	– 0.25	
$Pb^{2+}(aq) + 2e^-$	⇌	$Pb(s)$	– 0.13	
$2H^+(aq) + 2e^-$	⇌	$H_2(g)$	0.00	
$Cu^{2+}(aq) + 2e^-$	⇌	$Cu(s)$	+0.34	
$I_2(s) + 2e^-$	⇌	$2I^-$	+0.54	
$Ag^+(aq) + e^-$	⇌	$Ag(s)$	+0.80	*Oxidizing power increases*
$Br_2(l) + 2e^-$	⇌	$2Br^-(aq)$	+1.09	
$Cl_2(g) + 2e^-$	⇌	$2Cl^-(aq)$	+1.36	
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	⇌	$Mn^{2+}(aq) + 4H_2O(l)$	+ 1.51	
$F_2(g) + 2e^-$	⇌	$2F^-(aq)$	+2.87	*Strongest oxidizing agent*

Note: The table only shows the **conventional** way of listing standard electrode potentials, whether they are metals or non-metals. During calculations, if we reverse the equation for a metal, the sign in front of the E^θ value will also change. For example:

$$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s) \qquad E^\theta = -0.76V$$

To show that zinc reacts by donating electrons, we reverse the equation as follows.

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^- \qquad E^\theta = +0.76V.$$

Note the following about the E^θ values in Table 3.4:

- The more negative the E^θ value, the more reactive the metal. It gives electrons easily, hence it is a powerful reducing agent. It is also relatively difficult for the ions formed to gain electrons. For example, if this metal is potassium(K) it is difficult for potassium(K^+) ion to gain electrons.
- The more positive the E^θ value is, the less reactive the metal. The metal atoms do not donate electrons easily.An example of this is iron(Fe)
- For non-metals, the more +ve the E^θ value, the more readily that element gains electrons. Since electrons are removed from another species, those elements are strong oxidising agents. In Table 4.4. fluorine has the largest positive E^θ value; +2.87V. It is the strongest oxidising agent in the table. It gains electrons most easily.

$$F_2(g) + 2e^- \longrightarrow 2F^-(aq) \qquad E^\theta = +2.87V.$$

To make it easy for us to understand, let us put some of the E^θ values in Table 4.4 in a number line as shown below.

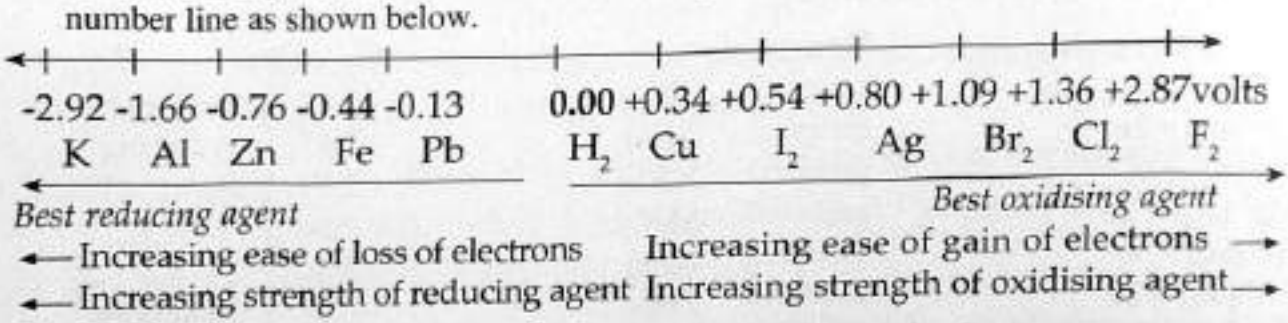


Fig. 3.8: Some electrode potentials on the number line

What is the use of electrode potentials?

- to calculate e.m.f of a cell.
- to predict whether or not a redox reaction will occur.

Sample calculations using E^θ values

Place the E^θ values in a number line as shown above. Check the E^θ value on the right hand side, (RHS) and subtract the E^θ value on the left hand side (LHS).

By convention, the e.m.f of the cell is obtained from the following relationship.

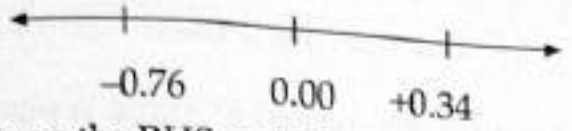
$$\mathbf{E^\theta\ cell = E^\theta_{(RHS\ electrode)} - E^\theta_{(LHS\ electrode)}}$$

Note: If the final value above is +ve, it indicates that the reaction will occur.

Example 1

Calculate the e.m.f. of an electrochemical cell, Zn(s) | Zn2+(aq) || Cu2+ | Cu(s). Refer to Table 4.4 for Eθ values.

Solution



Cu is on the RHS and Zn on the LHS of the number line, therefore e.m.f of the cell

$= E^{\theta}_{Cu} - E^{\theta}_{Zn}$

$= +0.34 - (-0.76)$ V

$= +1.10$ V

Note: Whenever, you choose a pair of electrodes to construct a cell, you will always have one on RHS and the other on LHS on the number line irrespective of whether they both have positive values or negative values or one has a positive value and the other, a negative value.

Remember the following:

- the more reactive a metal is, the more negative the E^{θ} value and is always placed on the LHS when we draw an electrochemical cell.
- oxidation-reduction reactions that have a positive overall cell potential are spontaneous.

Example 2

A student wanted to find out whether a reaction would occur or not. The student set up the experiment as shown in Fig. 4.9.

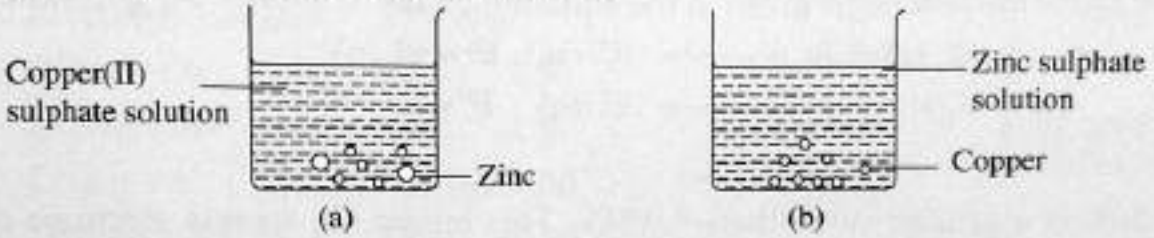


Fig. 3.9: Displacement of metals from their salt solutions

- Using the standard electrode potentials, E^{θ}, given on Table 3.4, explain in which beaker the reaction took place and state the observations the student made. Write an ionic equation for the reaction.

Solution

The problem demands that we use E^{θ} values to solve it.

From Table 3.4, the half-reactions given are:

$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)\ E^{\theta} = -0.76V.$

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)\ E^{\theta} = +0.34V.$

The –ve E^0 value indicates that this metal loses electrons easily. These electrons are taken up by Cu^{2+} ions(oxidising agent).This is the equation we must *reverse* first then add up the two equations as follows:

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^- \quad E^0 = +0.76V$$
$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s) \quad E^0 = +0.34V$$

Overall reaction:- $Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s) \quad E^0 = +1.10V$

The reaction will take place in beaker (a) because the overall E^0 value is **positive**, i.e +1.10V. Cu^{2+} ions are reduced by zinc and displaced from the solution as copper.

Observation: The blue colour fades and a red-brown solid, which is copper is deposited. The ionic equation for the reaction is the same as overall equation above.

Ionic equation

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

Now, let us reverse the half reaction of copper in beaker (b)

$$Cu(s) \longrightarrow Cu\ (aq) + 2e^- \quad E^0 = -0.34V$$
$$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s) \quad E^0 = -0.76V$$

Overall equation $Cu(s) + Zn^{2+}(aq) \longrightarrow Cu^{2+}(aq) + Zn(s) \quad E^0 = -1.10V$

The reaction cannot take place because E^0 value is **negative**, i.e.–1.10V.

A negative final E^0 means that the reaction is not possible. In this case, it means copper cannot displace zinc from its salt solution.

Example 3

Can chlorine displace iodine from a solution of potassium iodide?

Use electrode potentials given in the equation below to answer the question.

$$Cl_2\ (g) + 2e^- \longrightarrow 2Cl^-(aq) \quad E^0 = +1.36V$$
$$I_2(s) + 2e^- \longrightarrow 2I^-(aq) \quad E^0 = +0.54V$$

Solution

+1.36V is a greater value than + 0.54V. This means Cl_2 accepts electrons more readily than I_2 (see the 'number line' below).

We reverse the 2nd equation and add the two equations.

$$Cl_2\ (g) + 2e^- \longrightarrow 2Cl^-\ (s) \quad E^0 = +1.36V$$
$$2I^-(aq) \longrightarrow I_2(s) + 2e^- \quad E^0 = -0.54V$$

Overall equation: $Cl_2\ (g) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(s)\ E^0 = +0.82V$

The answer has a positive value. Therefore, chlorine can displace iodine from potassium iodide solution.

- Alternatively, we can put the values on a number line and subtract the smaller value (on LHS) from the greater value (on RHS) as shown below:

0.00	+0.54	+1.36
H_2	I_2	Cl_2

$$E^0 = E^0_{RHS} - E^0_{LHS}$$
$$E^0 = E^0_{Cl2} - E^0_{I2}$$
$$= +1.36 - 0.54$$
$$= +0.82V$$

- Now, show that iodine cannot displace chlorine from potassium chloride solution.

Self assessment exercise 3.8

Use the electrode potentials in Table 4.4

1. Which of the following species can be reduced by Ni?
   I_2, Cu^{2+}, Fe^{2+}
2. Calculate the standard e.m.f of each of the following cells.
   (a) $Ag(s) \mid Ag^+ \mid\mid Cu^{2+}(aq) \mid Cu(s)$
   (b) $Fe(s) \mid Fe^{2+}(aq) \mid\mid Cu^{2+}(aq) \mid Cu(s)$
   (c) $Zn(s) \mid Zn^{2+}(aq) \mid\mid Ni^{2+}(aq) \mid Ni(s)$
3. The letters A,B,C,D, and G given below do not represent actual symbols of elements.

   Use the E^θ values given to answer the questions that follow.

Electrode reaction		E^θ (volts)
$A^{2+}(aq) + 2e^-$	$\longrightarrow$	$A(s)$ -2.90
$B^{2+}(aq) + 2e^-$	$\longrightarrow$	$B(s)$ -2.38
$C^+(aq) + e^-$	$\longrightarrow$	$^1/_2C_2(g)$ 0.00
$D^{2+}(aq) + 2e^-$	$\longrightarrow$	$D(s)$ +0.34
$\frac{1}{2}G_2(g) + e^-$	$\longrightarrow$	$G^-(aq)$ +2.87

   (i) Which element is likely to be hydrogen? Give a reason for your answer.
   (ii) Which element is the strongest reducing agent? Give a reason for your answer.
   (iii) Write the representation of a cell obtained when A and D are combined and calculate e.m.f. of that cell.

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3.8 Corrosion and rusting

Corrosion is a reaction process which occurs when some metals are exposed to oxygen and moisture. In the process, the metal is dissolved (ionised) slowly and gradually wears out.

Most metals undergoing corrosion form oxides. When the corroding metal is iron, the process is called **rusting**. Rust is the name of the **oxide** compound formed when iron corrodes. It is usually a **brown substance**.

Experiment 3.4

Aim: *To investigate conditions necessary for rusting*

Apparatus and reagents

- Test tubes and test tube rack.
- Nails
- Distilled water and tap water
- Anhydrous calcium chloride
- Oil
- Cotton wool
- Emery paper
- Sodium chloride (common salt)
- Rubber bung

Procedure

1. Label five test tubes A, B, C, D and E.
2. Clean five iron nails thoroughly with emery paper.
3. Put one nail in each test tube and place the test tubes in a test tube rack.
4. Do not add anything else to test tube A.
5. Add tap water to cover the nail in test tube B.
6. Push a plug of cotton wool above the nail in test tube C. Add anhydrous calcium chloride. Close the test tube with a rubber bung.
7. Add boiled or distilled water to the fourth test tube D. Add a layer of oil and close the tube as shown in Fig. 3.10.
8. Add sodium chloride solution to test tube E. Leave the test tubes undisturbed for one week.

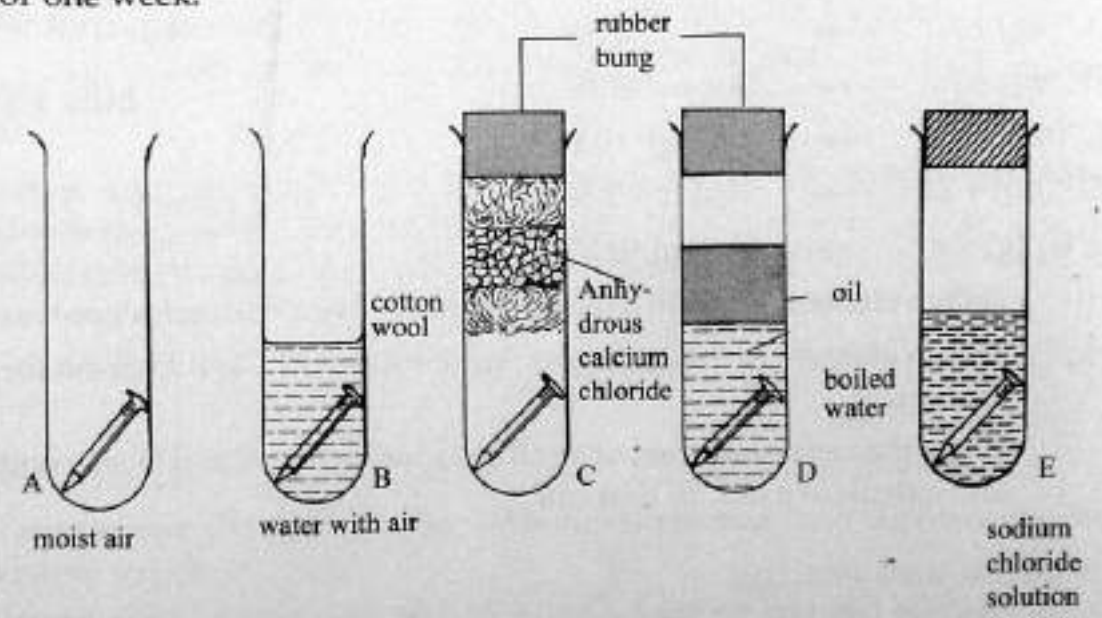

Fig. 3.10: Investigating the conditions necessary for rusting.

You must have observed that nails in test tubes C and D remained clean while those in test tubes A, B and E rusted. Rusting occurred in nails in test tube A and B because both oxygen and moisture are present. In test tube C, anhydrous calcium chloride absorbed water vapour in the air, making the air inside dry (no moisture). This is why rusting did not take place. In test tube D, boiling the water drove out the dissolved oxygen and the layer of oil prevented any air from dissolving into the water as it cooled. Therefore, rusting did not take place.

This experiment shows that the presence of both **water** and **oxygen** is necessary for rusting to take place. In the absence of any one of the two, iron will not rust.

A lot more rusting took place in test tube E which contained a solution of sodium chloride. Sodium chloride speeds up rusting process.

Factors that accelerate rusting

These include

- Presence of sodium chloride
- High temperatures
- Acidic conditions

Research Activity

Find out with a classmate how these factors affect the rate of rusting.

Sea water usually is salty i.e contains high levels of sodium chloride. Sodium chloride speeds the process of rusting and therefore rusting of ships is a major problem in seas. This also explains why cars used in coastal areas rust faster and get holes in their bodies than cars used further inland. In countries that experience cold winter seasons like South Africa, North America and Europe, the roads are often salted during winter to melt ice. However this accelerates the rusting process on vehicles bodies.

Note

- Rusting is of great economic importance. Large sums of money are spent each year to replace damaged structures made of iron and steel. This is because rust forms blisters which peel off thus exposing more iron or steel to rusting. This weakens structures such as bridges, church rafters and houses which can eventually collapse. Therefore, it is important to prevent rusting process when such structures are being made. This explains why most iron or usually painted.

How rust is formed in terms of redox reaction

When iron dissolves in water it forms ions by losing electrons i.e it undergoes oxidation.

$$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^- \quad \text{.....(i)}$$

These electrons are taken up by water and air i.e they are reduced.

$$2H_2O(l) + O_2(g) + 4e^- \longrightarrow 4OH^-(aq) \quad \text{.....(ii)}$$

Because hydroxide ions (OH^-) are formed, the process is strongly affected by presence of acids. At low pH, the process is accelerated as the OH^- reacts with the H^+ ions to form water and providing the electrons required for the oxidation process in equation (i) above.

To get the overall equation:

1. Equalise the number of electrons in the two equations by multiplying equation (i) by 2. We get:

   $2Fe\ (s) \longrightarrow 2Fe^{2+}(aq) + 4e^-$ .....(iii)

2. Adding equation (ii) and equation (iii) gives:

   $2Fe\ (s) \longrightarrow 2Fe^{2+}(aq) + \cancel{4e^-}$ .....(iv)

   $2H_2O + O_2(g) + \cancel{4e^-} \longrightarrow 4OH^-(aq)$ .....(v)

   (electrons on opposite side of the two equations cancel out).

3. The overall equation becomes:

   $2Fe(s) + 2H_2O + O_2(g) \longrightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$

The hydroxyl ions (OH^-) react with iron(II) ions (Fe^{2+}) and more air (O_2) to form rust which is hydrated iron(III) oxide ($Fe_2O_3 . nH_2O$). The ionic equation for this reaction is:

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

Methods of preventing rusting

We have seen that for iron to rust, oxygen and moisture must be present. We must therefore protect iron objects and structures from either oxygen, moisture or both. Most of the modern methods of preventing rusting make use of this principle.

Research Activity

Find out from text books in the library or the internet how rusting process can be prevented. Write a report and share with your class members.

Some methods of preventing rusting include:

1. Painting, greasing and oiling

Painting is a method of preventing rusting used in ships, bridges, cars and corrugated iron sheets. Rust should be removed before painting the surface that needs protection. If the layer of paint is scratched off, rust is formed below the surface. If rust begins under a paint coat, it will force the paint coat away from the metal by forming a blister which then splits, allowing fresh metal surface to be exposed.

Greasing and **oiling** are commonly used to protect moving engine parts from rusting. Such parts cannot be painted since the paint would get scratched off when the parts are in motion. Oiling and greasing is not a permanent method of preventing rusting and must be repeated frequently.

2. Coating with other metals or plastics

(a) Galvanising

This involves dipping iron sheets in molten zinc. Zinc reacts with air forming a thin coating of zinc oxide which protects iron from rusting.

Galvanising has an advantage over other plating methods because even if the zinc layer is badly scratched, still zinc protects iron from rusting. This is because zinc is more reactive than iron.

(b) Tin-plating

This involves dipping iron into molten tin to form a layer of tin on top of the iron. If the tin layer is scratched off, rusting proceeds very extensively under it. This is because iron is more reactive than tin. Tin coating is used in food cans because tin is **unreactive** and **non-toxic**.

(c) Plastic coating

Plastic coating is widely used to protect equipment made of iron from rusting. Some examples are refrigerators, garden chairs, pliers, dish racks, among others.

(d) Electroplating

Electricity is used to form a very thin coat of some metals on iron. Chromium, gold, silver, nickel and tin can be used. Besides protecting iron from rusting, electroplating makes iron articles **shiny** and **attractive**.

3. Alloying

Two or more molten metals are mixed together to form alloys. In this case, iron can be mixed with small proportions of molten chromium, nickel, carbon and other elements. Alloys do not rust and therefore can be used to make cutlery and other types of kitchenware.

4. Silica gel

Some equipment made of iron are packed and supplied with small bags of silica gel. The silica gel will absorb any moisture in the package and therefore prevent rusting.

5. Sacrificial protection

This is used to protect underground iron pipes from rusting, for example, zinc or magnesium blocks are attached to the pipes at intervals. Since these metals are more reactive than iron, they corrode in preference to the iron. This is called **sacrificial protection** because zinc or magnesium is **sacrificed** to protect iron. This principle is also used to protect the bodies of ship.

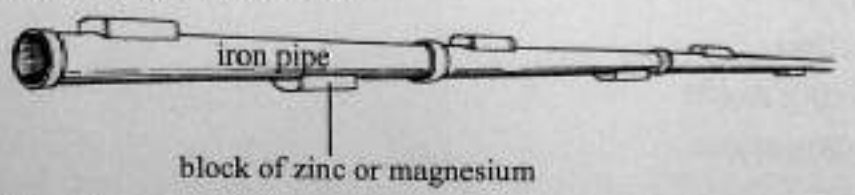

Fig. 3.11: Sacrificial protection of iron pipe by zinc or magnesium

Self assessment exercise 3.9

1. What is the chemical name for rust?
2. What are the conditions necessary for rusting to occur?
3. Name two factors that accelerate rusting.
4. A coat of paint by itself is not sufficient protection from rusting. Explain why?
5. Differentiate between rusting and corrosion of metals.
6. List the items made of iron in your home and school laboratory. With the help of your teacher or friend, write down the methods used to prevent them from rusting.
7. Explain why a piece of magnesium in contact with a steel pipeline prevents the pipeline from rusting.

Revision exercise 3

1. Name four methods apart from painting, greasing and oiling that can be used to prevent iron from rusting.
2. Name two metals that can be used to electroplate items and give two reasons for electroplating.
3. Describe the reaction between copper sulphate solution and zinc metal in terms of oxidation and reduction.
4. Using reactivity series, predict which of the following pairs would displace each other from their salt solutions and which would not be displaced.
   (a) Copper and iron
   (b) Aluminium and iron
   (c) Zinc and iron
   (d) Copper and aluminium
5. Distinguish between galvanising and electroplating
6. Explain what you understand by the following terms:
   i) Oxidation
   ii) Reduction
   iii) Oxidizing agent
   iv) Reducing agent

7. In terms of electron transfer, explain the meaning of oxidising and reducing power.

8. (a) What is the oxidation number (state) of manganese in the following substances?

   (i) Mn_2O_7

   (ii) MnO_4^-

   (b) What is the oxidation number of germanium in $GeCl_4$?

9. Calculate the oxidation number of carbon in the equation below and state whether the element has undergone oxidation or reduction.

$$C_2O_4^{2-}(aq) \longrightarrow 2CO_2(g) + 2e^-$$

Topic 4 Electrolysis

Sucess criteria

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

(a) Describe electrolysis.
(b) Write electrolytic half and overall equations.
(c) Predict products of the electrolysis of molten and aqueous ionic compounds.
(d) Explain some uses of electrolysis.
(e) Explain the applications of electroplating.

Introduction

Electrolysis is the chemical process by which ionic compounds in molten or aqueous state get decomposed by passing electric current through them.

- An **electrolyte** is a compound which when in molten form or when dissolved in water (in aqueous form) can conduct electric current and get decomposed by it. Examples of electrolytes include sodium chloride and dilute sulphuric acid solutions.
- An **non-electrolyte** is a substance that does not conduct electric current when in molten or in solution form. Examples of non-electrolytes include sulphur, glucose, ethanol and distilled water.

Electrolytes conduct electric current in molten or aqueous form because they contain charged particles (ions) that are free and can move. Non-electrolytes on the other hand, do not conduct electric current because they exist as neutral molecules in solution and in molten form.

4.1 Electrolysis process

The process fo electrolysis is carried out in an electrochemical cell. The cell consists of several components as you will learn shortly.

Components of an electrolytic cell

An electrolytic cell usually consists of a dry cell, an ammeter, bulb, graphite rods and an electrolyte.

Dry cells are usually used in an electrolytic cell because they supply direct current (DC). Sometimes an ammeter can be used to measure the amount of current passing through as a bulb may not light if the voltage is low.

The graphite rods which carry the current into and out of the electrolytes are called **electrodes.** Graphite is chosen because it does not react with the electrolyte or with the

products formed during electrolysis. Unreactive metal electrodes such as platinum and copper can also be used as electrodes. The electrode connected to the positive terminal of the battery is called **anode**. The electrode connected to the negative terminal is called **cathode.** Fig. 4.1 shows the various components of an electrolytic cell.

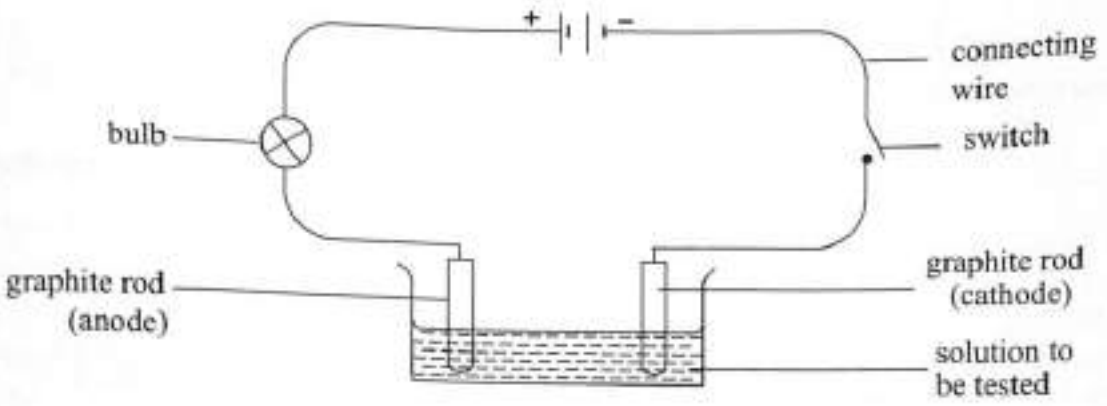

Fig. 4.1: Electrolytic cell

An electrolytic cell is used to pass an electric current through a molten or dissolved ionic compound. After a compound has been melted, heat must be supplied to keep the salt in molten state. Fig. 4.2 shows how ions occur in a binary compound before heating.

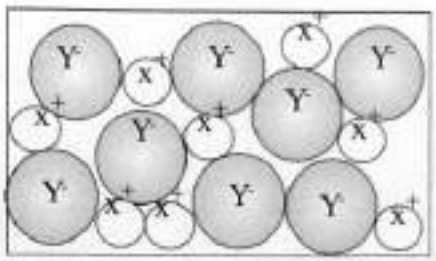

Fig. 4.2: Ions in a binary compound XY

When the binary compound for example X^+Y^- is heated, it melts and produces X^+ and Y^- ions as follows:

$$X^+Y^-(l) \xrightarrow{heat} X^+(l) + Y^-(l)$$

When the switch is closed to complete the circuit, all positive (+) ions migrate to the cathode and all negative (-) ions migrate to the anode. This generates **electric current**. As the current flows, a chemical reaction occurs at the electrodes. The ions lose the charge (they are discharged) i.e. Y^- ions lose electrons(s) while X^+ ions gain electron(s). The respective products are deposited on the surface of each electrode as shown in Fig. 4.3.

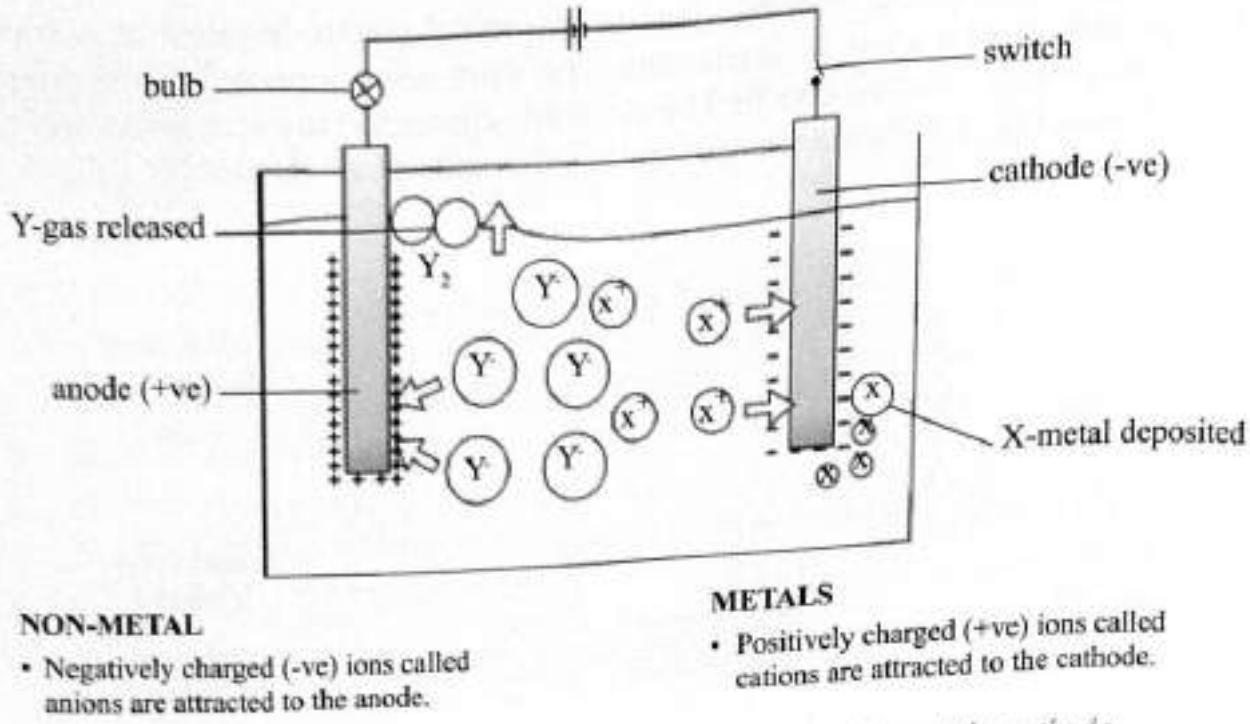

Fig. 4.3 Discharging of anions at the anode and cations at the cathode

We can write the reactions taking place at the electrodes in form of equations. The reactions at each of the electrodes is represented by half equation. When both half equations are combined, we obtain overall ionic equation as follows:

At the cathode

$$X^+(l) + e^- \longrightarrow X(s)$$

At the anode

$$Y^-(l) \longrightarrow Y(g) + e^-$$

$$X^+(l) + Y^-(l) \longrightarrow X(s) + Y(g)$$

The role of water in electrolysis

Water dissolves most ionic compounds. This is because its molecules are polar and therefore, attract the positive and negative ions that make the ionic compound. When the forces of attraction holding the ions in the solid are overcome, the solid dissolves.

$$H_2O_{(l)} \rightleftharpoons H^+(aq) + OH^-(aq)$$

The ions are set free and move at random. Water is a weak electrolyte because it is only partially ionised.

During electrolysis, the $H^+(aq)$ and $OH^-(aq)$ ions move to the electrodes, hence when an aqueous solution is electrolysed, apart from the ions of the salt, $OH^-(aq)$ ions and $H^+(aq)$ ions from the water will also compete to be oxidised or reduced at the electrodes. At each electrode, only one type of ions get discharged. This is called **preferential discharge**.

Factors affecting preferential discharge

We are already aware that during electrolysis, certain ions are preferentially discharged at their respective electrodes. There are factors which determine the ions to be discharged.

Research Activity

Find out from textbooks in the library or the internet which factors affect preferential discharge of ions during electrolysis.

1. *Position of the metal or ion in the electrochemical series*

The ease of discharge of ions during electrolysis depends on their position in the electrochemical series. Metals in this series are placed in an increasing order of their ability to gain electrons. The most reactive metals for example potassium release electrons easily to form ions. These cations require more energy to gain back electrons i.e. to be reduced and therefore, are more difficult to discharge. Ease of discharge therefore increases down the electrochemical series as shown in Table 4.1.

Table 4.1: The ease of discharge of ions during electrolysis of various solutions

Cations (+ve ions)		Anions (-ve ions)	
Ease of discharge of cations increases downwards	K^+ Na^+ Ca^{2+} Mg^{2+} Zn^{2+} Fe^{2+} Pb^{2+} H^+ Cu^{2+} Ag^+	SO_4^{2-} NO_3^- Cl^- Br^- I^- OH^-	Ease of discharge of anions increases downwards

2. *Concentration*

We have seen that ions lower in the electrochemical series will preferentially be discharged under normal conditions. However, if the ion that requires more energy to discharge is present in greater concentration than the other one lower in the electrochemical series, it will be discharged preferentially. For example, in electrolysis of dilute sodium chloride solution, OH^- ions are discharged preferentially to give oxygen since Cl^- ions are present in low concentrations. But, if the concentration of Cl^- ions is much higher than that of OH^- ions, the Cl^- ions are discharged first to give chlorine as seen in the electrolysis of brine.

Note: The number of Cl^- ions in the solution gradually decreases as their discharge continues. When they are no longer more than OH^- ions, the OH^- ions will be discharged preferentially to produce oxygen, because OH^- are lower in electrochemical series than Cl^- ions and are easier to discharge.

3. Nature (type) of electrode

The choice of electrodes during electrolysis can alter the order of discharge of the ions. For example, if a copper electrode(active electrode) is used as the anode during electrolysis of copper sulphate solution, neither OH^- nor SO_4^{2-} ions are discharged. Instead, the copper electrode dissolves. But if we change and use graphite (inert elecrode) as the anode, OH^- ions are discharged and we get oxygen gas. Another example where the type of electrode can change the order of discharge of ions is in the electrolysis of concentrated sodium chloride solution using a mercury cathode. Na^+ ions are discharged instead of H^+ ions, which should be the case if we consider the position of sodium in the electrochemical series. Sodium metal forms an amalgam with mercury in the process.

Test for conduction of electric current by aqueous solutions

Water is the most common solvent. A solution of a substance in water is known as **aqueous** solution. Experiment 4.1 below will be used to test whether substances dissolved in water conduct electric current.

Experiment 4.1

Aim: *To investigate the conductivity of aqueous solutions.*

Apparatus and reagents

- Battery/dry cells
- Bulb (6V) or ammeter
- Graphite rods
- Acidified water
- Connecting wires
- Beaker 250 cm^3
- Distilled water
- Aqueous solutions of: sugar, copper sulphate, sodium chloride and and sulphuric acid.

Procedure

1. Set up the apparatus as shown in Fig. 4.4. Before dipping the graphite rods into the solution, let the two electrodes touch and switch on the current to check whether the bulb lights up.
2. Place the graphite rods apart in the beaker.
3. Pour distilled water into the beaker. Switch on the current. Does the bulb light?
4. Switch off the current. Add a few drops of concentrated sulphuric acid to the water in the beaker. Switch on the current. Does the bulb now light?
5. What do you observe on the graphite rods? What does this indicate?
6. Repeat the above experiment using same graphite rods and the solutions of

Note: The number of Cl^- ions in the solution gradually decreases as their discharge continues. When they are no longer more than OH^- ions, the OH^- ions will be discharged preferentially to produce oxygen, because OH^- are lower in electrochemical series than Cl^- ions and are easier to discharge.

3. *Nature (type) of electrode*

The choice of electrodes during electrolysis can alter the order of discharge of the ions. For example, if a copper electrode(active electrode) is used as the anode during electrolysis of copper sulphate solution, neither OH^- nor SO_4^{2-} ions are discharged. Instead, the copper electrode dissolves. But if we change and use graphite (inert elecrode) as the anode, OH^- ions are discharged and we get oxygen gas. Another example where the type of electrode can change the order of discharge of ions is in the electrolysis of concentrated sodium chloride solution using a mercury cathode. Na^+ ions are discharged instead of H^+ ions, which should be the case if we consider the position of sodium in the electrochemical series. Sodium metal forms an amalgam with mercury in the process.

Test for conduction of electric current by aqueous solutions

Water is the most common solvent. A solution of a substance in water is known as **aqueous** solution. Experiment 4.1 below will be used to test whether substances dissolved in water conduct electric current.

Experiment 4.1

Aim: *To investigate the conductivity of aqueous solutions.*

Apparatus and reagents

- Battery/dry cells
- Bulb (6V) or ammeter
- Graphite rods
- Acidified water
- Connecting wires
- Beaker 250 cm^3
- Distilled water
- Aqueous solutions of: sugar, copper sulphate, sodium chloride and and sulphuric acid.

Procedure

1. Set up the apparatus as shown in Fig. 4.4. Before dipping the graphite rods into the solution, let the two electrodes touch and switch on the current to check whether the bulb lights up.
2. Place the graphite rods apart in the beaker.
3. Pour distilled water into the beaker. Switch on the current. Does the bulb light?
4. Switch off the current. Add a few drops of concentrated sulphuric acid to the water in the beaker. Switch on the current. Does the bulb now light?
5. What do you observe on the graphite rods? What does this indicate?
6. Repeat the above experiment using same graphite rods and the solutions of

the following compounds: sugar, copper sulphate, sodium chloride and dilute sulphuric acid.

7. Note the changes (if any) occurring around the graphite rods and in the contents of the beaker. Rinse the beaker and the rods thoroughly with distilled water before each new solution is used.

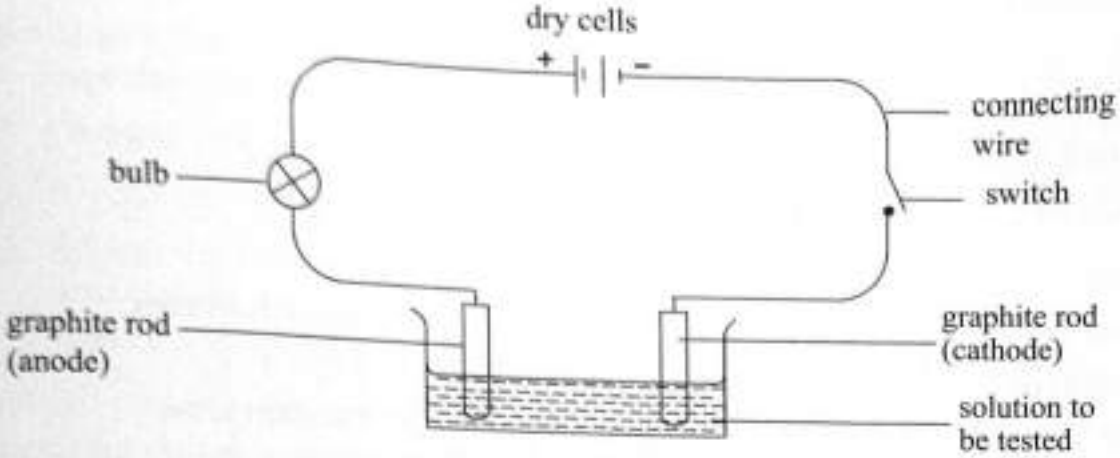

Fig. 4.4: The apparatus for testing conductivity of aqueous solutions

8. Copy Table 4.2 and fill your results, include all the other substances being investigated.

Table 4.2: Electrolytes and non-electrolytes

Solution	Does the bulb light		Observation at the electrodes	
	Yes	No		
			Anode	Cathode
Sugar				
Sodium chloride				
Concentrated sulphuric acid				
Sodium chloride				
Copper sulphate				
Pure water				
Acidified water				

Different results are achieved from the apparatus in Fig. 4.4 when used to test different solutions. Liquids such as acidified water, sodium chloride solution, mineral acids and copper sulphate solution conduct electric current. They are hence referred to as **electrolytes.** The current decomposes the electrolyte as it passes through it. On the

other hand, Aqueous solutions such as that of sugar does not conduct electric current. The are known as **non-electrolytes.**

Test for conduction of electric current by molten substances

You have seen in the Experiment 4.1 that when some compounds are dissolved in water they conduct electric current whereas others do not. In our next experiment, we will investigate whether some compounds and elements conduct electricity when melted.

Experiment 4.2

Aim: *To investigate the conductivity of molten substances.*

Apparatus and reagents

- Crucible
- Graphite rods
- Sulphur
- Connecting wire fitted with crocodile clips
- Battery/dry cells with cell holder
- Pipe-clay triangle
- Lead bromide or lead iodide
- Bunsen burner
- Tripod
- Paraffin wax
- Bulb

Cation: Lead bromide and lead iodide produce **poisonous** fumes. The experiment should be performed in a fume cupboard or in the open.

Procedure

1. Set up the apparatus as in Fig 4.5.

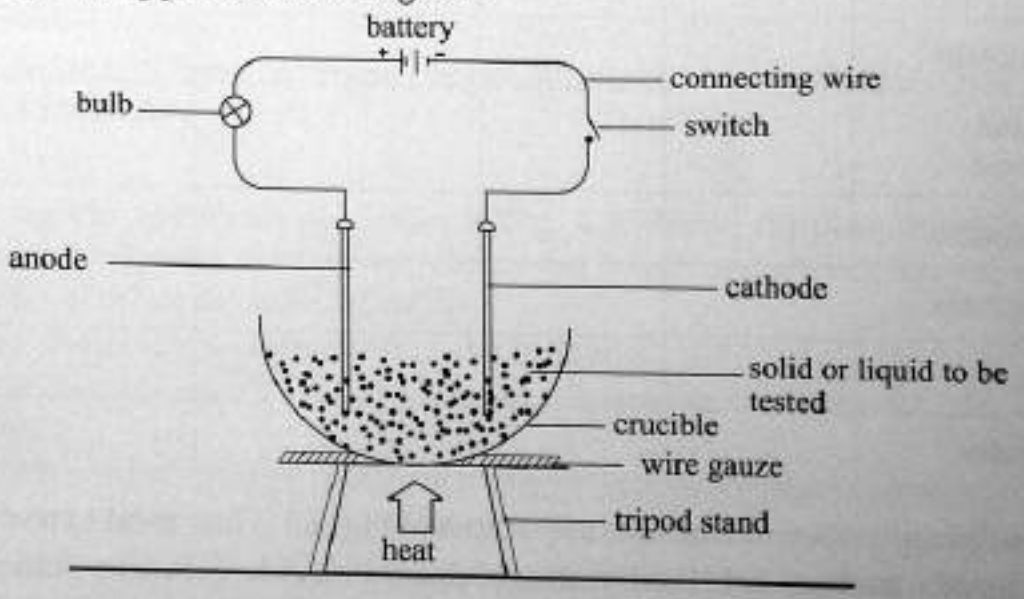

Fig 4.5: Testing conductivity of molten substances

2. Fill the crucible with either lead bromide or lead iodide to approximately two thirds.
3. Place the crucible and its contents on a pipe clay triangle supported by a tripod.
4. Insert the graphite electrodes into lead bromide and close the switch. Does the bulb light?
5. Open the switch and gently heat the crucible until lead bromide melts.
6. Adjust the burner so that the substance remains in molten state as the experiment progresses. Close the switch. Does the bulb light?
7. Observe whether there is any chemical reaction at the electrodes.
8. Record your observations as shown in Table 4.3.
9. Repeat the experiment with other substances such as paraffin wax, sulphur, lead iodide, sugar among others. Use a fresh crucible for each substance and clean the electrodes thoroughly before using it with the next substance.

Table 4.3 Electrolysis of molten substances

Substance	Does bulb light		Observation at the electrodes	
	Solid	Molten	Anode	Cathode
Lead bromide	No	Yes	Red brown fumes	Grey beads
Sulphur				
Wax				
Lead iodide				

When we insert the electrodes in solid lead bromide, the bulb does not light. But the bulb lights when lead bromide is heated until it melts. This indicates that molten lead bromide conducts electric current. It is an **electrolyte**. Metal compounds in molten state are electrolytes. It is also observed that other substances like sulphur, paraffin and wax do not conduct electric current when in solid or molten form. They are non-electrolytes. Other examples of electrolytes and non–electrolytes are given in Table 4.4.

Table 4.4: Other examples electrolytes and non-electrolytes

Electrolytes	Non-electrolytes
Sodium chloride solution	Pure water
Copper chloride solution	Sugar solution
Sodium hydroxide solution	Solid sodium chloride
Molten leadiodide	Ethanol (alcohol)
Hydrochloric acid	Urea
Ethanoic acid	
Most salts in molten or aqueous form	

Self assessment exercise 4.1

1. Explain why solid sodium chloride does not conduct electricity but aqueous or molten sodium chloride does.
2. Explain the difference in conductivity of sodium metal and solid sodium chloride.
3. What is meant by the following terms:
   (i) Electrolyte
   (ii) Non-electrolyte
   (iii) Electrolysis
   (iv) Anode
   (v) Cathode

4.2 Electrolysis of acqueous and molten compounds

In this section, you will learn what happens when various compound are subjected to the process of electrolysis.

(a) Electrolysis of dilute sulphuric acid

Electrolysis of dilute sulphuric acid is also known as electrolysis of water. By itself, water is a very poor conductor of electric current because it is a weak electrolyte. We therefore add an electrolyte to water to provide more ions that flow through the solution thereby completing the electric circuit. Sulphuric acid is a strong electrolyte and when added to water, it provides more mobile ions.

Ions present

- From sulphuric acid – $H^+(aq)$; $SO_4^{2-}(aq)$
- From water – $H^+(aq)$, $OH^-(aq)$

Reactions at electrodes

At the cathode	At the anode
Ions present: $H^+(aq)$ H^+ ions move to the cathode. H^+ ion gain an electron to form H atom i.e. $H^+(aq) + e \longrightarrow H(g)$. Since hydrogen is a diatomic gas, two atoms combine to form hydrogen (H_2) molecule. The $4e^-$ from the anode are taken up by H^+ ions giving 2 moles of hydrogen gas. Overall ionic equation: $4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$	**Ions present:** $SO_4^{2-}(aq)$, $OH^-(aq)$ Both ions move to the anode. OH^- ion is lower in the electrochemical series and is therefore discharged in preference to SO_4^{2-} i.e. $OH^-(aq) \longrightarrow OH(g) + e^-$ $4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$

Note: In the electrolysis of sulphuric acid solution, $OH^-(aq)$ and $H^+(aq)$ are discharged. The solution becomes more concentrated as water is continually decomposed. This is why it is sometimes known as electrolysis of water.

Self assessment exercise 4.2

1. During electrolysis of dilute sulphuric acid, at which electrode does oxidation take place?
2. In the electrolysis of water, at which electrode will hydrogen be produced? Give a reason for your answer.

Electrolysis of molten lead bromide

Remember: Perform this experiment in a fume cupboard or an open place. Lead is poisonous hence can pollute the environment.

Lead bromide is an ionic compound. The electrolyte consists of two types of ions - lead(II) ions and bromide ions.

$$\text{Lead bromide} \xrightarrow{\text{heat}} \text{Lead(II) ions} + \text{Bromide ions}$$

$$PbBr_2(s) \longrightarrow Pb^{2+}(l) + 2Br^-(l)$$

Lead bromide is a binary salt. In molten state, it produces two types of ions to be considered for discharge. It is easy to predict the products at the electrodes.

Experiment 4.3

Aim: *To investigate electrolysis of molten lead bromide*

Apparatus and reagents

- A crucible
- Carbon electrodes
- Bulb
- Dry cells
- Source of heat
- Tripod stand, wire gauze, Lead bromide

Procedure

1. Set up the apparatus as shown in Fig. 4.6.
2. Fill a crucible to about two-thirds with lead bromide.

3. Heat the salt until it melts. Keep on supplying heat to keep the salt in molten state.
4. Observe the bulb to see whether it lights.
5. Let the current pass for about 10 minutes.
6. Pay particular attention to any chemical changes taking place around the electrodes. Record any observation made.
7. Remove the electrodes and examine them carefully. Record any observation.
8. After the crucible has cooled, observe its contents. What do you observe around each electrode? Write equations to explain your observations.

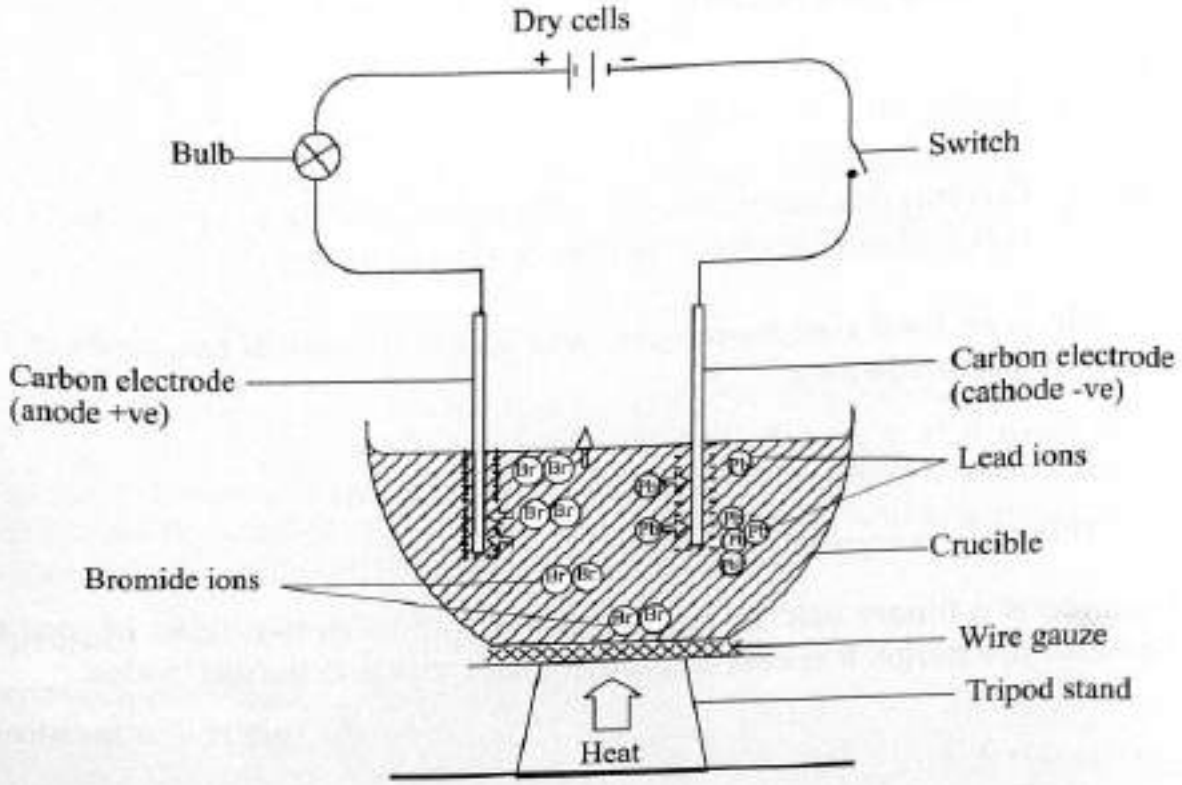

Fig. 4.6: Movement of ions during the electrolysis of molten lead bromide.

The bulb lights brightly because molten lead bromide conducts electric current i.e. it is an electrolyte. Brown vapours are observed at the anode. These are bromine fumes produced during electrolysis. Beads of a grey solids are deposited at the cathode and at the bottom of the crucible. This is lead metal.

We can conclude that as the electric current passes through molten lead bromide, it decomposes to form lead metal and bromine.

Ions present: Br^-, Pb^{2+}

Reactions at electrodes

At the anode	At the cathode
• $Br^-(aq)$ ions move to the anode and are discharged by losing an electron and becoming a bromine atom. $Br^-(l) \rightarrow Br(g) + e^-$ • Bromine atoms combine in pairs to form molecules. $Br(g) + Br(g) \rightarrow Br_2(l)$ It is common to write one equation once we understand that there are very many Br^- ions in the molten solution. $2Br^-(l) \rightarrow Br_2(l) + 2e^-$	• $Pb^{2+}(aq)$ ions move to the cathode and get discharged by gaining electrons to become lead atoms i.e. grey lead metal is formed. The metal is deposited on the cathode. $Pb^{2+}(l) + 2e^- \rightarrow Pb(s)$

Self assessment exercise 4.3

1. What products and at which electrode would they be formed if molten copper chloride was electrolysed?
2. Solid sodium chloride does not conduct electricity but in molten or solution form it does. Explain.

(c) Electrolysis of dilute sodium iodide

Experiment 4.4

Aim: *To investigate electrolysis of dilute sodium iodide solution.*

Apparatus and reagents

- Beaker
- Carbon (graphite) rods
- Wooden splint and a match box
- Connecting wires fitted with crocodile clips
- Dry cells (1.5V) with cell holder/ battery
- Stirrer
- Bulb/ ammeter
- Test tube rack with test tubes

- Distilled water
- Sodium iodide

Procedure

1. Set up the apparatus as shown in Fig 4.7 below.

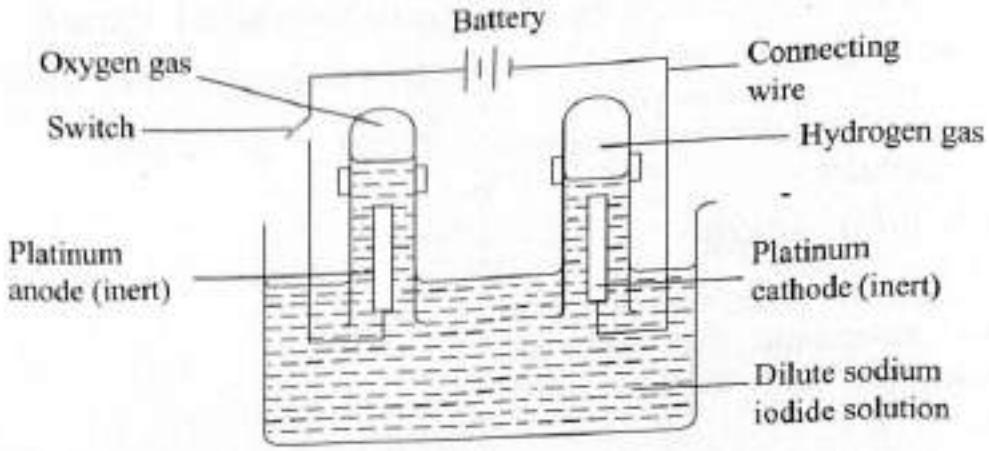

Fig. 4.7: Electrolysis of dilute sodium iodide solution

2. Half-fill the beaker with a dilute solution of sodium iodide.
3. Switch on the current. What do you observe on the electrodes?
4. What is the colour of the gas collected at the
   (a) anode?
   (b) cathode?
5. Test the gases collected with glowing and burning splints.
6. Copy Table 4.5 in your notebook and record your observations.

Table 4.5: Electrolysis of dilute sodium iodide solution

Gas collected at	Colour	Test with glowing/burning splint
Anode		
Cathode		

A colourless gas, which re-lights a glowing splint is collected at the anode. The gas is **oxygen**. At the cathode, a colourless gas is collected which burns with a 'pop' sound. This gas is hydrogen gas.

Now, let us see how these gases are formed at the respective electrodes. We start by finding out which ions are present in the electrolyte, then see how they migrate and finally find out which ones are selectively discharged.

Ions present

- From sodium iodide $Na^+(aq)$, $I^-(aq)$
- From water $H^+(aq)$, $OH^-(aq)$

The positive ions, i.e. Na^+ and H^+ migrate to the cathode. The negative ions, I^- and OH^-, migrate to the anode. This is because they are attracted by the oppositely charged electrodes.

Reactions at the electrodes

At the cathode	At the anode
Ions present: $Na^+(aq)$, $H^+(aq)$ Both ions move to the cathode. H^+ ion is discharged in preference to Na^+ ion because it is lower than Na^+ in the electrochemical series. It gains an electron to form H atom. $H^+(aq) + e^- \longrightarrow H(g)$ Two hydrogen atoms combine to form H_2 (g) molecule. $H(g) + H(g) \longrightarrow H_2(g)$ Therefore; $2H^+(aq) + 2e^- \longrightarrow H_2(g)$ But the $4e^-$s produced at the anode must react hence overall equation is $4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$	**Ions present:** $I^-(aq)$, $OH^-(aq)$ Both ions move to the anode. OH^- ion is discharged because it is lower then I^- in the electrochemical series. OH^- is discharged by losing an electron. $OH^-(aq) \longrightarrow OH + e^-(g)$ Four OH^- ions are discharged, the OH^- groups interact to form one mole of oxygen and two moles of water as shown below, releasing 4 electrons which move to the cathode. Hence overall reaction is given by $4OH^-(aq) \longrightarrow O_2(g) + 2H_2O(l) + 4e^-$

Note: The $4e^-$s released at the anode must be taken up by H^+ ions. So we must multiply the last equation under cathode by 2 to get the overall equation.

Finally, Na^+ and I ions remain in solution forming sodium iodide; a solution of higher concentration but the pH does not change.

(d) Electrolysis of brine

Experiment 4.5

Aim: *To investigate electrolysis of brine (concentrated sodium chloride solution).*

Apparatus and reagents

- The apparatus is the same as in Experiment 4.4 and the set up is as in Fig. 4.7 or glass cell fitted with inert electrodes such as platinum or carbon as in Fig 4.8.
- 1M sodium chloride solution.

Procedure

1. Set up the apparatus as shown in Fig. 4.8 below.

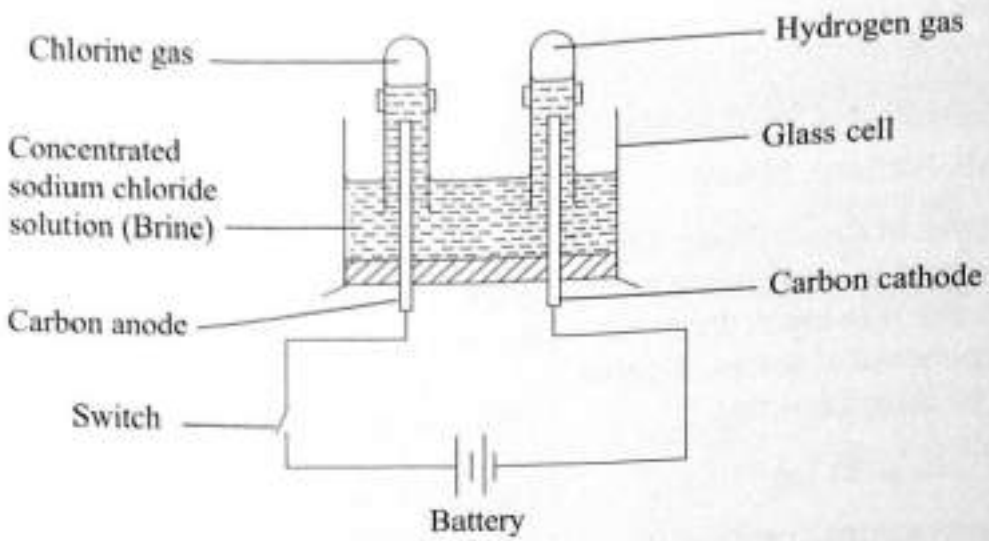

Fig. 4.8: Set-up for electrolysis of brine

2. Half-fill the beaker or glass cell with 1M sodium chloride solution.
3. Switch on the current.
4. Note the colour of the gas collected at the anode and at the cathode.
5. At which electrode is a gas with a smell produced?
6. Test the gas produced at each electrode with a burning splint. At which electrode does the gas burn with a "pop" sound?

When the circuit is completed, bubbles are observed at the electrodes. A green-yellow poisonous gas which has a choking irritating smell is collected at the anode. This gas bleaches moist litmus papers and therefore the gas is **chlorine**. At the cathode, a colourless gas which burns with 'pop' sound is collected. This gas is hydrogen.

Ions present

- From sodium chloride: $Na^+(aq)$, $Cl^-(aq)$
- From water: $H^+(aq)$, $OH^-(aq)$

Reactions at the electrodes

At the cathode	At the anode
Ions present $Na^+(aq)$, H^+ H^+ ions are discharged in preference to Na^+ ions because H^+ ion is lower in electrochemical series than Na^+ ion. $2H^+(aq) + 2e^- \longrightarrow H_2(g)$	**Ions present** $Cl^-(aq)$, $OH^-(aq)$ Although OH^- ion is lower in the electrochemical series than Cl^-, the concentration of Cl^- ions at the anode is greater than the OH^- ions. Therefore, Cl^- ions are discharged in preference to OH^- ions. $2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$ This leads to production of chlorine gas at the anode.

Note

(a) The **concentration** in the above reaction has determined the product at the anode.

(b) H^+ and Cl^- ions are discharged leaving Na^+ and OH^- ions. Finally, sodium chloride solution is converted to sodium hydroxide solution.

- What would be formed if the chlorine produced was allowed to react with sodium hydroxide solution?
- Write an ionic equation for the reaction.

Chlorine would react with sodium hydroxide solution to form sodium hypochlorite which is a bleaching agent, and sodium chloride.

Ionic equation

$$Cl_2\,(g) + 2OH^-(aq) \longrightarrow Cl^-\,(aq) + OCl^-\,(aq) + H_2O\,(l)$$

Self assessment exercise 4.4

1. Write the equations that occur at the electrodes during the electrolysis of dilute and concentrated sodium chloride solutions.

(e) Electrolysis of copper sulphate solution using inert electrodes

Inert electrodes are electrodes which allow passage of electric current into and out of electrolyte without themselves getting decomposed. Carbon (graphite) and platinum are examples of inert electrodes.

Experiment 4.6

Aim: *To conduct electrolysis of copper sulphate using carbon electrodes.*

Apparatus and reagents

- Battery/dry cells
- Carbon rods
- Connecting wires fitted with crocodile clips
- Beaker
- Bulb (6V) or ammeter
- Copper sulphate solution

Procedure

1. Half-fill a beaker with copper sulphate solution.

2. Set up the apparatus as shown in Fig. 4.9 below.

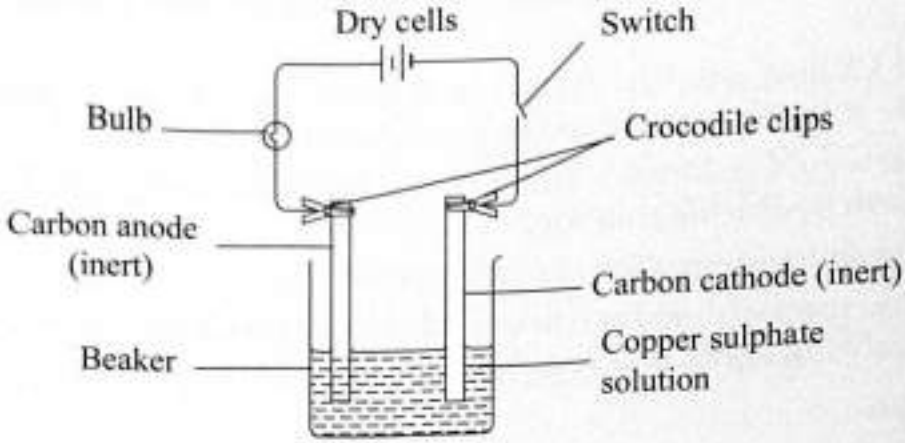

Fig. 4.9: Electrolysis of copper sulphate using carbon electrodes

3. Switch on the current and after a short while, observe what happens on the carbon electrodes at the:
   (i) anode
   (ii) cathode
4. Record your observations in your notebook.
5. What do you conclude?

Ions present

- From copper sulphate – $Cu^{2+}(aq)$, $SO_4^{2-}(aq)$
- From water – $H^+(aq)$, $OH^-(aq)$

At the cathode	At the anode
$Cu^{2+}(aq)$, $H^+(aq)$ move to the cathode.	$SO_4^{2-}(aq)$, $OH^-(aq)$ both move to the anode.
Cu^{2+} ions gain electrons to form copper which appears as a brown deposit at the cathode i.e. Cu^{2+} ions are preferentially discharged.	OH^- ions are preferentially discharged lose electrons being lower in the electrochemical series.
Ionic equation:	They combine to form water and oxygen.
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	Ionic equation:
Account for the $4e^-$s from the anode.	$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$

Copper (Cu^{2+}) ions, give copper sulphate solution the blue colour. As they are discharged and deposited at the cathode as copper solid, the colour of the solution fades and if we continue with the electrolysis, the solution finally turns colourless and acidic. The solution turns acidic because $H^+(aq)$ and $SO_4^{2-}(aq)$are unaffected during this process. They accumulate in the solution.

Note: Further electrolysis of the resultant colourless solution would give the same products as electrolysis of dilute sulphuric acid.

(f) Electrolysis of copper sulphate solution using copper electrodes

Experiment 4.7

Aim: *To investigate electrolysis of copper sulphate solution using copper electrodes*

Apparatus and reagents

- As in experiment 4.6 but using copper electrodes
- Electric balance (sensitive)
- Distilled water
- Propanone

Procedure

1. Put copper sulphate solution in a beaker then set-up the apparatus as shown in Fig. 4.10 below.

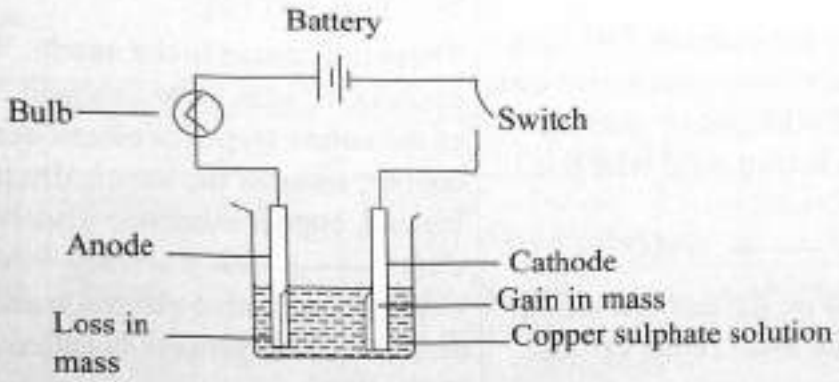

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Fig. 4.10: Electrolysis of copper sulphate using copper electrodes

2. Clean the copper electrodes until they shine.
3. Weigh the electrodes and record the mass in your notebook as in Table 4.7.
4. Switch on the electric current.
5. After 10 minutes, note the colour of the electrolyte and rinse the electrodes with distilled water.
6. Dip the electrodes in propanone and then let the propanone dry.
7. When dry, re-weigh the electrodes.
8. Record the mass after the experiment as shown in Table 4.7.
   - Why must electrodes be cleaned before electrolysis?
   - Why must the electrodes be washed with distilled water and then propanone after electrolysis?

Table 4.6: Results of electrolysis of copper sulphate solution using copper electrodes

	Before passing electricity	After passing electricity
Colour of copper sulphate solution		
Mass of electrodes		
Change in mass of electrodes		

The electrodes must be cleaned to remove any oxide that might be on the surface. The electrodes are washed with water and propanone to ensure they are completely dry.

Ions present:

- From copper sulphate – $Cu^{2+}(aq)$, SO_4^{2-} (aq)
- From water – $H^+(aq)$, OH^- (aq)

At the cathode	At the anode
Ions present $Cu^{2+}(aq)$, $H^+(aq)$ These ions move to the cathode. Cu^{2+} ions are lower in the electrochemical series and are preferentially discharged by gaining electrons to form a brown solid which is copper. $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$ Copper is deposited on the cathode and it causes an increase in mass of the cathode	**Ions present** SO_4^{2-} (aq), OH^- (aq) These ions move to the anode. We would expect OH^- ions to be oxidised. But because of the nature (type) of electrodes used i.e copper; **none** of the ions is discharged. Instead, copper electrode dissolves i.e $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$ Copper is an active electrode and it therefore dissolves. This process is called **electrode ionisation**.

Note: Copper (Cu^{2+}) ions make copper sulphate solution blue. For each ion removed at the cathode, it is replaced when copper dissolves at the anode. Therefore, the solution remains blue and the concentration of the solution remains constant.

When we electrolyse copper sulphate solution using copper electrodes, the mass of both electrodes change. The anode loses mass while the cathode gains mass. The loss in mass at the anode is equal to the mass gained at the cathode (see Fig. 4.11). Remember that metals are always formed at the cathode. This is the reason why even if we use carbon (graphite) as the cathode, the metal will still be deposited on it.

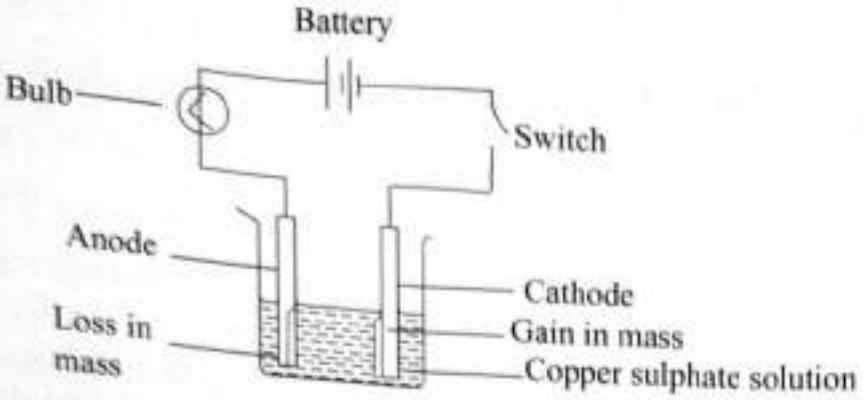


Fig. 4.11: Loss in mass at the anode is equal to gain in mass at the cathode

Table 4.7: Summary of electrolysis of some aqueous solutions

Electrolyte	Electrodes	Reaction of the anode	Reaction at the cathode
Dilute sulphuric acid (water)	Graphite/ platinum	$4OH^-(aq) \longrightarrow 2H_2O(l)+O_2(g)+4e^-$	$2H^+(aq)+2e^- \longrightarrow H_2(g)$
Dilute sodium chloride	Graphite/platinium	$4OH^-(aq) \longrightarrow 2H_2O(l)+O_2(g)+4e$	$2H^+(aq)+2e^- \longrightarrow H_2(g)$
Concentrated sodium chloride (brine)	Graphite/platinium	$2Cl^-(aq) \longrightarrow Cl_2(g)+2e^-$	$2H^+(aq)+2e^- \longrightarrow H_2(g)$
Magnesium sulphate	Graphite/platinium	$4OH^-(aq) \longrightarrow 2H_2O(l)+O_2(g)+4e^-$	$2H^+(aq)+2e^- \longrightarrow H_2(g)$
Copper sulphate	Carbon/platinium	$4OH^-(aq) \longrightarrow 2H_2O(l)+O_2(g)+4e^-$	$Cu^{2+}(aq)+2e^- \longrightarrow Cu(s)$
Copper sulphate	Copper	Copper dissolves $Cu(s) \longrightarrow Cu^{2+}(aq)+2e^-$	Copper deposited $Cu^{2+}(aq)+2e^- \longrightarrow Cu(s)$

4.3 Uses of electrolysis

Electrolysis has many important applications. These include:

- Extraction of metals
- Refining of metals
- Electroplating of metals

(a) Extraction of metals

Electrolysis is used to extract highly reactive metals such as sodium, potassium, calcium, magnesium and aluminium from their molten compounds. Because of their reactivity, these elements occur in combined states.

(i) Extraction process of sodium

Sodium is extracted by electrolysis of molten sodium chloride in a Down's cell; hence the process is commonly referred to as the **Down's process.**

Down's process

This takes place in a Down's cell (see Fig. 4.12 below).

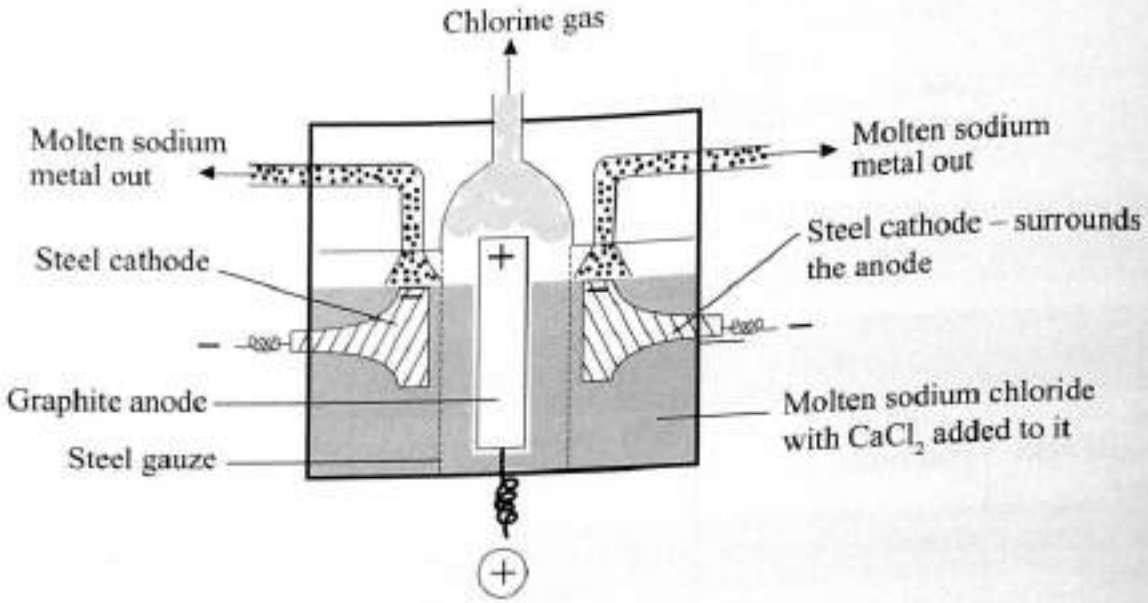


Fig. 4.12: Down's cell for the extraction of sodium

The cell consist of:

- an iron box lined with fire-bricks.This iron box contains the mixture of salts to be electrolysed.
- a carbon (graphite) anode.
- a circular steel cathode from which metallic sodium rises into the receiver.
- steel gauze for separating sodium and chlorine as these can react to form sodium chloride again.

Sodium chloride is heated to temperatures of upto 800°C in the cell in order to melt it. This high temperatures is expensive to attain and maintain. At this temperatures fused salt is also very corrosive. Therefore, calcium chloride ($CaCl_2$) is added to sodium chloride to bring its melting point to about 600°C. At this temperature; the reaction proceeds satisfactorily.

In molten state, sodium chloride occurs in ionic form i.e.

$$NaCl(s) \longrightarrow Na^+(l) + Cl^-(l)$$

Na^+ ions migrate to the cathode whereas Cl^- ions migrates to the anode where they are discharged at their respective electrodes thus obtaining sodium metal and chlorine gas.

Table 4.8: Reactions at the electrodes in Down's cell

At the cathode	At the anode
• The product obtained in sodium metal • $2Na^+(l) + 2e^- \longrightarrow 2Na(l)$ Sodium has a relatively low melting point, thus under a very hot environment, sodium is formed as a liquid. This liquid sodium is collected, solidified and can be stored for use. It floats on the brine due to its low density from where it is tapped off.	• The product obtained is chlorine gas • $2Cl^-(l) \longrightarrow Cl_2(g) + 2e^-$ Chlorine gas reacts vigorously with sodium to form sodium chloride. Thus, it is important to keep the chlorine gas away from the sodium in the cell. This is done using the steel as shown in Fig.4.12.

(b) Extraction process of aluminium

Aluminium is extracted from its chief ore **bauxite.** Extraction involves concentration of the ore and electrolysis.

Electrolysis of aluminium oxide

The aluminium oxide is first melted and then electrolysed. When the aluminium oxide is in molten state,the Al^{3+} and O^{2-} ions are free to move.

However, aluminium oxide has a very high melting point (2050°C) and also does not dissolve in water which would otherwise be used as a solvent to separate the ions. To solve these problems during the process of electrolysis , the purified aluminium oxide is dissolved in molten cryolite (Sodium aluminium flouride, Na_3AlF_6). This produces a suitable electrolyte at a lower temperature, about 900°C -1000°C. The cryolite does not decompose during the process. Electrolysis of aluminium oxide takes place in the Hall's cell. (See Fig 4.13 below)

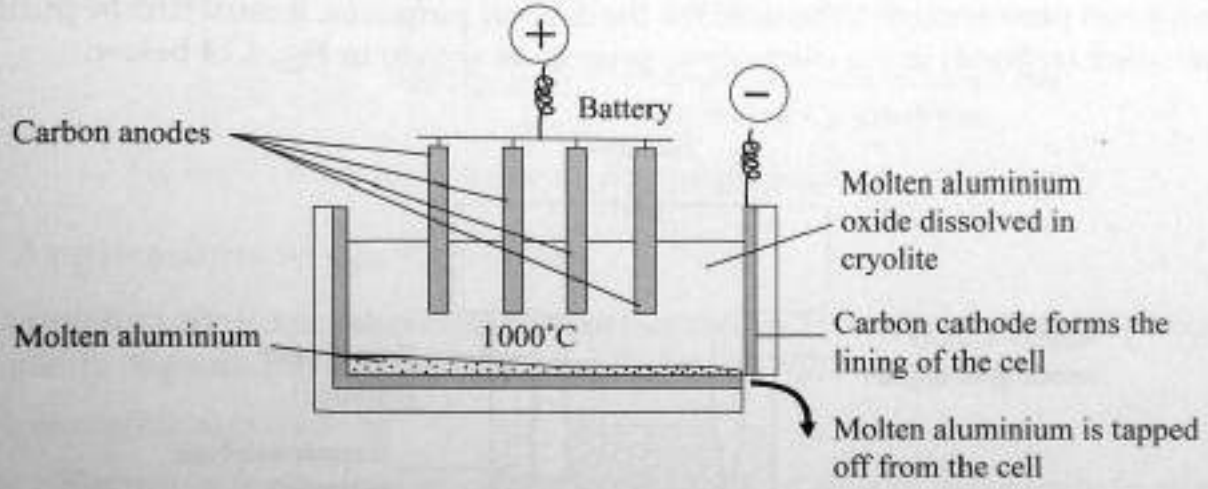


Fig. 4.13: Hall's cell for electrolysis of aluminium oxide

The electrolysis of aluminium oxide starts as soon as it melts. The aluminium ions , (Al^{3+})move to the cathode, whereas the oxide ions (O^{2-}) move to the anode. The aluminium(Al^{3+}) ion gain three electrons forming an aluminium atom. The oxide ion lose two electrons to form oxygen molecule (O_2).

Reaction at the electrode

At the cathode	At the anode
$Al^{3+}(l) + 3e^- \longrightarrow Al(l)$ To balance the number of electrons at the cathode and anode multiply the equation above by 4: $4Al^{3+}(l) + 12e^- \longrightarrow 4Al(l)$	$2O^{2-}(l) \longrightarrow O_2(g) + 4e^-$ To balance the number of electrons at the cathode and anode, multiply the equation above by 3. $6O^{2-}(l) \longrightarrow 3O_2(g) + 12e^-$
Overall equation $4Al^{3+}(l) + 6O^{2-}(l) \longrightarrow 4Al(l) + 3O_2(g)$	

Oxygen forms at the anode and metallic aluminium settles at the bottom of the cell and it is tapped off at intervals. Fresh aluminium oxide is continually added as the reaction proceeds so that the process becomes continuous. The cryolite is not used up during the process.

Note: The oxygen gas produced at the anode reacts with the carbon anodes to produce carbon dioxide or carbon monoxide. This reaction burns away the anodes. Therefore, the anode must be replaced at intervals.

This process requires high input of electrical energy and therefore, it is only economical in countries where electricity is cheap and readily available.

2. Refining of metals

Copper is an excellent conductor of electricity. It is therefore used widely as a conductor in various electrical appliances and power transmission lines. Copper extracted from its ores is not pure enough to be used for the desired purposes. It must first be purified. It is purified (refined) using electrolysis process as shown in Fig. 4.14 below.

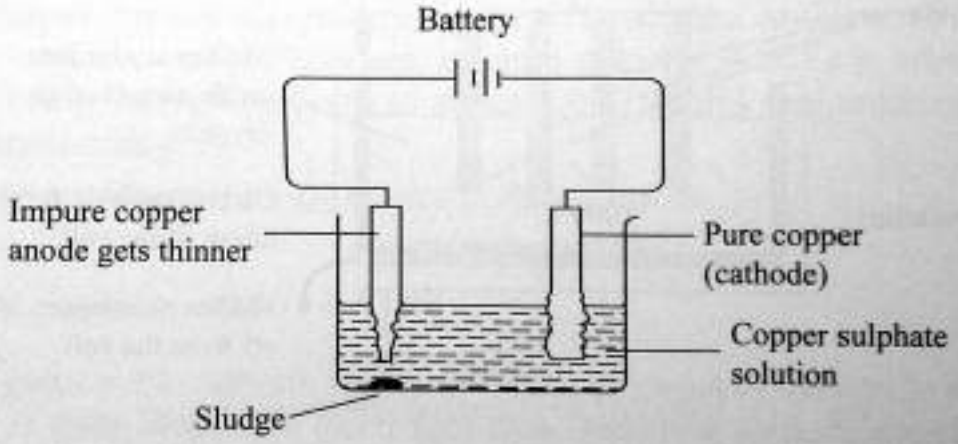


Fig 4.14: Refining impure copper

During refining of copper, the following must be done:

- the cathode must be made of thin sheet of pure copper.

- the anode must be made of impure copper.
- the electrolytic solution must contain copper ions e.g. copper sulphate solution.

(a) At the anode

The copper atoms lose electrons and enter into solution as ions.

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

(b) At the cathode

The positively charged (Cu^{2+}) ions are attracted to the cathode and get discharged by accepting two electrons. They then get deposited on the cathode as pure copper solid.

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

Thus, the impure block at the anode dissolves and pure copper is deposited at the cathode. The impurities from the anode drop to the bottom of the cell as **sludge**. Note: The sludge contains valuable metals like silver and gold which can be extracted.

3. Electroplating of metals

Electroplating is a process whereby a thin layer of a metal is applied on the surface of a material of interest through electrolysis process. As learnt earlier, electroplating is vastly done on stainless steel to prevent rusting. It is also done on some objects to make them look shiny and attractive.

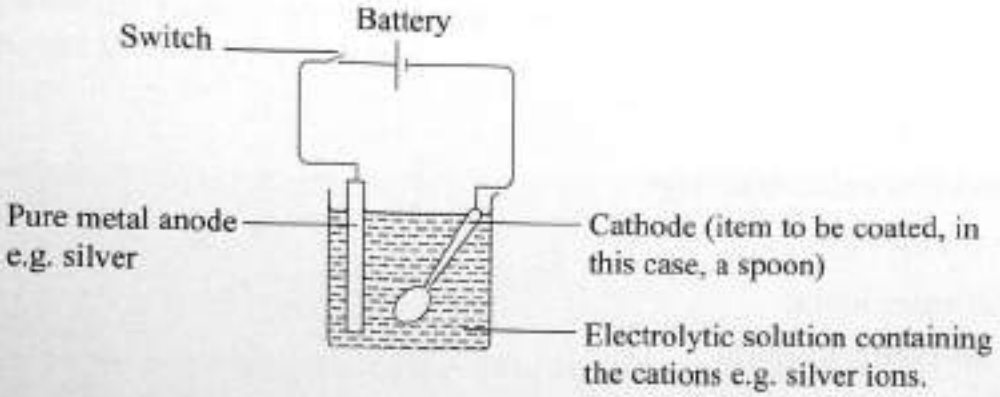


Fig. 4.15: Conventional electrolytic cell that can be used for electroplating

4.4 Applications of electroplating

Electroplating offers a number of important features and benefits that can help improve the quality of products. The process is applied in the following areas:

1. *Automotive applications*

- Electroplating is done on old bumpers and tyre rims to make them look new.
- The process is also applied on light weight parts made of plastics.
- Palladium plating is used in the manufacture of catalytic converters due to its ability to absorb excess hydrogen.

2. *Corrosion protection*

- Nickel plating, tin plating and their various alloys are all used to offer protection on nuts and bolts.
- Gold plating also provides superior protection against tarnishing, but it is more expensive than other plating processes. Plated products last longer.

3. *Electrical components*
   - Silver electroplating is often used on copper or brass connectors due to its superior conductivity. It is also cost effective.
4. *Manufacture of computer hard drives*
   - Nickel plating is used in making of hard drives so that the discs can easily be read.
5. *Zinc-nickel plating in the manufacture of engine parts*

   Zinc-nickel plating is capable of withstanding extremely high temperatures. Plating with these metals protect engine parts from damage due to extreme temperatures.
6. *Jewelry industry*

   Jewelry is often plated with a thin layer of a precious metal such as gold and silver to make it more lustrous and attractive.
7. *Electroplating*

   Electroplating is sometimes used to make brittle materials stronger and more durable. They become less susceptible to damage when struck. This increases their lifespan.

Experiment 4.8

Aim: *To demonstrate electroplating*

Apparatus and reagents

- Impure copper metal
- Copper sulphate solution
- D.C. battery
- Connecting wire
- Steel spoon
- Beaker
- Switch

Procedure

1. Clean the copper metal thoroughly.
2. Make it the anode.
3. Clean the spoon to be electroplated thoroughly and make it the cathode.
4. Connect the wire to the battery with the switch.
5. Connect one end of the wire to the anode and the other end of the cathode.

6. Pour copper sulphate solution into the beaker until both the electrodes are immersed in the solution
7. Connect the switch.
8. Set up the apparatus as in Fig. 4.16 then wait for sometime. Record your observations.

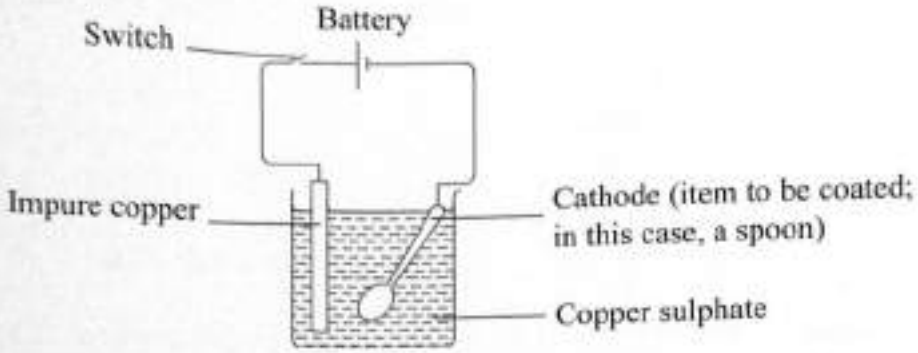


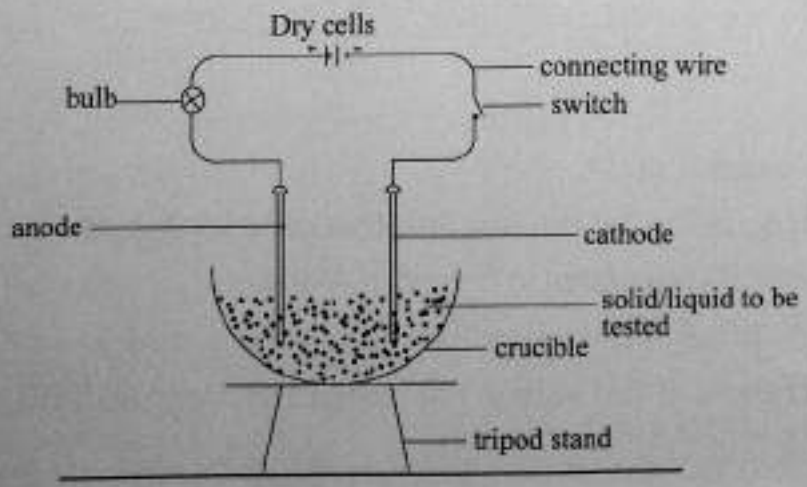
Fig. 4.16: Electroplating a spoon

Self assessment exercise 4.5

1. Give two applications of electrolysis.
2. A student wanted to coat a steel knife with chromium. On which electrode should the knife be attached.

Revision exercise 4

1. In an experiment to investigate conductivity of substances, a student used the set-up shown below.

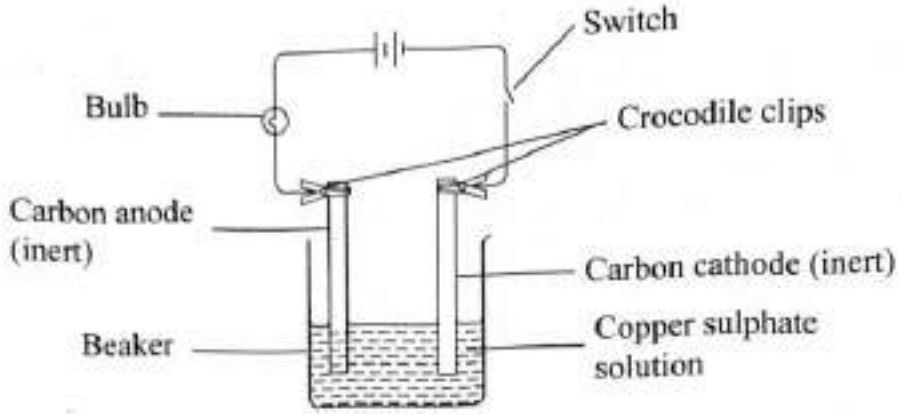


The students noticed that the bulb did not light.

a) What had been omitted in the set-up?

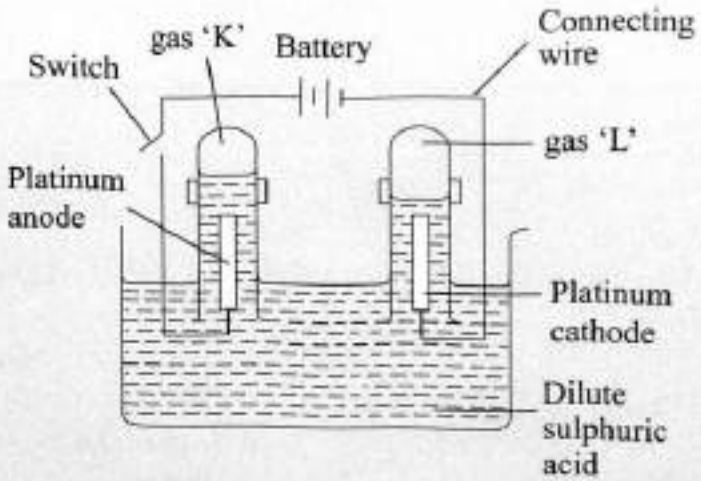
b) Explain why the bulb would light when the omission is corrected?

2. The set-up below was used to electrolyse aqueous copper sulphate.



Write the half equation of the reaction that took place at the cathode.

3. 100 cm^3 of 2M sulphuric acid was electrolysed using the set-up shown below.



i) Identify gases K and L.

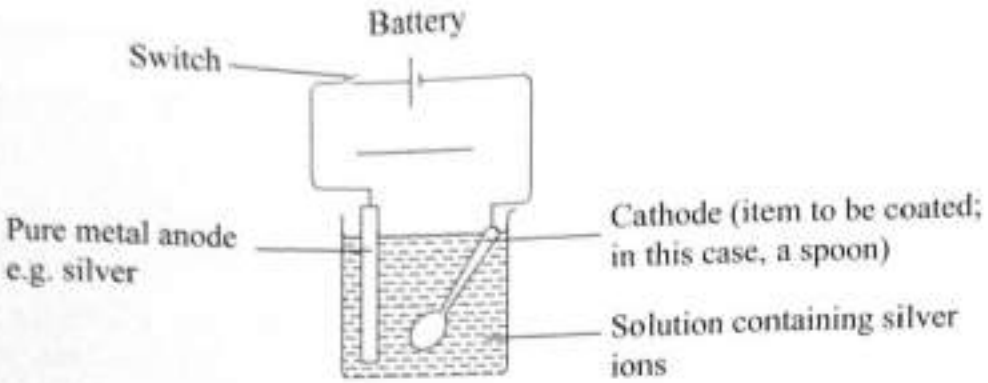
ii) Write an equation for the reaction that produced gas L.

iii) Explain the difference in:

a) the volume of the gas produced at the electrodes.

b) brightness of the bulb if 100 cm^3 of 2M acetic acid was used in place of 100 cm^3 2M sulphuric acid.

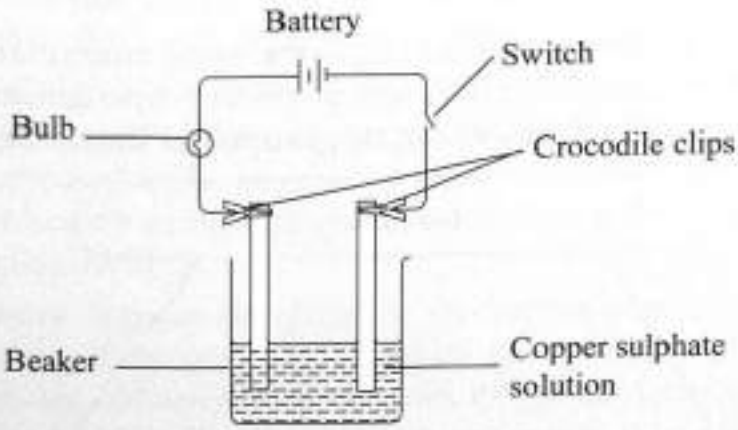
4. The set-up below was used to electroplate a metal spoon. Study it and answer the questions that follow.



a) Write the half equation for the reaction that occurred at the cathode.

b) State and explain what happens at the anode.

5. The set-up below was used during the electrolysis of aqueous copper sulphate using inert electrodes.



a) Label the electrodes in the set-up above.

b) Identify the anions and cations in the solution.

c) Explain the change that occurred to the concentration of copper sulphate solution during the experiment

6. One of the uses of electrolysis is electroplating

(a) What is electroplating?

(b) How can we make use of electroplating?

7. Name the particles responsible for the electrical conduction in

(a) copper metal.

(b) molten lead iodide.

(c) sodium chloride solution.

Topic 5 Isomerism

Success criteria

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

- Define isomers
- Draw the structures of isomers of alkanes, alkenes, alkanols, alkanals, alkanoates and alkanoic acids.
- Describe the effect of branching on physical properties in different organic families.

Introduction

Isomerism refers to the occurrence of two or more compounds with the same molecular formula but different structural formulae. It is derived from the word **isomer**.

Isomers

Isomers are two or more compounds that have the same molecular formula but with different structures. Isomers exbihit different physical properties however, they share the same chemical properties. Isomers can be grouped as shown below.

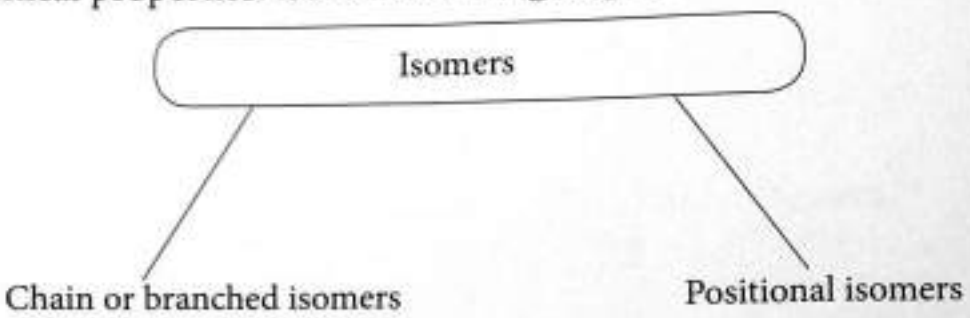


Fig. 5.1 Types of isomers

(a) Chain or branched isomers

In branched isomers, the position of the substituent groups for example substituent atoms or alkyl groups is shifted to different carbon atoms but the position of the functional group such as the double bond or the alcohol group remains the same.

(b) Positional isomers

In positional isomers, the position of the functional group for example double bond, or triple bond or alcohol group can be shifted to a different position within the molecule.

Isomerism occurs in various families of organic compounds. Such organic families are alkanes, alkenes, alkanols, alkanones and alkanoic acids.

When naming isomers, the International Union of Pure and Applied Chemistry (IUPAC) naming system is applied.

5.1 Isomerism in alkanes

Alkanes exhibit chain or branched isomerism. This is where the carbon chain is branched or the methyl group shifts its position in the chain.

Systematic (IUPAC) naming of branched alkanes

The following are rules used in naming branched alkanes:

1. Determine the **longest continuous chain** of carbon atoms. This chain does not have to be a straight line. It can be zigzag. This will give us the parent name.
2. Identify the **branching (side) groups** attached to this continuous chain of carbon atoms.
3. **Number the carbon atoms** in the longest continuous chain and ensure that the groups would be attached to the carbon assigned to the lowest number possible. This means that you have to decide whether to number from the right or left end of the chain or from top to bottom and vice versa.
4. The names of the side groups and the numbers of the carbon atoms to which they are attached will form prefixes of the parent name.
5. If more than one of the same kind of branched groups are attached to the chain, identify the numbers on the carbon atoms bearing each group and write this as a series of numbers separated by commas between each number and in an increasing order, then a hyphen and finally use a Greek prefix attached to the branch name as discussed under rule No. 6 below.
6. **Substituent groups** – Substituent groups are kinds of side or branch groups and are named according to the number of carbon atoms. A branch with one carbon atom ($-CH_3$) is known as **methyl** group, while that with 2 carbon atoms is known as **ethyl** group ($-CH_2CH_3$).

In alkanes and many organic compounds, hydrogen atoms which are bonded to the continuous carbon chain are substituted by groups. The commonest substituent groups or branches are obtained from alkanes whose molecules have one hydrogen atom less. The general name of such a substituent group is an **alkyl group**. The name is obtained from the name of the alkane by replacing **'–ane'** with **'–yl'**.

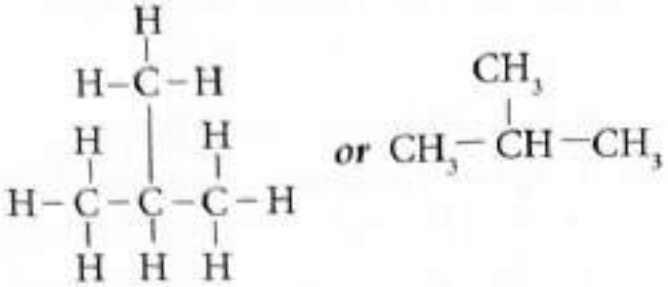
Naming the alkyl groups

Alkyl groups are named by adding *suffix -yl* to the prefix of the parent name as given in Table 3.1

Table 3.1: Naming alkyl groups

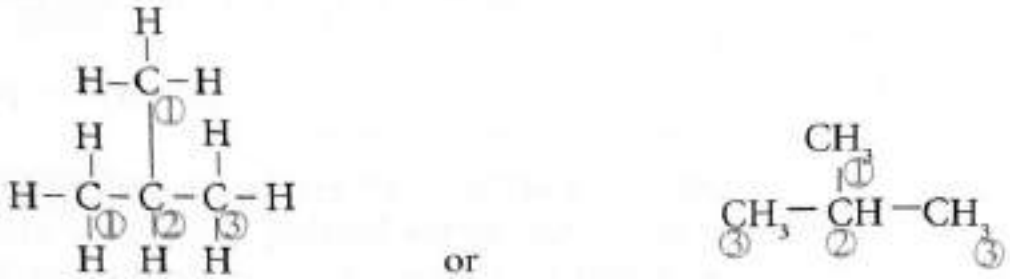
Number of carbon atom in the longest continuous carbon chain	Name of alkane	Prefix	Name of alkyl group	Formula of the alkyl
1	methane	meth-	methyl	CH_3-
	ethane	eth-	ethyl	CH_3CH_2-
3	propane	prop-	propyl	$CH_3CH_2CH_2^-$
4	butane	buty-	butyl	$CH_3CH_2CH_2CH_2^-$

Let us now apply the rules listed above to name a branched alkane which has the following structures.



Rule no. 1 - for parent name

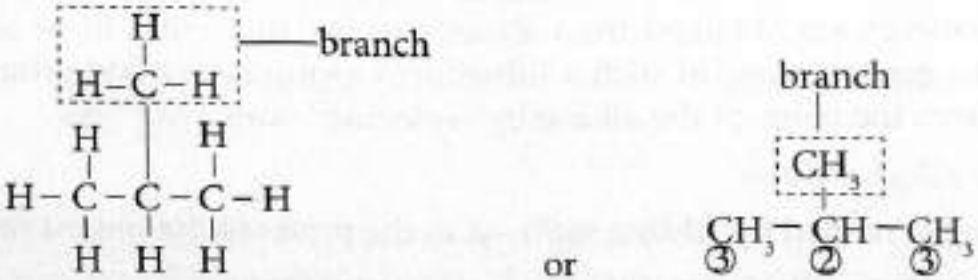
Identify the longest continuous chain of carbon atoms



- When you trace the carbon chain from all sides, the longest continuous carbon chain has three carbon atoms. Therefore, the ending or parent name will be *propane*.

Rule no. 2 - for branches

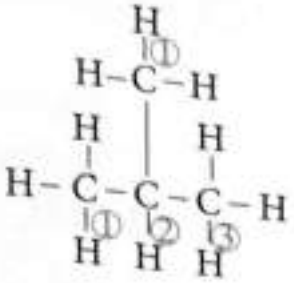
Identify the branches (substituent groups)



There is one **methyl group** branched off the continuous carbon chain. The prefix will therefore be **methyl**, hence the name is **methylpropane.**

Rule no. 3 - for position of branches on the carbon chain

Number the carbons in the continuous chain so that the methyl group will be attached to the carbon atom assigned the lowest number on the continuous carbon chain.In this case, it is important to number carbon atoms from all sides of the chain to be able to identify the lowest possible numbering of the position(s) of the branch(es).



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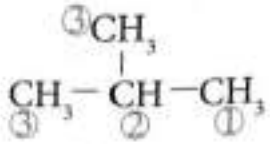
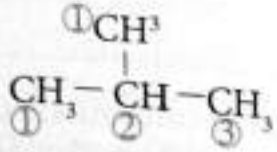
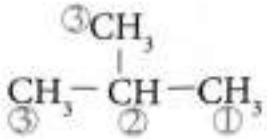






- From all sides, the lowest number on the carbon atom within the chain that is bearing the methyl group is number two (2).

Rule no. 4 - for inserting a hyphen

Write carbon number bearing the methyl group followed by a hyphen, then the name of the alkyl group, for example, 2-methyl.

Rule no. 5 - for ending name

Use the parent alkane name corresponding to the number of the carbon atoms on the longest continuous carbon chain and write this next to the alkyl group as:

2-methylpropane

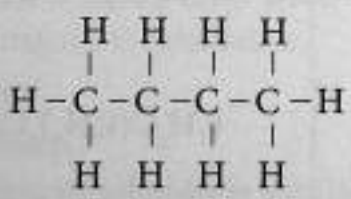
Rule no. 6 - for identical

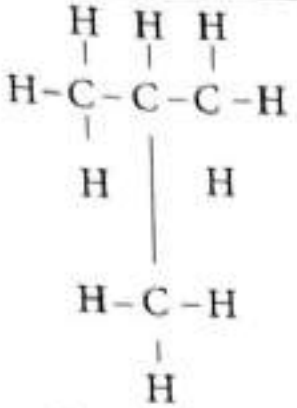
For branches on the same carbon atom, we write the number on that carbon twice before the name of the compound

Examples of isomers of alkanes

1. Isomers of butane, (C_4H_{10})

Butane has four carbon atoms. The four carbon atoms can be arranged in two different ways to give two isomers as shown in the examples below.

	Structural formulae	Condensed formulae
(a)	 Butane	$CH_3CH_2CH_2CH_3$

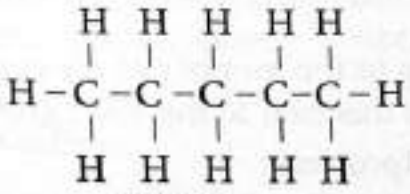
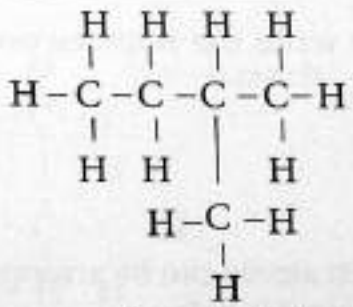
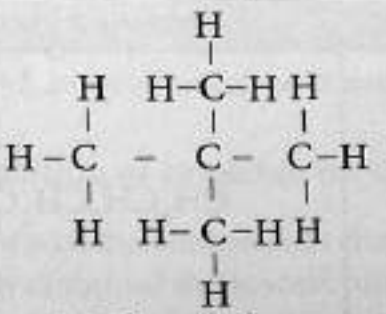
(b)	 2-methlypropane	$CH_3CH(CH_3)CH_3$

Notice that the structure in (b) a carbon atom is branched off the continuous chain of carbon atoms in the molecule. Therefore, the isomers are butane and 2–methylpropane.

Note: The branched isomers have lower melting point and boiling point than unbranched isomers.

2. *Isomers of Pentane, (C_5H_{12})*

Pentane has five carbon atoms hence its structural and condensed formulae are as shown below.

	Structural formulae	Condensed formulae
(a)	 Pentane	$CH_3CH_2CH_2CH_2CH_3$
(b)	 2-methylbutane	$CH_3CH(CH_3)CH_2CH_3$
(c)	 2,2-dimethylpropane	$CH_3C(CH_3)_2CH_3$

Therefore, the isomers of pentane are

- Pentane
- 2–methylbutane
- 2, 2 – dimethylpropane

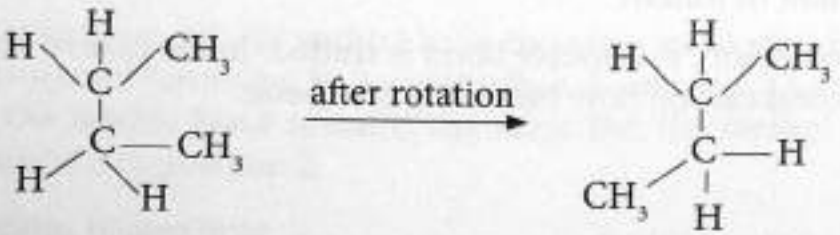
Self assessment exercise 5.1

1. Define the following terms:
   (a) Isomerism
   (b) Isomers
2. Draw and name all the structural isomers of:
   (a) Pentane
   (b) Butane

Note: In the condensed formulae, (CH_3) and $(CH_3)_2$ indicate branches. The branched alkanes have the same molecular formulae as straight chain alkanes but the way the carbon atoms are linked to one another differs. This is the reason why they are called **isomers.**

Conformational isomers (conformers)

Single bonds between carbon-carbon atoms in alkanes can rotate. By rotation of the bond, one can convert to another isomer. Such isomers are referred to **conformation isomers** or **conformers.**



Conformer

Can you draw other conformers of butane shown above

5.2 Isomerism in alkenes

Earlier, we learnt that alkanes only show chain isomerism. Unlike alkanes, alkenes show two types of structural isomerism:

- Chain isomerism
- Positional isomerism

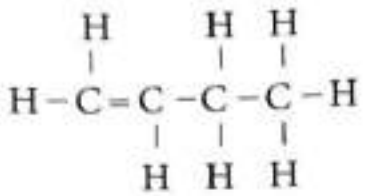
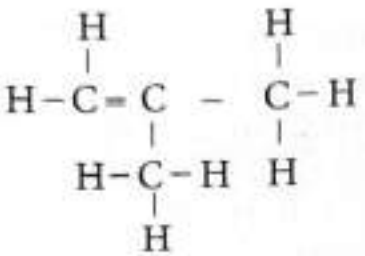
IUPAC naming system of alkenes is similar to that of alkanes except that the longest continuous chain of carbon atoms with a double bond between two carbon atoms is considered to give the parent name. The chain is numbered starting from the side that gives the double bond the lowest number.

Examples of isomers of alkenes

1. *Isomers of butene*

Chain or branched isomerism in butene

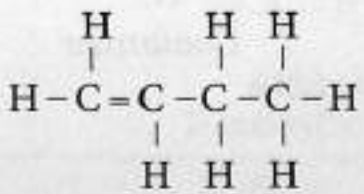
In this type of isomerism, the position of the double bond remains the same but a methyl group is shifted to a different location in the carbon chain e.g the second carbon in (b).

	Structural formulae	Condensed formulae
(a)	 But-1-ene	$CH_2{=}CHCH_2CH_3$
(b)	 2-methyl prop-1-ene	$CH_2{=}C(CH_3)CH_3$

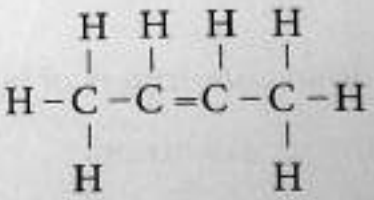
Position isomerism in butene

In this type of isomerism, the double bond is shifted. In the case of but-1-ene it can be shifted to the second carbon now becoming but-2-ene.

For example:



But-1-ene

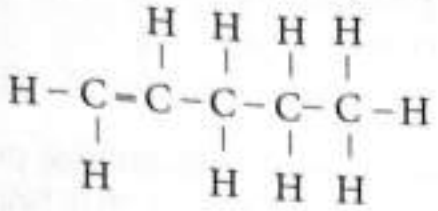
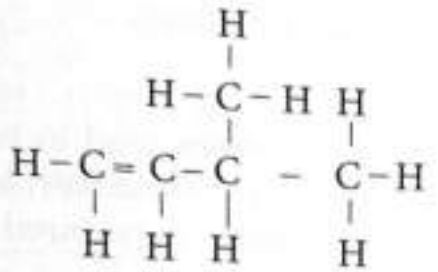
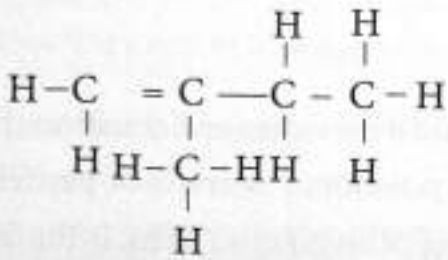


But-2-ene

2. *Isomers of pentene*

Chain isomerism in pentene

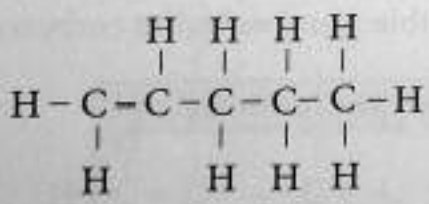
It is also possible to draw other different structural formulae of pent-1-ene. The atoms of carbon are arranged in a different way in each structure as shown below:

	Structural formulae	Condensed formulae
(a)	 Pent-1-ene	$CH_2CH(CH_2)_2CH_3$
(b)	 3-methylbut-1-ene	$CH_2CHCH(CH_3)CH_3$
(c)	 2-methylbut-1-ene	$CH_2C(CH_3)CH_2CH_3$

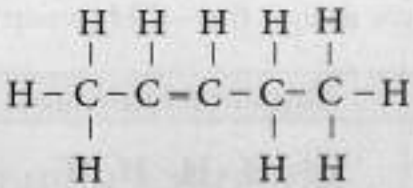
The above compounds i.e. (a), (b) and (c) have the same molecular formula but they have different structural formulae. In 3 – methylbut–1–ene and 2 – methylbut–1–ene, the ***position of the double bond*** remains the same but the ***methyl group*** has been moved from position 3 to position 2.

Position isomerism in pentene

The position of the double bond changes to carbon number 2 as shown below



Pent-1-ene



Pent-2-ene

Note that in the above isomers, the position of the *double bond* shifts while the position of the carbon atoms remain the same.

Effect of branching on physical properties of hydrocarbons

Branching in hydrocarbons affects physical properties such as density and melting and boiling points of individual hydrocarbon compounds.

(a) Boiling point

Generally, branched hydrocarbons have lower melting and boiling points than the corresponding straight chain isomers. This is due to the fact that branching of the chain makes the molecule more compact and thereby decreases the surface area. This lowers Van der Waal's forces of interactions. The forces thus require little amount of energy to break during transition from liquid to gaseous phase.

(b) Melting point

However, melting points of highly branched hydrocarbons tend to be higher than their corresponding branched chain isomers. This is because molecules are packed closer together in their solid phase, hence more energy is required to break the intermolecular forces during melting.

Self assessment exercise 5.2

1. What are isomers?
2. Is there any difference between branched isomerism and positional isomerism?
3. Draw and name all the branched and positional isomers of pentene.
4. What is the main feature in the bonding of an alkene? Why is the first member of this series not methene?

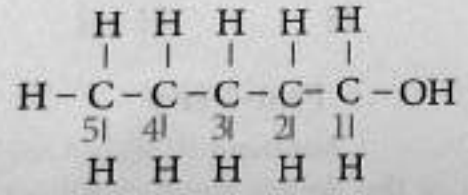
5.3 Isomerism of alkanols

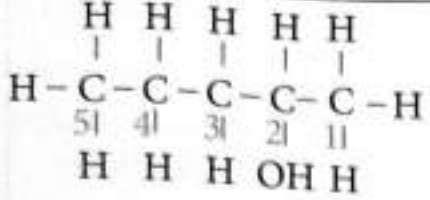
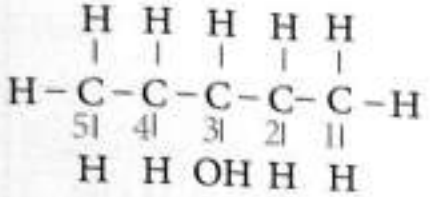
Alkanols contain – OH group. We name them using the suffix – ol. A number is usually used to indicate the position of the functional group within the compound.

Examples of isomers of alkanols

1. Isomers of pentanol

Pentanol is derived from an alkane with five carbon atoms i.e. pentane. In naming the alkanol, we assign the – OH group the lowest possible number in the carbon chain.

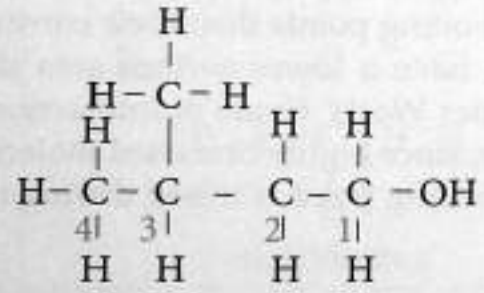
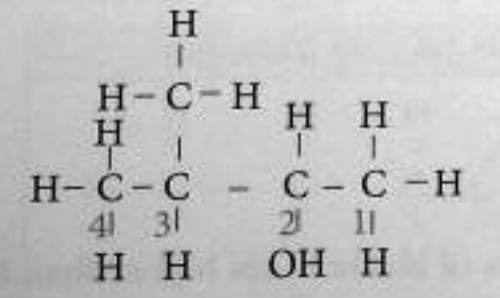
	Structural formulae	Condensed formulae
(a)	 Pent-1-ol	$CH_3CH_2CH_2CH_2CH_2OH$

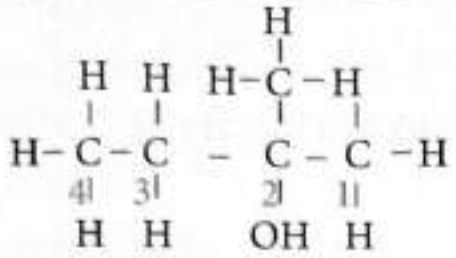
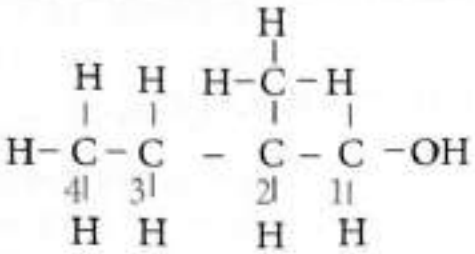
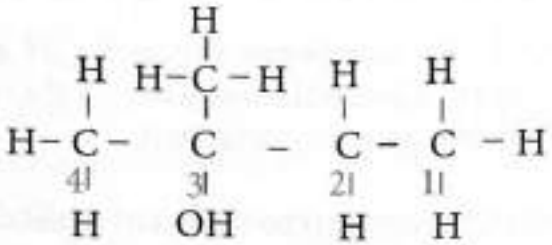
(b)	 Pent-2-ol	$CH_3CH(OH)CH_2CH_2CH_3$
(c)	 Pent-3-ol	$CH_3CH_2CH(OH)CH_2CH_3$

The above are positional isomers of pentanol.

Note: In the above chain isomers of pentanol, the functional group – OH group is attached on 1st, 2nd and 3rd carbon atom. Hence the number in the name indicates the carbon to which the functional group is attached.

We can get structural isomers by shifting the methyl group to different carbon atoms as shown below.

	Structural formulae	Condensed formulae
(a)	 3-methylbutan-1-ol	$CH_3CH(CH_3)CH_2CH_2OH$
(b)	 3-methylbutan-2-ol	$CH_3CH(CH_3)CH(OH)CH_3$

(c)	 2-methylbutan-2-ol	$CH_3CH_2CC(CH_3)OHCH_3$
(d)	 2-methylbutan-1-ol	$CH_3CH_2C(CH_3)CH_2OH$
(e)	 3-methylbutan-3-ol	$CH_3C(CH_3)OHCH_2CH_3$

Branched alcohols exhibit lower melting and boiling points than their corresponding straight chain alcohols. This is because they have a lower surface area than their unbranched isomers. So there are lower van der Waals' forces of interactions. These forces require lesser energy to break. However, since highly branched molecules tend to pack closer together in their solid phase, melting point is raised during transition from solid to liquid.

Self assessment exercise 5.3

1. Name each of the following compounds.
   i) $CH_3CH_2CH_2CH_2OH$
   ii) $CH_3CH_2CH_2OH$
   iii) CH_3CH_2OH
2. Explain briefly why the boiling points of alcohols are higher than those of alkanes of corresponding molecular mass.
3. Draw the structural formula of the following compounds:
   $CH_3CH_2CH_2OH$

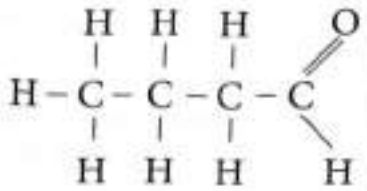
5.4 (a) Isomerism in alkanals

These are compounds with a functional group called **carbonyl group**. The carbonyl group (-CHO) is placed at the end of the carbon chain in an **alkanal**.

The general formular of alkanals is $(R-\overset{\overset{\displaystyle O}{\|}}{C}-H)$ usually written as RCHO where R represents an alkyl group or a hydrogen atom.

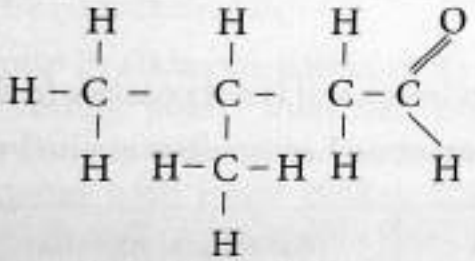
Nomenclature

When naming alkanals, the suffix **'e'** in the corresponding alkane is replaced with **'–al'**.

condensed formulae	structural formulae
$CH_3CH_2CH_2CHO$ Butanal	

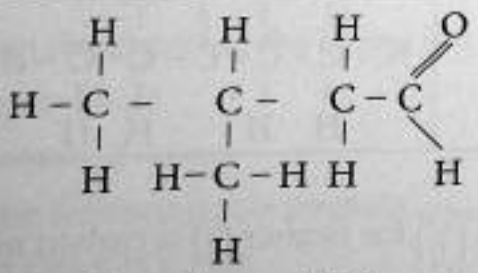
1. *Isomers of butanal*

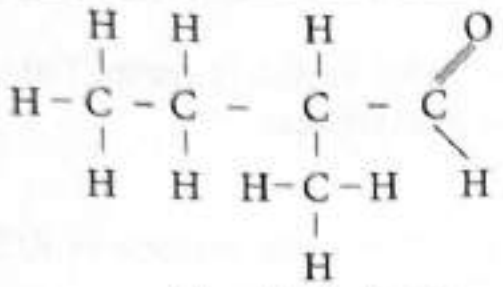
Branched isomers of butanal

Structural formulae	Condensed formulae
 3-methylbutan	$CH_3CH(CH_3)CH_2CHO$

2. *Isomers of Pentanal*

Branched isomers of pentanal

Structural formulae	Condensed formulae
 3- methypentanal	$CH_3CH_2(CH_3)CH_2CHO$

 2-methylbutanal	$CH_3CH_2CH(CH_3)CHO$

Note: Alkanes do not exhibit positional isomerism because the functional group (-CHO) is placed at the end of an alkanal molecule.

5.4 (b) Isomerism in alkanones

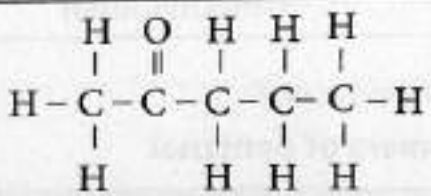
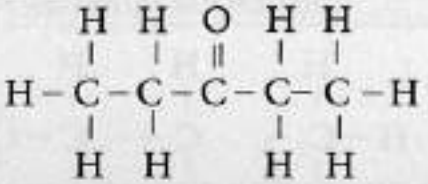
Alkanones are relative of alkanals the general formula of alkanones is RCOR or ... $R-\overset{\overset{O}{\|}}{C}-R'$ ..., where R and R' are carbon chains (alkyl group) whose number of carbon atoms may be the same or different. Note that in alkanones, the 'R' group replace 'H' in alkanals.

Nomenclature of alkanones

In naming alkanones, the suffix **– one** replaces the **'e'** in the corresponding alkane. Usually a number, if necessary is used to indicate the position of the carbonyl group. The first ketone is **propanone**. The name is derived from the alkane - **propane**. The other ketones are butanone and pentanone.

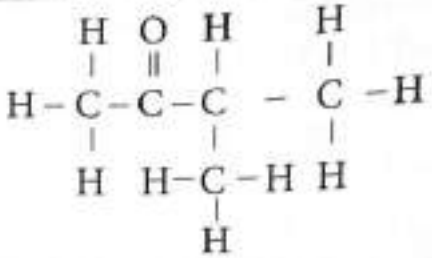
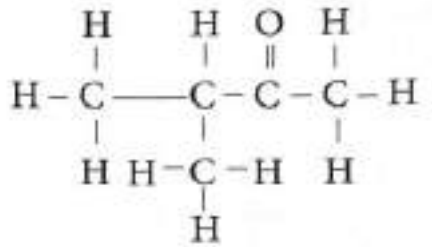
Examples of Isomers in alkanones

Propanone and butanone have no isomers because it is not possible to shift the carbonyl group $\left(\overset{\overset{O}{\|}}{C}\right)$ to a **different** position, but pentanone has isomers as shown below.

	Condensed formula	Structural formula
(a)	$CH_3COCH_2CH_2CH_3$ Pentan - 2 -one	
(b)	$CH_3CH_2COCH_2CH_3$ Pentan - 3 -one	

Note: Location of the carbonyl group i.e. $\left(\overset{\overset{O}{\|}}{C}\right)$, for isomer (a) is carbon number 2 and for isomer (b) is carbon number 3.

(c) We can also shift the location of a substituent group CH_3 to carbon number 3 in isomer (a) and carbon number 2 in (b) as indicated below. This forms structural isomers of pentanone.

	Condensed formula	Structural formula
(a)	$CH_3COCH(CH_3)CH_3$ 3-methylbutan-2-one	
(b)	$CH_3COCH(CH_3)CH_3$ 2-methylbutan-2-one	

Note; When naming the alkanones, always number the carbon atoms from the end that gives the functional group the lowest number possible. If you number carbon atoms in (b) from left to right, the carbonyl group will bear number 3 which is wrong.

Physical properties of alkanones

The carbonyl group in alkanone compounds is a **polar group**. This causes alkanones to have higher boiling points than the corresponding hydrocarbons of the same molecular mass. But since they do not form strong hydrogen bonds between their molecules, alkanones have lower boiling points than their corresponding alcohols. In contrast, the carbonyl oxygen allows the molecules of alkanones to form strong hydrogen bonds with water molecules. As a result, low-molecular mass alkanones show appreciable solubility in water.

Self assessment exercise 5.4

1. Draw the structures of the following compounds:
   i) Butan-1-one
   ii) Pentan – 2 – one
2. Name the following compounds.
   i) $CH_3CH_2CH_2CHO$
   ii) $CH_3CH_2COCH_2CH_3$

5.5 Isomerism in alkanoic acids (carboxylic acids)

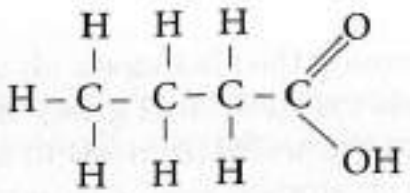
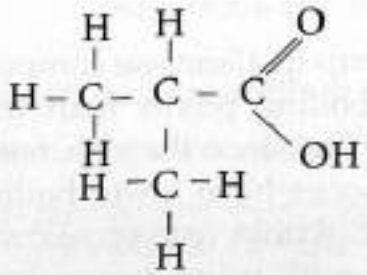
Alkanoic acids are also know as carboxylic acids. IUPAC names for alkanoic acids are obtained by replacing the final **'e'** of the corresponding alkane with **'oic** acid'.

The general formula of carboxylic acids is R-COOH or $R-\overset{\overset{O}{\|}}{C}-OH$ whereby R represents the carbon chain except in methanoic acid where it represents H and –COOH or $-\overset{\overset{O}{\|}}{C}-OH$ is the functional group. It is referred to as the **carboxyl group**.

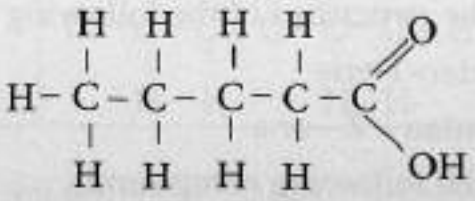
Examples of Isomerism in alkanoic acids

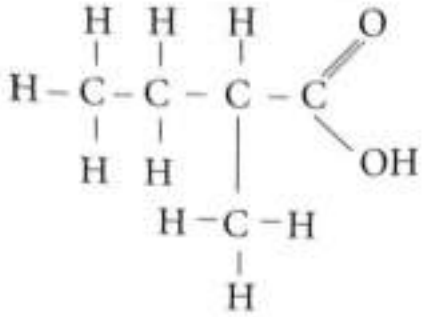
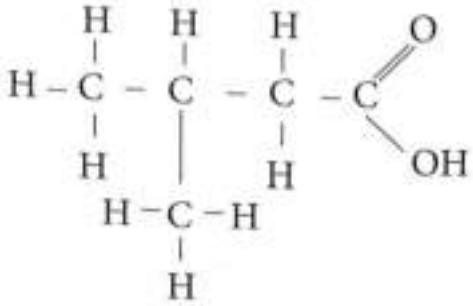
The isomers of alkanoic acids are formed as a result of shifting a methyl group ($-CH_3$) to different positions.

1. Isomers of butanoic acid

	Condensed formula	Structural formula
(a)	$CH_3CH_2CH_2COOH$ butanoic acid	
(b)	$CH_3CH(CH_3)COOH$ 2 -methylpropanoic acid	

2. Isomers of pentanoic acid

	Condensed Formula	Structural Formula
(a)	$CH_3CH_2CH_2CH_2COOH$ Pentanoic acid	

(b)	$CH_3CH_2CH(CH_3)COOH$ 2 -methylbutanoic acid	
(c)	$CH_3CH(CH_3)CH_2COOH$ 3 -methylbutanoic acid	

The relative formula mass of the above branched acids are the same as their corresponding straight chain acid. But the melting and boiling point of straight chain molecules will be higher. The solubility in water will also be less than that of straight chain carbonxylic acids.

Polarity on alkanoic acids

Alkanoic acids are polar substances. Their molecules can form intermolecular forces of attraction called **hydrogen bonds** with each other. For example, ethanoic acid molecule can form intermolecular forces of attraction with each other as follows.

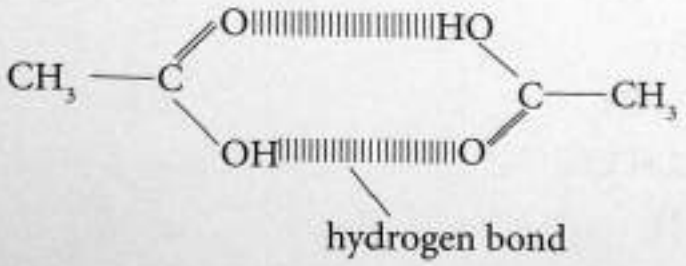


Fig. 5.2: Hydrogen bond

Effects of hydrogen bonding and branching on melting and boiling points of alkanoic acids

As a result of hydrogen bonding in alkanoic acids, they have high melting and boiling points. Branching and increase in length of the carbon chain decreases the boiling point of alkanoic acids. Low molecular mass alkanoic acids are appreciably soluble in water. The first four members of the alkanoic acids series are miscible with water in all proportions. But as the length of the carbon chain increases, their solubility decreases.

Self assessment exercise 5.5

1. Name the following compounds:
   (a) HCOOH
   (b) $CH_3CH2COOH$
   (c) $CH_3CHCOOH$
   $\quad\quad\quad

(d) $CH_3CH_2\overset{\overset{\large O}{\|}}{C}OH$

(e) $CH_3\overset{\overset{\large CH_3}{|}}{C}HCOOH$

(f) $CH_3\underset{\underset{\large CH_3}{|}}{C}HCOCH_3$

3. Explain the effects of branching on melting and boiling point of alcohols.
4. Draw the structural formulae of
   (a) Pentan-3-one
   (b) Ethanol
5. Name the straight-chain compound having the formula C_5H_{12}.

Topic 6 Polymerisation

Success criteria

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

(a) Define polymer.

(b) State types of polymerisation.

(c) Describe properties and uses of synthetic polymers.

(d) Explain the difference between thermosoftening and thermo setting plastics.

Introduction

All living things contain small molecules linked together. When very many of these molecules are linked together they form substances known as **polymers**. Polymers are large organic molecules made up of small repeating molecules known as **monomers** and are also referred to as **macromolecules.** Fig. 6.1 shows an example of polymer. A polymer can be made of similar or different monomers. The process of joining up the monomers is known as **polymerisation.**

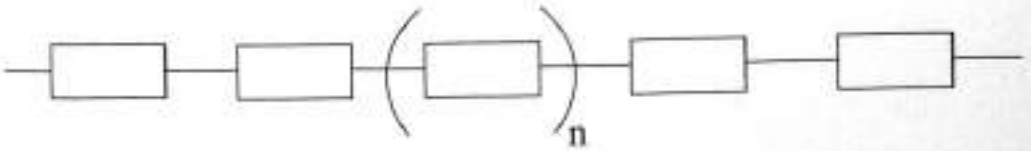


Fig. 6.1: Example of a polymer.

The monomers may be repeated very many times. This can be represented by letter'n' which may range from over few units to a million. Polymers having the same kind of monomers are known as **homopolymers,** for example in polyethene, polypropene and polychloroethane.

Polymers made of two or more different monomers are known as **copolymers.** Examples of copolymers are nylon and proteins. Proteins are made from 20 different monomers of amino acids.

6.1 Types of polymers

Research Activity

Find out about the various types of polymers from the library or internet. Write a report and share with your friends.

Polymers can be classified into two categories:

- Natural polymers.
- Synthetic (artificial) polymers.

(a) Natural polymers

All living things contain natural polymers. Examples include proteins, carbohydrates, starch, silk, wood, natural rubber and DNA. Can you name places or things where the above polymers are founds.

Some examples of natural polymers, their monomers and sources are listed in Table 6.1

Table 6.1: Some natural polymers and their sources.

Polymer	Monomer	Sources
Carbohydrates for example Starch	Glucose	Potato, maize and wheat flour
Cellulose	Glucose	Paper, wood and cotton
Sucrose	Fructose and glucose	Sugarcane and beetroot
Protein	Amino acids	Muscles, tissue, wool and silk
DNA	Nucleotides	Genes and chromosomes

Note: Natural polymers do not pollute the environment because bacteria can break them down hence causing decomposition. i.e. they are **biodegradable.**

Advantages of natural polymers

- They are biodegradable and can easily be disposed.
- They do not produce poisonous gases when they burn.
- Are made from renewable resources such as trees.

Disadvantages of natural polymers

- They cannot be moulded into complicated shapes like synthetic polymers.
- They can be expensive.
- They often do not last for very long.

(b) Synthetic Polymers

These are man-made polymers and fibres which are made by copying the natural process of linking monomers during polymerisation. Can you name some sythetic polymers that you know.

Some examples of synthetic polymers, their monomers and products made from them are given in Table 6.2 below.

Table 6.2: Some synthetic polymers and their products

Polymer	Monomer	Products
Polyethene	Ethene	Plastic bags and bottles, bowls, packaging materials,etc
Polychloroethane (PVC)	Chloroethane	Crates and boxes, water pipes, artificial leather and electrical insulations.
Polytetrafluoroethene (PTFE) "Teflon"	Tetrafluoroethene	Non-stick coating on pans
Nylon	1,6 – diaminohexane hexane -1,6 –dioyl dichloride	Stockings, carpets and clothing.
Terylene (a polyester)	Ethane -1,2 –diol and 1,4-dicarboxylic acid	Clothing for example shirts, trousers, sails and ropes.

Advantages of using synthetic polymers

- They are light and therefore portable.
- They are soft, and easy to manufacture.
- They can be made into complicated shapes more easily that is, they are malleable.
- They are good insulators.
- They are resistant to corrosion from acids, alkalis, water and air.
- They are colourful and attractive.
- They are cheap and therefore affordable.

Disadvantages of using synthetic polymers

- They have comparatively low strength.
- They have poor heat resistance.
- They have poor weather resistance.
- They are non-biodegradable hence difficult to dispose.
- Some synthetic polymers give off poisonous gases such as carbon monoxide when they are burnt. These gases pollute the air.
- They increase levels of gases such as carbon dioxide in the atmosphere when they burn. Excess accumulation of these gases cause greenhouse effect and global warming.

Remember!

We should always avoid burning plastics and tyres or burying them underground as this causes pollution. We should all participate in ensuring a clean and safe environment.

Self assessment exercise 6.1

1. What do you understand by the following terms:
   (a) Monomer
   (b) Polymerisation
   (c) Polymer.
2. Give two examples of:
   (a) Natural polymers.
   (b) Synthetic polymers.
3. Differentiate between biodegradable and non-biodegradable polymers.
4. Explain briefly why synthetic polymers are a big 'nuisance' to our environment.
5. Where can you find the following polymers?
   (i) DNA (ii) Proteins
   (iii) PVC (iv) Teflon

6.2 Types of polymerisation

There are two types of polymerisation:

- Addition polymerisation
- Condensation polymerisation

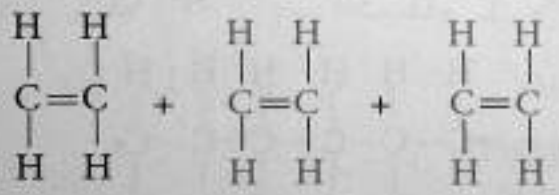
(a) Addition polymerisation

In this type of polymerisation, the monomers simply add into each other to form a longer molecule. During polymerisation of ethene for example,the double bond between the carbon atoms is broken down enabling many molecules of ethene to join to each other to form a large molecule called **poly-ethene** which is commonly called **polythene**. The prefix "poly" means "many". Polythene is therefore an addition polymer. The reaction can be summarised in the following equations. Suppose we represent ethene by a box ▭ ,we can join many boxes ▭-▭-▭-▭-▭- to represent polyethene. Compare this with the actual reaction below.

(i) Ethene

$$\text{Ethene} \xrightarrow[\text{heat, catalyst}]{\text{high pressure}} \text{polythene}$$

Monomer (Ethene)



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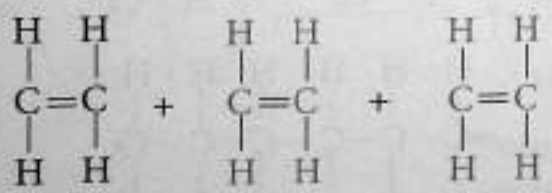
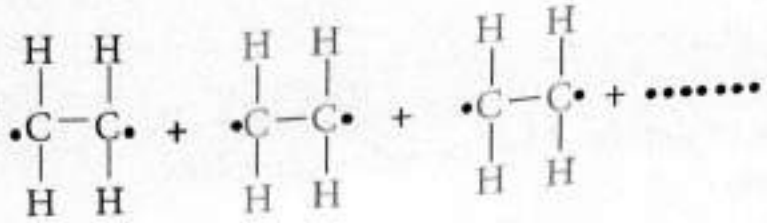
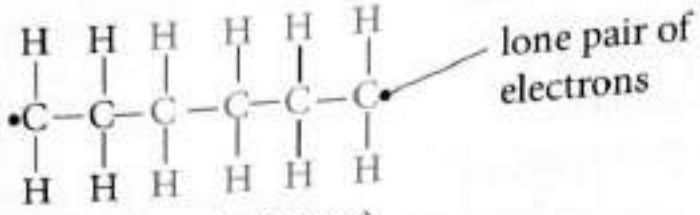


Double bond broken



Uniting the three monomers



Polyethene (polymer)

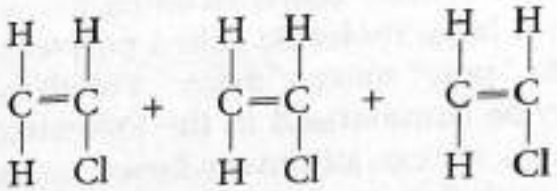
Note that at each end of the polymer, there is a lone electron indicating that more monomers can be joined to form a longer polymer. In this way, a very long chain of carbon atoms is formed. The chain may contain over 50,000 carbon atoms or less. If n molecules of ethene combine, the process can be written as:

Polyethene

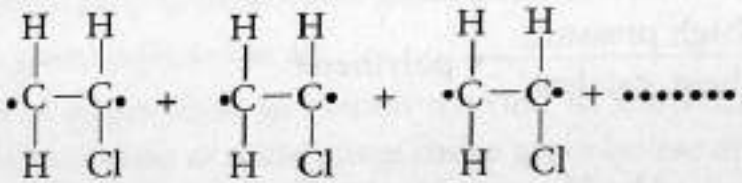
Other molecules also form polymers in a similar manner as shown below.

(ii) Polychloroethene (PVC)

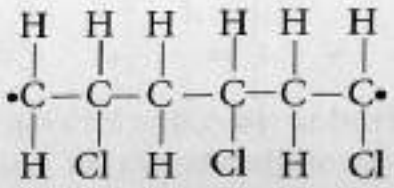
Monomers (chloroethene)



Double bond broken



Uniting the three monomers



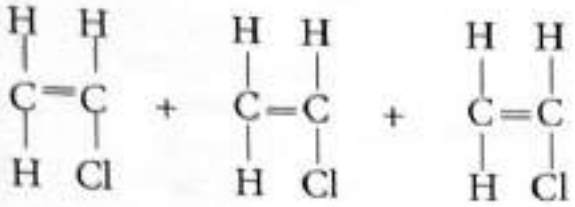
Polychloroethene (polymer)

Again if 'n' molecules of chloroethene combine we get;

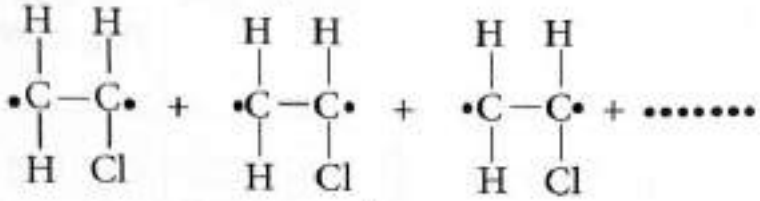
Polychoroethene(PVC)

OR

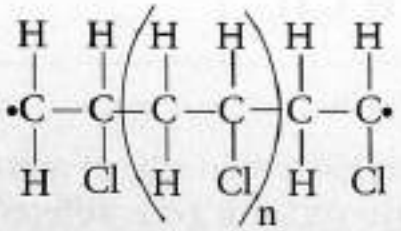
Monomers (Chloro-ethene)



Double bond broken



Uniting the monomers



Polychloroethene (polymer)

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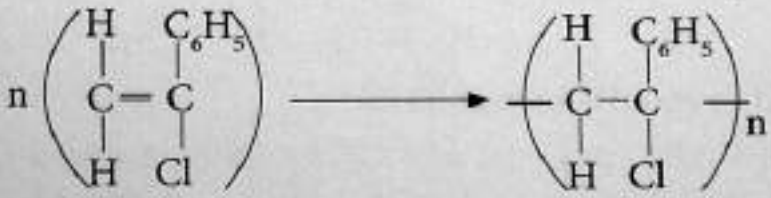
(iii) Polyphenylethene

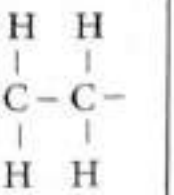
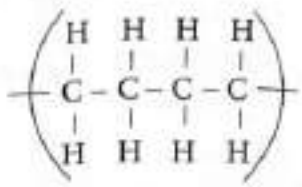
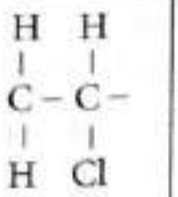
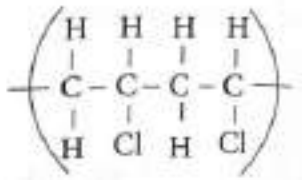
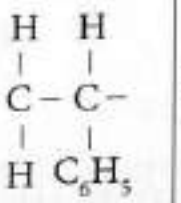
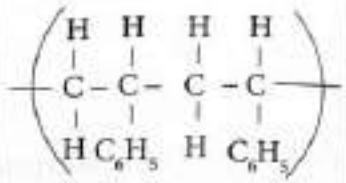
In the formation of polyphenylethene, n-molecules combine as follows:



Polyphenylethene
(or polystyrene)

Note that C_6H_5 can be represented as . It is known as the benzene ring

Table 6.3: Summary of polymers formed by addition polymerisation

Name of polymer	Formula of monomers	Name of monomers	Reaction conditions	Part of polymer chain
Polyethene (high density form)		Ethene (gas)	Low temperature and pressure, special catalyst	
Polychloroethene		Chloroethene	60°C, high pressure, H_2O_2 catalyst	
Polyphenylethene (Polystyrene)		Phenylethene (styrene)	Catalyst, heat	

Note; If you replace one H atom in ethene with a different atom such as Cl,C_6H_5 or substituent group, you will get a different polymer as in the formation of:

- Polychloroethene (Polyvinylchloride).
- Polyphenylethene above.

Experiment 6.1

Aim: *To investigate the process of polymerisation*

Apparatus:

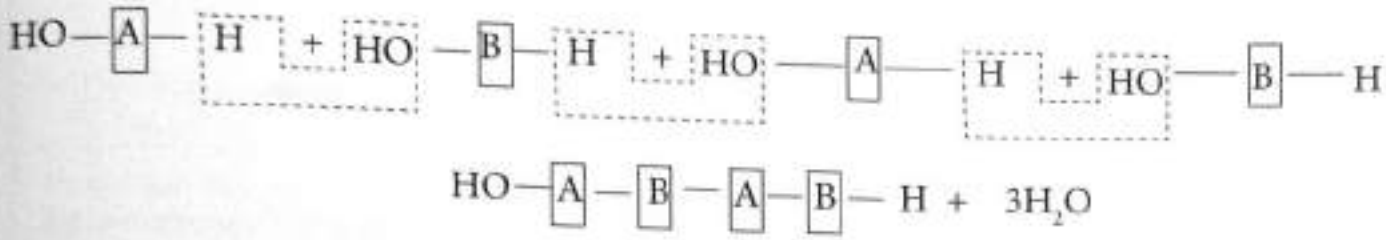
- Blocks of wood of equal dimensions
- Super glue
- Working table
- Cutting instrument, for example, a knife

Procedure

1. Sharpen some blocks on the width to form a pointer on one side and curve out holes in other blocks into which the pointers can enter to fit on the blocks.
2. Fix the blocks one into the other continuously and apply super glue to make sure they are firmly held to one another.
3. Observe how from one small block (monomer), we can produce a totally new product which can be used for purposes that one block could not do. That is how a polymer is formed.

(b) Condensation polymerisation

This is the formation of a single molecule by the combination of two or more molecules. This process is usually accompanied by **elimination of a simple molecule** such as water or hydrogen chloride. For example, if we want to join the boxes below, we can do that by giving off simple water molecules as shown.



"Poly AB"

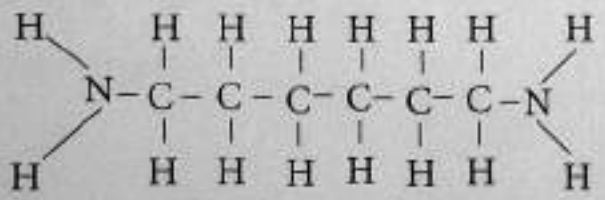
Examples

Nylon and terylene are made by condensation polymerisation.

Nylon

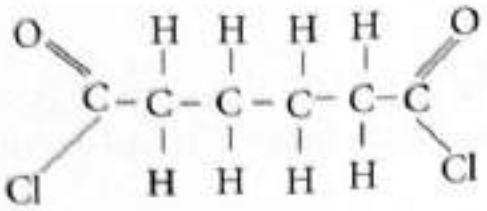
Let us consider condensation polymerisation of nylon. Nylon is a polymer of 1,6-diaminohexane and hexane-1,6-dioyl dichloride.

The structure of 1,6-diaminohexane is



1,6-diaminohexane

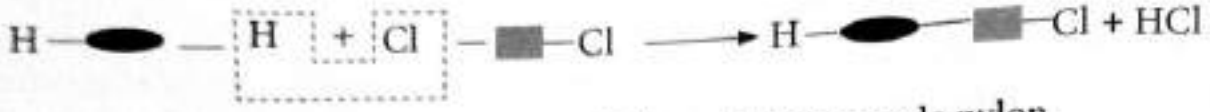
Which can be represented as H — ■ — H for simplicity while the structure of hexane-1,6- dioyldichloride is



Hexane-1,6-dioyl dichloride

Which can be represented as Cl — ■ — Cl for simplicity.

When a molecule of 1,6-diaminohexane reacts with a molecule of hexane–1,6–dioyl dichloride, the two join by eliminating a hydrogen chloride molecule as follows:

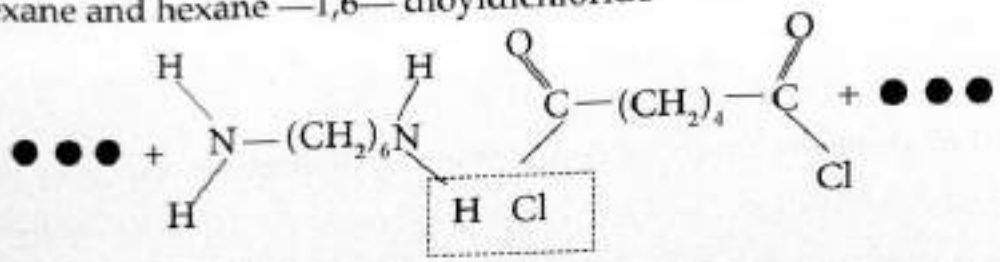


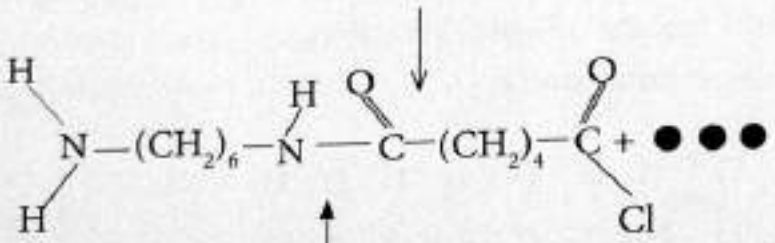
'Polymer' for example nylon

This process can occur again and again forming a long polymer which is called nylon as shown below.

Longer polymer

Now compare the linking of actual condensation formula of the monomers 1,6–diaminohexane and hexane —1,6— dioyldichloride





bond joining the two monomers

When the process is repeated several times, one of the H atoms on the left hand side and Cl atom on the right hand side combines with Cl and H respectively and HCl molecule is formed.

6.3 Properties of synthetic polymers

Although different polymers have unique characteristics, many synthetic polymers have the following common properties.

- They do not corrode in air or water.
- They do not decompose easily.
- They are relatively light.
- Some synthetic polymers are soft and malleable while others are very strong.
- They are cheap to produce.
- Some are transparent.

Table 6.4: Summarry of properties and uses of some natural and synthetic polymers

Name of polymer	Common name	Properties	Uses
Polyethene	Polythene	• Tough • Durable	Used in making: • pipes and plastic bags • bowls and bottles • packaging material • insulation for electric wiring
Polychloroethene	Polyvinyl chloride (PVC	• Tough • Durable	Used in making: • crates and boxes • plastic ropes. • artificial leather • water pipes • insulation for electrical wires
Polyphenylethene	Polystyrene	• Light • Poor conductor of heat	Used in making: • insulation material. • packaging (foam) • ceiling tiles • clothes • plastic model kits
Terylene	Terylen (a polyester)	• Light • Tough • Durable	Used in making: • clothings like shirts and trousers • sails and ropes • plastic model kits
Nylon	Nylon	• Light • Tough • Durable	Used in making: • ropes • clothings • carpets

6.4 Thermoplastics and thermosetting polymers

Plastics form a group of polymers which can withstand deformity and which can easily be moulded, particularly at high temperatures. Plastics are generally tough, resistant to the action of acids and alkalis and are not easily affected by a fair range of temperature. The properties of a plastic depend on the extent of polymerisation. The nature and arrangement of the molecular chains in different polymers will affect the strength and melting point of a polymer. These properties will increase with the increase in length of the chain. The structure of monomers also influence the properties and the structure of the polymer. For example, branched monomers will form branched polymers. Polymers made of such monomers have irregularly packed structure thus leading to amorphous polymers with low tensile strength and low melting point. Polymers with no branched chains are regularly arranged and well packed. This gives them high tensile strength and high melting points.

There are two types of plastics:

- Thermosoftening (thermoplastic) plastics
- Thermosetting plastics

(a) Thermosoftening plastics

- Thermosoftening (or thermoplastic) are those plastics which soften when heated and harden when cooled. They can be re-melted indefinitely without losing their properties. Due to their ability to soften on warming, they can be moulded into different shapes.

Examples of thermosoftening plastics are polyethene, polymethyl - 2 - methyl-propanoate.

(b) Thermosetting plastics

Thermosetting plastics on the other hand cannot be remelted once they have been formed. They can be heated and moulded only once. For example, **melamine**.

The chains in these polymers are cross-linked to each other. These **cross-links** in the plastics are permanent chemical bonds. They make the structures rigid when moulded and no softening can take place on heating. Thermosetting plastics can be re-converted into thermoplastics by opening the cross-links. Note that, the thermosetting plastics can be made soft and readily moulded by addition of certain compounds known as **plasticisers**. Polychloroethene is stiff and hard but the addition of **tricrosy phosphate** makes it soft and rubber-like. These cross-links make the structure rigid when moulded, and no softening take place on heating. Examples include bakelite, rigid, hard and brittle.

Table 6.5: Comparison between thermosoftening plastics and thermosetting plastics

	Thermosoftening plastics	Thermosetting plastics
1.	They can soften on heating and can be re-moulded.	They do not soften on heating and cannot be re-moulded.

2.	They have no cross-link chains but are held together by Van der Waal forces of attraction.	They have cross-link chains which are held together by permanent chemical bonds.
3.	They are highly branched hence have lower tensile strength and low melting points.	They have few or no branched chains hence high tensile strength and high melting points.

Self assessment exercise 6.2

1. Differentiate between thermo-softening and thermosetting plastics.
2. Which of the following plastics cannot be re-melted or re-moulded?
   (a) Polyethene (b) Bakelite
3. (a) Polymerisation reactions may be classified as addition or condensation reactions. Which of the processes is used to manufacture bakelite.
   (b) Give one advantage PVC could have over bakelite.
4. Explain why dumping of plastics in holes underground and burning are not recommended for disposal of plastics.

Revision exercise 6

1. Define the following terms:
   (a) Synthetic polymer
   (b) Natural polymer
2. How many products are formed from:
   (a) addition polymerisation?
   (b) condensation polymerisation?
3. State two uses of each of the following polymers:
   (a) Nylon (b) Terylene
4. Ethene is used in the manufacture of plastic polyethene. Ethene is heated under high pressure in the presence of a catalyst. Many ethene molecules join together to form a giant molecule called polyethene.
   (a) What is the name given to this reaction?
   (b) Describe how ethene molecules join together to form polyethene.
5. Joining together very many small molecules to make one macromolecule is known as _____.

Topic 7 Water

Success criteria

After studying this topic, you should be able to:

(a) Identify natural sources of water.
(b) Explain water cycle.
(c) Describe the physical properties of water.
(d) Explain the importance of water.
(e) Describe water hardness and its effects.
(f) Describe methods used to remove water hardness.
(g) Describe sources, effects and prevention of water pollution.

Introduction

Water is the most abundant substance on earth, thus it is considered essential for life. It is necessary for maintaining life processes. Huge quantities of water are used all over the world everyday for domestic consumption, transport, generating hydro-electric power and as a solvent in many industrial processes. Water can be obtained by chemical reactions or naturally.

Remember!

Water is life. We should protect our water sources at all costs and avoid polluting water.

7.1 Natural sources of water

Water occurs on the earth's surface from varying sources. The major natural sources of water can be categorised as: underground, surface and atmospheric water sources.

(a) Underground water sources

These sources of water are found in certain zones called **water tables**. They include;

(i) Spring water-: This kind of water comes from the impervious layers of the rock.

(ii) Wells-: These are dug out holes which supply water.

(iii) Boreholes-: These are deep holes drilled into the ground using machines. They require submersible pumps to lift the water.

(b) Surface water sources

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

(c) Atmospheric water source

An example of this kind of water source is rainwater collected from rooftops during rainy seasons using gutters and directed into storage tanks.

Which of these sources of water do you use in your community?

7.2 Water cycle

Water cycle is also known as **hydrological cycle**. It refers to the continuous circulation of water between the earth's surface and the atmosphere. Water cycle involves the following processes: heating, transpiration, expansion and cooling, condensation, precipitation, infiltration or percolation and surface runoff. The figure below shows how these processes inter-relate in a water cycle.

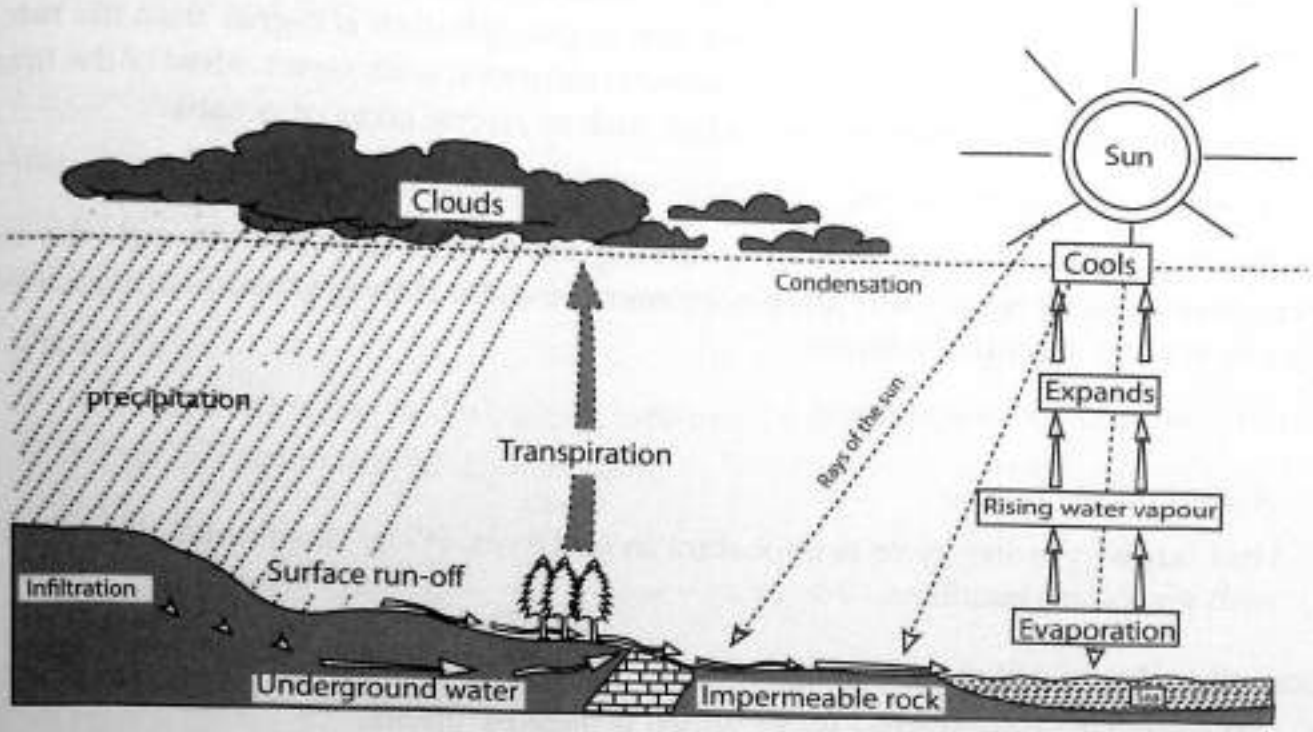


Fig. 7.1: Some of the processes involved in water cycle

(a) Evaporation and transpiration

When the water bodies are heated during the day with the heat from the Sun, water vapour rises in the atmosphere through the process of **evaporation**. Evaporation is the change of liquid water to water vapour from open water sources such as ocean and lakes into the atmosphere. Water vapour is also added to the atmosphere from green plants through the process of **transpiration**. You already learnt in Biology that transpiration is the process by which plants lose excess water to the atmosphere in form of water vapour. Sometimes, the two process are combined to give **evapotranspiration**.

(b) Expansion and cooling

As water vapour rises in the atmosphere, it **cools** due to the cool temperatures in the atmosphere. It also undergoes expansion as well. When it reaches certain heights where it is very cold, the water vapour begin to form clouds.

(c) Condensation

This is the process by which water vapour is changed into a liquid as a result of cooling. It happens high up in the atmosphere where the temperatures are very low.

(d) Precipitation

Precipitation involves accumulation of clouds in the sky for many hours. This activity causes the clouds to become heavy. The heavy clouds release moisture back to the earth's surface in form of rain, hailstorm, snow or fog. This process is generally referred to as **precipitation**.

(e) Surface run-off

This refers to the water that runs on the surface of the earth especially after it has rained. Surface run-off occurs when the rate of precipitation is higher than the rate of infiltration. It can also occur because the soil is saturated with water. Most of the times, surface run-offs end in other water bodies such as rivers, lakes or oceans.

(f) Infiltration or percolation

Infiltration is the physical process involving movement of water through the soil. **Percolation** on the other hand is the movement of water through the soil and its layers by gravity and cappilarity forces.

Significance of the water cycle

Research Activity

Find out why water cycle is important in our lives. Write short notes then share with your class members.

Some benefits of water cycle include:

1. It provides underground water which is used by plants.
2. It provides moisture in the atmosphere through the process of evapotranspiration. The moisture controls other factors such as humidity and temperature.
3. It ensures constant supply of fresh water in lakes and rivers. The water is used by animals, human being and aquatic plants.
4. It maintains a balance between the water on the surface, underground and in the atmosphere.
5. It encourages soil formation through leaching process.

Self assessment exercise 7.1

1. What is "water cycle".
2. Describe the process of water cycle.
3. State three sources of underground water.
4. Explain four ways in which the water cycle is of significance to us.

7.3 Physical properties of pure water

Some of the properties of pure water include:

1. Pure water is a colourless, tasteless, odourless liquid. Tap water is however tasty due to dissolved salts.
2. It boils at 100°C at 1 atmosphere pressure and it freezes at 0°C.
3. Its density is 1.00g/cm³. However, frozen water is less dense than liquid water. It floats on liquid water.

Experiment 7.1

***Aim**: To investigate the boiling point of pure water*

Apparatus and reagents

Boiling tube, thermometer, Bunsen burner and pure or distilled water.

Procedure

1. Add distilled water to the boiling tube until it is halfway.
2. Set up the apparatus as shown below.

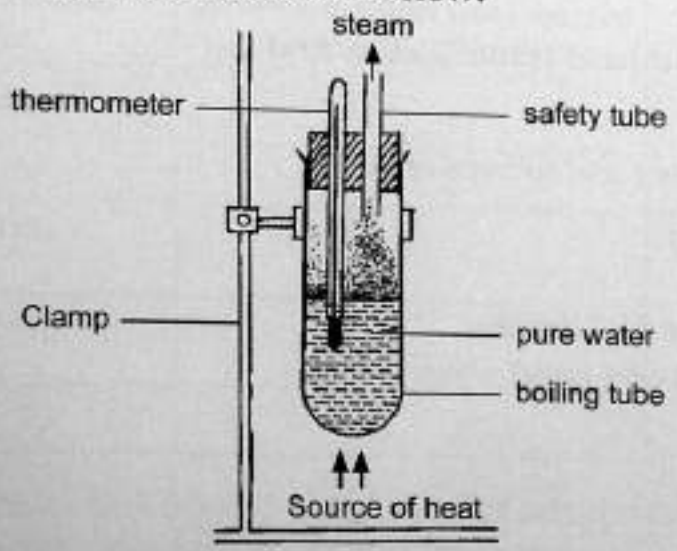


Figure 7.2: Investigating the boiling point of water

3. Heat the water until it boils and note the highest temperature reached in the thermometer.

The boiling point of water is measured by placing thermometer inside the boiling water as shown above. Water without impunities (pure or distilled water) boils at 100°C.

7.4 Importance of water

Research Activity

Find out about the importance of water to:
(a) Plants
(b) Animals
(c) Human beings
Write a report and present to the rest of the class

Water plays an important role in most aspects of our lives as shown below.

1. As a **coolant** – in simple distillation, it facilitates liquefaction process, while in automobiles and industries it is used to maintain low temperatures in engines.
2. As a **solvent** – Water is considered a universal solvent because it dissolves many substances in the human body and in industrial processes.
3. Generation of **hydroelectric power** generation – moving water is used to turn huge turbines that generate electricity at hydro electric power stations.
4. **Domestic** uses – water is used at home for cooking, washing, drinking and irrigating farms.
5. As a **reactant** – water is used in many chemical reactions which occur in aqueous media.
6. **Recreational** purposes – clean water is used in facilities such as swimming pools and for surfing.
7. **Transport** medium – ships, ferries, boats and canoes move on water in the process of transporting people and goods.

7.5 Water hardness

What is soft water? How about hard water? Let us find out

Experiment 7.2

Aim: *To investigate the hardness and softness of water*

Apparatus and reagents

- Distilled water
- Tap, river, well, pond and lake water
- Boiled tap or river, well and pond water
- Rain water
- Boiled rain water
- Sea water
- Boiled sea water
- Water to which magnesium sulphate has been added
- Boiled magnesium sulphate solution
- Water to which calcium chloride has been added

- Aqueous calcium hydrogen carbonate
- Boiled solution previously containing calcium hydrogen carbonate
- Distilled water containing calcium carbonate powder
- Sodium chloride solution
- Aqueous calcium hydroxide
- Soap
- 250 cm^3 conical flasks
- Burette

Note: Calcium hydrogen carbonate solution can be prepared by passing carbon dioxide into aqueous calcium hydroxide for a long time (until the precipitate formed dissolves).

Procedure

1. Dissolve 20g of soap in hot distilled water. Add water to make up one litre of soap solution.
2. Put 25 cm^3 of distilled water in a conical flask. Add about 2cm^3 of soap solution from a burette and shake the mixture. Does lather form?
3. If lather does not form in the conical flask, add the soap solution drop by drop, shaking it after every addition until a permanent lather (lasting for about 2 minutes without breaking) is formed.
4. Copy Table 7.1 in your notebook and record the volume of soap solution used in each case.

Table 7.1: Action of soap on soft and hard water

Sample solution/ water	Volume of soap used before it is boiled (cm^3)	Volume of soap used after it is boiled (cm^3)
Distilled water		
Tap water		
Calcium chloride solution		
Magnesium sulphate solution		
Rain water		

5. Repeat procedure 2 and 3 with 25 cm^3 of calcium chloride solution, magnesium sulphate solution, rain water and tap water.
6. Repeat the same procedure using the other solutions and the suspension of calcium carbonate. Record the volumes of soap solution used in a table like the one shown in Table 7.1.

Study questions

1. Which of your sample solutions lather readily and which ones lather least with soap?
2. Which cations are present in solutions that formed lather easily with the soap solution?
3. Which cations are present in solutions that did not form lather easily with the soap solution?
4. Compare the amount of soap used when the sample solution is not boiled and when boiled.
5. List the sample solutions that lather readily when they are not boiled and when boiled.
6. List the sample solutions that lather readily when boiled.
7. List the sample solutions that do not lather readily even when boiled.
8. What happens when a solution of calcium hydrogen carbonate is boiled?
9. What happens when a solution of magnesium sulphate is boiled?

Water which forms lather easily with soap is said to be **soft**. Examples of soft water include **distilled water** and **rainwater**. The water which does not form lather easily is said to be **hard**. An example of hard water is **tap water**. From experiment 7.2 above distilled and rain water form lather very easily with soap. Tap, well, river, lake and sea water will vary in the amount of soap used for them to form lather depending on the extent of hardness. Sodium chloride solution also produces lather easily. Solutions containing calcium and magnesium ions did not produce lather easily. In these solutions which did not lather easily, there was the formation of grey–white suspension. This suspension is known as **scum**. You must also have noted that distilled water containing calcium carbonate powder lather very easily. Calcium carbonate is insoluble in water and therefore it does not form calcium ions.

From the observations above, distilled water or pure water contains no dissolved solid impurities. It lathers easily with soap. Rain water lathers almost the same as distilled water. The water that readily forms lather with soap is known as **soft water**. **Hard water** does not readily form lather with soap.

When calcium hydrogen carbonate solution is boiled, it decomposed into calcium carbonate which precipitates from the solution. The water which is formed lathers easily with soap.

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

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

When magnesium sulphate solution is heated, the resulting hot solution does not lather easily with soap. This is because magnesium sulphate does not decompose on heating. Thus we can therefore conclude that **calcium ions** and **magnesium ions** cause hardness in water. Hence there are two types of water hardness:

- Temporary hardness
- Permanent hardness

(a) Temporary hardness

Water contains dissolved carbon dioxide. Part of this gas forms carbonic acid.

Carbon dioxide + water ⟶ carbonic acid

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

When this water passes over limestone (calcium carbonate) or dolomite (magnesium carbonate), **calcium hydrogen carbonate** and **magnesium hydrogen carbonate** are formed respectively.

Calcium carbonate + carbonic acid ⟶ calcium hydrogen carbonate

$$CaCO_3(s) + H_2CO_3(aq) \longrightarrow Ca(HCO_3)_2(aq)$$

Magnesium carbonate + carbonic acid ⟶ magnesium hydrogen carbonate

$$MgCO_3(s) + H_2CO_3(aq) \longrightarrow Mg(HCO_3)_2(aq)$$

Temporary hardness in water is caused by the presence of dissolved calcium hydrogen carbonate and/or magnesium hydrogen carbonate. It is called temporary hardness because it can be removed by **boiling**.

(b) Permanent hardness

Permanent hardness in water is caused by dissolved calcium sulphate and magnesium sulphate. This occurs when water dissolves small amounts of calcium and magnesium sulphate from ground rocks as it passes on its way to reservoirs, lakes, rivers or to the sea. This type of hardness cannot be removed by boiling. It can however be removed through **chemical treatment** or **distillation**.

Note: Methods used to remove permanent hardness in water can also be applied in removing temporary hardness.

Effects of water hardness

As we mentioned earlier, hard water does not lather easily with soap. Instead, a grey-white suspension or **scum** is formed in the water.

The main constituent of common soap is a soluble sodium salt known as **sodium octadecanoate**. When soap dissolves in water, its octadecanoate ion is produced.

$$\text{Sodium octadecanoate} \xrightarrow{\text{water}} \text{octadecanoate ion} + \text{sodium ion}$$

The octadecanoate ions combine with calcium or magnesium ions from the hard water to form an insoluble salt which is called **scum**.

$$\text{Octadecanoate ion} + \text{calcium ion} \longrightarrow \text{calcium octadecanoate (scum)}$$

$$\text{Octadecanoate ion} + \text{magnesium ion} \longrightarrow \text{magnesium octadecanoate (scum)}$$

The scum often stains clothes when the clothes are washed using hard water from rivers, lakes or ocean.

Advantages and disadvantages of hard water

Hard water has its advantages and disadvantages as shown below.

(a)Advantages of hard water

1. Calcium compounds present in hard water is good for development of healthy teeth and bones.
2. If conveyed in pipes of lead, it helps to form lead carbonate coat in the water. This prevents lead from dissolving in water. **Lead is a poisonous metal.**
3. Hard water is good for formation of shells in animals. Snails obtain calcium from hard water.
4. Calcium and magnesium contained in hard water are necessary for healthy growth of both plant and animals.

(b) Disadvantages of hard water

1. Washing using hard water is accompanied with wastage of soap due to scum formation.
2. Calcium and magnesium carbonates exist as 'fur' on electrical appliances, for example, kettle and boiler thereby making them inefficient as their conductivity is reduced.
3. Scum leaves dirty marks on clothes after washing. This act lowers the quality of the fibre.
4. Calcium carbonate and magnesium carbonates formed in hard water blocks pipes. These deposits reduces heat conductivity.

Self assessment exercise 7.2

1. Differentiate between hard water and soft water.
2. Define the following terms
   (a) Temporary hardness of water.
   (b) Permanent hardness of water.
3. Name four salts that cause water hardness.

7.6 Methods of removing water hardness

The method used to remove water hardness depends on whether the hardness is permanent or temporary.

(a) Removing temporary hardness in water

Temporary water hardness can be removed by:

(i) Boiling

Boiling decomposes the soluble hydrogen carbonates to insoluble carbonates, carbon dioxide and water as shown in the equations below.

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

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

Magnesium hydrogen carbonate $\xrightarrow{\text{boiling}}$ magnesium carbonate + water + carbon dioxide

$$Mg(HCO_3)_2(aq) \xrightarrow{\text{boiling}} MgCO_3(s) + H_2O(l) + CO_2(g)$$

The insoluble carbonates formed eliminate the calcium and magnesium ions from the water hence the water becomes soft.

(ii) Addition of calcium hydroxide or sodium carbonate

The method of removing temporary water hardness by boiling is too expensive to be used on a large scale. Commercially, temporary hardness can be removed by adding calculated amounts of calcium hydroxide or sodium carbonate powder. This causes precipitation of calcium and magnesium ions inform of carbonates as shown below.

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

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

Calcium hydrogen carbonate + sodium carbonate ⟶ calcium carbonate + sodium hydrogen carbonate

$Ca(HCO_3)_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaHCO_3(aq)$

Note: Sodium ions from sodium hydrogen carbonate do not cause water hardness.

Self assessment exercise 7.3

1. Write the chemical equation for the reaction which takes place when magnesium hydrogen carbonate is removed from hard water using calcium hydroxide solution and also when it is removed by sodium carbonate solution.
2. What will happen if excess calcium hydroxide solution is used in the removal of temporary water hardness?

(b) Removing permanent hardness in water

Permanent hardness can be removed by:

(i) Addition of sodium carbonate

Addition of sodium carbonate causes precipitation of calcium and magnesium ions as insoluble carbonates. The equations of the reactions are shown below.

Calcium sulphate + sodium carbonate ⟶ calcium carbonate + sodium sulphate

$CaSO_4(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + Na_2SO_4(aq)$

Magnesium sulphate + sodium carbonate ⟶ magnesium carbonate + sodium sulphate

$MgSO_4(aq) + Na_2CO_3(aq) \longrightarrow MgCO_3(s) + Na_2SO_4(aq)$

(ii) Distillation

This method removes both permanent and temporary hardness because it removes all the dissolved ions. However, it is expensive to use it to soften large amounts of water. It can, also be used in the laboratory to make distilled water for analytical work or for batteries.

(iii) Use of ion-exchange method

This method involves use of a **resin**. Resin is made of an insoluble complex of a sodium salt known as **sodium permutit**. For this reason, this method is also referred to as the **permutit process**. Permutit releases sodium (Na^+) ions which exchange with the calcium and magnesium ions in the hard water. Sodium ions $Na^+(aq)$ in water do not cause water hardness. Calcium and magnesium ions now become part of the insoluble complex compound.

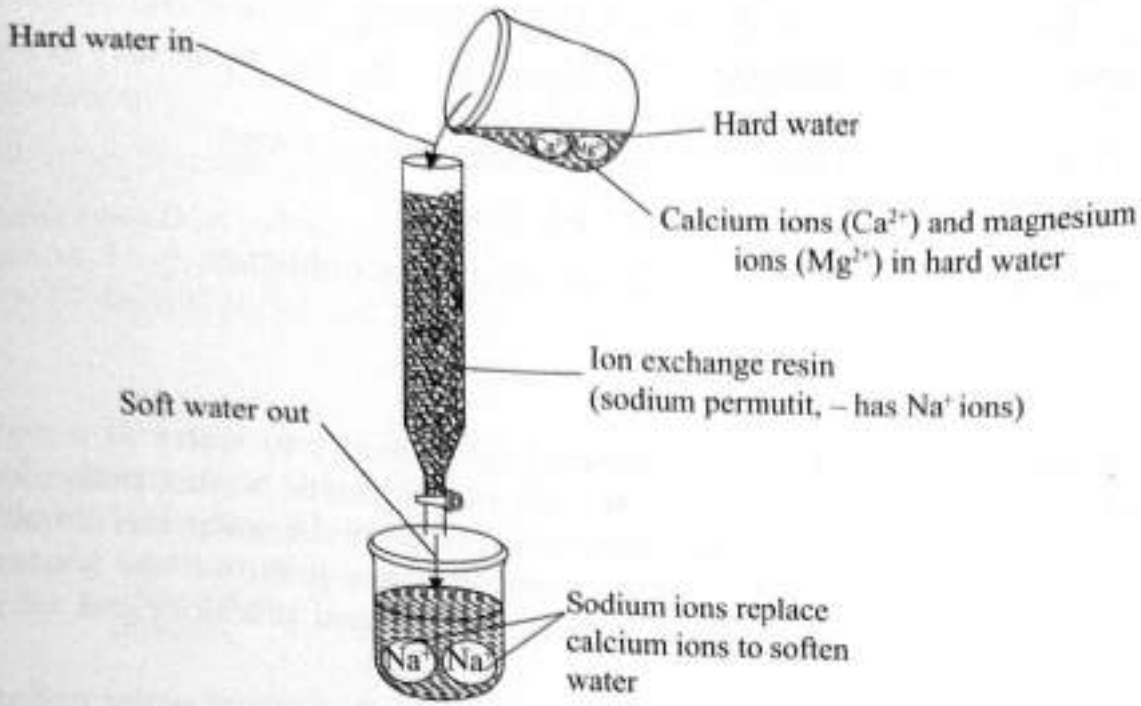


Fig. 7.3: Permutit process

Note: When all the sodium (Na^+) ions on the permutit column have been exchanged with calcium or magnesium ions; the resin containing ions of calcium and magnesium is now said to be **discharged**. The permutit can be **re-generated** or **recharged** by passing concentrated sodium chloride solution (brine) through it. Another ion–exchange process takes place.

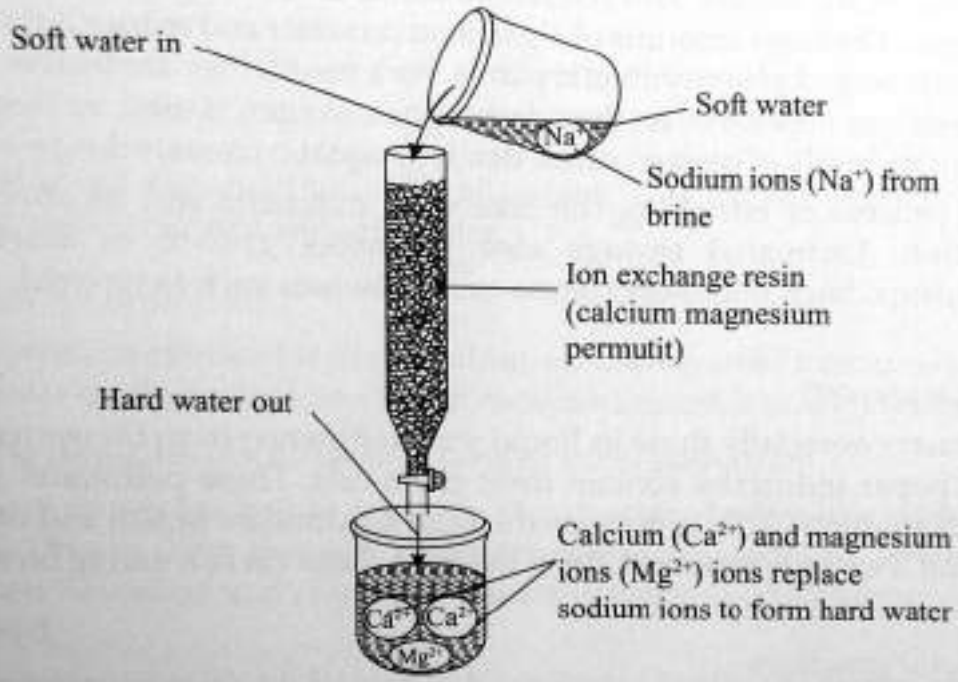


Fig. 7.4: Ion-exchange column during re-generation

Calcium or magnesium ions are washed away leaving the permutit full of sodium ions (sodium permutit) and it is ready for re-use.

Self assessment exercise 7.4

1. State two advantages and two disadvantages of hard water.
2. How does permutit process soften hard water?
3. Explain ways in which hard water is a nuisance in industries.

7.7 Water pollution

Water pollution refers to addition of harmful substances into water as a result of human activities. These substances include toxic metals, plastic wastes, pesticides and fertilisers. These substances may kill aquatic organisms, make water less suitable for drinking or for use in agriculture. Other causes of water pollution are wastes and sewage from homes, industrial effluents, poor agricultural practices and oil spills from ships.

In the next section we will discuss the various sources and effects of water pollution.

(a) Wastes and sewage from homes

The release of organic wastes such as human faeces and sewage into water bodies such as lakes and rivers causes water pollution. The decomposition of these wastes by micro organisms uses up oxygen in water. As a result, fish and other aquatic organisms die due to lack of oxygen.

Decomposition of the wastes also release nutrients into the water which cause rapid growth of algae. The large amounts of algae float on water and reduce light penetration to the fully submerged photosynthetic plants. As a result, they are unable to carry out photosynthesis and hence die. As they decompose, oxygen is used up from the water. Also, inadequate levels of oxygen cause death of aquatic animals due to suffocation.

This whole process of enriching the lake with nutrients and its effects is called **eutrophication**. Untreated sewage also promotes growth of disease-causing micro-organisms. Such micro-organisms cause diseases such as typhoid, cholera and dysentery.

(b) Industrial wastes

Industrial wastes especially those in liquid state (effluents) from breweries, tanneries, textiles and paper industries contain toxic chemicals. These pollutants are harmful even in small amounts. This is because they can accumulate in fish and other aquatic organisms and then be transferred along the food chain via fish-eating birds to human beings.

(c) Agricultural practices

Pesticides and excess fertilisers applied to crops may enter into rivers and lakes through run off after rains. Some pesticides contain toxic metals such as copper. Excessive use of fertilisers, for example, nitrates promotes **eutrophication**, a process which causes death of aquatic organisms.

(e) Detergents from washing in homes and hospitals

Detergents release phosphates into water. Phosphates cause frothing and promote eutrophication.

(f) Hot water from industries

Hot water results in reduction of oxygen from water. This leads to death of fish due to suffocation. High temperatures contributed by hot water denature enzymes involved in many biological processes. This causes death of many acquatic organisms.

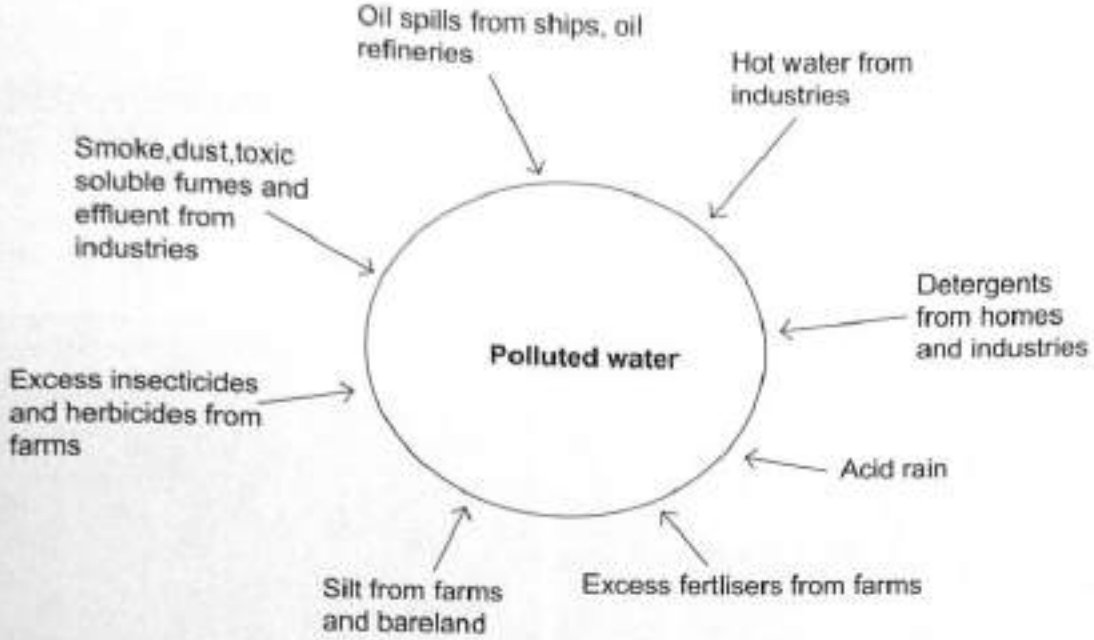


Fig. 7.5: Sources of water pollution

Remember

We should avoid water pollution by all means. It is the responsibility of every Malawian to ensure clean and safe water.

Control of water pollution

Due to the enormous effects of water pollution, effective control measures are required. These measures can help eliminate health problems associated with water pollution.

The following measures can be used to control water pollution:

1. Treatment of sewage before releasing it into water bodies like lakes, rivers and oceans. This is done through water treatment in plants in which pollutants contained in sewage and any other disease-causing micro-organisms present are destroyed.
2. Avoiding excessive use of chemicals, for example, pesticides in pest control. Adopt biological methods of controlling pests. This practice can reduce accumulation of harmful chemicals in the environment.
3. Encourage use of farm-yard manure instead of artificial fertilisers. This method of farming can reduce the amount of toxic chemicals finding their way into water bodies.

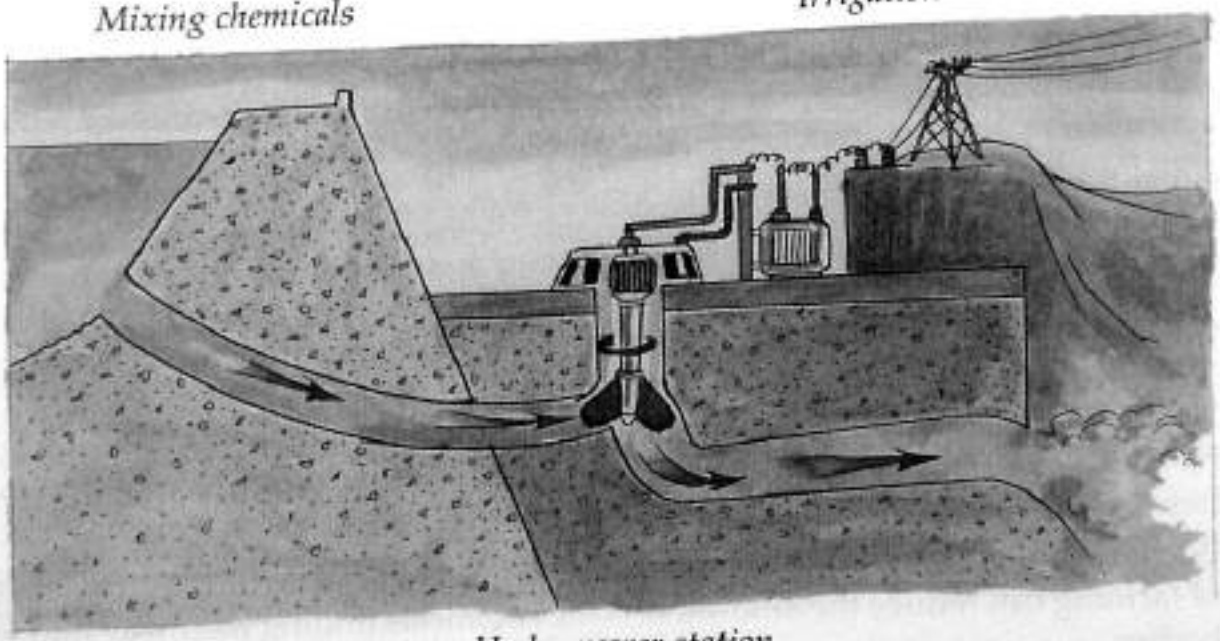
4. Pit latrines should be constructed away from water wells to avoid contamination of water sources with pollutants associated with sewage material. The slope of the land should also be considered in putting up a latrine to avoid drainage of sewage into water sources.
5. Proper legislation should be put in place by governments to reduce irresponsible damping of industrial wastes and other domestic wastes into water bodies.
6. Petroleum companies should put measures in place to avoid unnecessary accidents which usually result into spillage of oils into oceans and seas.

7.8 Water Quality Standards

We already learnt that water is important is sustaining life. Can you list some ways in which water is important to us? Now look at these pictures. Discuss what is happening in the pictures with your friend.

Mixing chemicals

Irrigation



Hydro power station

About 80% of Malawian households get water for domestic use from open shallow wells. Studies have shown that the waters from these wells are in most cases unfit for human consumption. We already know that in order to stay healthy and to avoid diseases, we should use safe and good quality drinking water. How do we know that drinking water is safe?

Look at the table below. It shows water quality guidelines from Malawi Bureau of standards (MBS) and the world health organization (WHO).

Table of water quality guidelines from Malawi Bureau of Standards (MBS) and World Health Organization (WHO)

Item	Ranges MBS	WHO guideline
(a) Physical qualities		
Conductivity (at 25°C)	70 - 150 mS/m	No data
pH value (at 25°C)	5.0 - 9.5	Not of health concern at levels found in drinking-water
(b) Macro element levels		
Calcium	80 - 150 mg/l	No data
Chloride	100 - 200 mg/l	Not of health concern at levels found in drinking-water
Fluoride	0.7 - 1.0 mg/l	1.5 mg/1
Magnesium	30 - 70 mg/1	No data
Nitrate and nitrite	6.0 - 10.0 mg/1	50 mg/1
Potassium	25 - 50 mg/1	Occurs in drinking-water at concentrations well below those of health concern
Sodium	100 - 200 mg/1	Not of health concern at levels found in drinking-water
Sulphate	200 - 400 mg/1	Not of health concern at levels found in drinking-water
Zinc	3.0 - 5.0 mg/1	Not of health concern at levels found in drinking-water
(c) Micro element levels		
Aluminium	0.150 - 0.300 mg/1	A health-based value of 0.9 mg/1 could be derived, but this value exceeds practicable levels based on optimization of the coagulation process in drinking-water plants using aluminium based coagulants: 0.1 mg/1 or less in large water treatment facilities and 0.2 mg/1 or less in small facilities
Cadmium	0.003 - 0.005 mg/1	0.003 mg/l
Chromium	0.050 - 0.100 mg/1	0.05 mg/1

Copper	0.500 - 1.000 mg/1	2 mg/1
Iron	0.010 - 0.200 mg/1	Not of health concern at levels causing acceptability problems in drinking-water
Lead	0.010 - 0.050 mg/1	0.1 mg/1
Manganese	0.050 - 0.100 mg/1	Not of health concern at levels causing acceptability problems in drinking-water

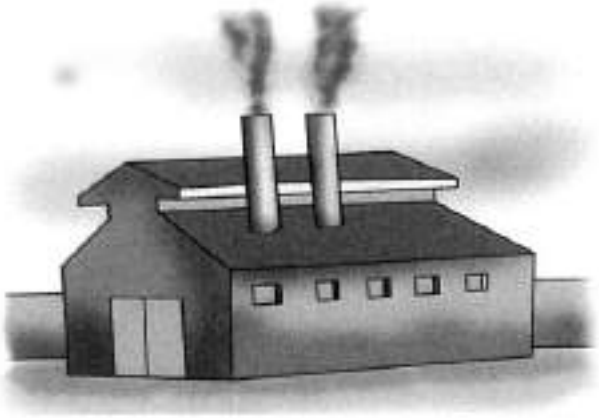
We should always ensure that the water that we drink adhere to these safety standards in order for us to remain healthy.

NOTE: Malawi Bureau of Standards has set guidelines for several elements that the world health organization considers to have no health effects or to be of health concern.

7.9 Impact of human activities on the atmosphere

Human beings engage in a variety of activities in the course of their lives. Can you name some of these activities?

Now, look at the following pictures. What is going on in the pictures? Are the activities good or bad for the environment? Why?



Factory producing smoke

Deforestation

Remember!

We should always avoid or minimize activities that cause pollution of any form in the environment!

Some of the activities above cause air pollution. Just like water pollution, air pollution refers to addition of harmful substances into the atmosphere. These substances are called **pollutants**. The pollutants contaminate the air thereby interfering with its quality. The most common air pollutants in Malawi include:

- Smoke and fumes from burning tyres, plastics and other forms of garbage.
- Burning fossil fuels for example from car exhausts and fumes from other automobiles.
- Emission of poisonous gases such as carbon dioxide, carbon monoxide and sulphur dioxide from industries.
- Dust arising from mining activities. For example, from quarrying and construction sites.
- Aerosol droplets from farm or household chemicals such as herbicides, insecticides and pesticides. Such droplets may contain mercury, chlorine or other chemicals which are harmful to health.

The presence of the air pollutants cause:

- Breathing difficulties or respiratory problems such as bronchitis, emphysema and asthma.
- Depletion of the ozone layer, which eventually leads to global warming or green house effect.
- Acid rains, which kill animals and destroy structures made of limestone and steel.
- Reduce visibility leading to air and road accidents

Activity

Your teacher will organise an academic trip to a polluted town. While in the town:

1. Look keenly around as you move. Can you see any smoke or dust particles. What causes them?
2. Smell around as you move. Can you feel any smell? What causes the smell?

Discuss with your friend the answers to the questions. What should we do to avoid such situations?

Revision Exercise 7

1. Name three natural sources of water.
2. Discuss the importance of water.
3. (a) What is meant by hard water?
   (b) What is the difference between hard water and soft water?
4. State the two types of water hardness.
5. Describe how the following methods remove water hardness.
   (a) Boiling
   (b) Addition of calcium hydroxide
   (c) Addition of sodium carbonate
6. Explain briefly the causes and effects of water pollution.

Topic 8 Greenhouse gases and ozone layer

Success criteria

After studying this topic, you should be able to:

(a) Explain the effects of natural disasters and human activities on the atmosphere.
(b) Explain the effects of greenhouse gases.
(c) Describe ways of mitigating the effects of global warming.
(d) Describe the importance of the ozone layer.
(e) Explain the depletion of ozone layer and its associated problems.

Introduction

Various natural disasters, for example, flooding and human activities such as burning of fossil fuels and deforestation have increased the abundance of gaseous particles in the atmosphere. These particles can change the amount of energy that is absorbed and reflected by the atmosphere. They are also believed to modify the properties of clouds, changing the amount of energy that they can absorb and reflect among others.

8.1 Natural disasters and their effects on the atmosphere

Natural disasters are any natural occurrence that causes widespread distress, usually loss of human life and damage to property. Some of these disasters are drought, windstorm, earthquake, fire, floods, hurricanes, thunderstorms and volcanic eruptions. Let us discuss each of these disasters in a little more detail.

(a) Flooding

It is the submerging of a normally dry area with water. Floods are caused by many factors including rainstorms, broken dams, underwater, volcanic eruptions and earthquakes.

Floods can have a negative impact on trees if they are submerged for too long i.e. the trees will be prevented from taking in carbon dioxide. This may lead to accumulation of the gas to toxic levels in the atmosphere. Floods also destroy structures such as buildings, roads, bridges among others. It can also cause death of animals and human beings.

(b) Earthquakes

Although the major effects of earthquakes are usually destruction of structures and loss of life, earthquakes can also result in adverse environmental and atmospheric consequences. It can cause damage to vegetation as well as change of landscape due to formation of physical features such as volcanoes, craters and escarpments.

(c) Volcanic eruption

Volcanic eruptions produce carbon dioxide and other toxic gases such as sulphur dioxide and nitrogen dioxide which are then released into the atmosphere. Accumulation of these gases over long periods of time lead to greenhouse effect which may eventually case global warming.

(d) Drought and wildfires

Drought occurs when the weather is **hot** and **dry**. The dust is usually blown up by the wind, causing huge clouds known as **dust storms**. Because they can turn the sky dark, they are sometimes known as **blizzards**. Long periods of drought leads to death of animals and vegetation, drying of water sources and increased chances of forest fires. Forest fires have a devastating power to wipe away large acres of forest land. The dry needlike leaves of trees provide ideal fuel for the fires. Broad leaves of evergreen trees also speeds up the forest fires. Drought and wild fires lead to contamination of the atmosphere through the release of smoke and toxic gases.

8.2 Greenhouse gases

Greenhouse gases are air components whose accumulation in the atmosphere cause rise in global temperatures. Can you have some greenhouse gases that you know.

Examples of greenhouse gases include:

- Carbon dioxide
- Sulphur dioxide
- Nitrogen dioxide
- Nitrogen monoxide
- Methane

Sources of greenhouse gases

(a) Carbon dioxide

The processes that add carbon dioxide to the atmosphere include:

(i) **Combustion** – Burning of carbon – containing compounds in sufficient oxygen produces carbon dioxide and water.

(ii) **Respiration** – This involves oxidation of glucose and in the process, water is formed as carbon dioxide is given out.

(iii) **Fermentation process** – in which yeast changes glucose to ethanol and carbon dioxide.

(iv) **Decomposition** of plants and animals remains this results in production of carbon dioxide as a by product.

(v) **Industrial manufacture of calcium carbonate** – results in production of carbon dioxide as a by product i.e.

$$\text{Calcium carbonate} \longrightarrow \text{calcium oxide} + \text{carbon dioxide}$$

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

(b) Sulphur dioxide

Sulphur dioxide is mainly produced when fossils fuels are burnt. It is also released through metallurgical processes when metal ores are roasted.

Zinc sulphide + Oxygen ⟶ zinc oxide + sulphur dioxide

$$2ZnS\,(s) + 3O_2\,(g) \longrightarrow 2ZnO\,(s) + 2SO_2\,(g)$$

Lead sulphide + Oxygen ⟶ lead oxide + sulphur dioxide

$$2PbS\,(s) + 3O_2\,(g) \longrightarrow 2PbO\,(s) + 2SO_2\,(g)$$

(c) Nitrogen monoxide and nitrogen dioxide

These two gases are introduced into the atmosphere through:

(i) Natural processes, for example, **lightning** and biological processes that occur in plants and animals.

(ii) Man-made sources mostly combustion of fossil fuel in the industries and internal combustion engines.

At high temperatures inside a car engine,oxygen and nitrogen in the air react to produce nitrogen monoxide.

Nitrogen + oxygen ⟶ nitrogen monoxide

$$N_2(g) + O_2(g) \longrightarrow 2NO_2(g)$$

Nitrogen monoxide can further be oxidised to form nitrogen dioxide as follows:

Nitrogen monoxide + oxygen ⟶ Nitrogen dioxide

$$4NO(g) + 2O_2(g) \longrightarrow 4NO_2(g)$$

(d) Sources of methane

There are both natural and artificial sources of methane in the atmosphere

The main natural source include:

- Wetlands
- Oceans
- Activities of termites

These natural sources account for upto 36% of methane emissions. However, the most important source of methane (that accounts for over 64%) is related to human activities associated with fossil fuels. For example, cooking gas in gas cylinders where methane leaks during their production, transportation and refilling account for significant amount of emissions. Other emissions come from land fills, biomass burning as well as biofuels.

Effects of greenhouse gases

Research Activity

Find out with your friend the effects of greenhouse gases to the atmosphere and the environment in general. Write a report and share with class members.

Greenhouse gases have the following impacts:

(a) Cause acid rain.

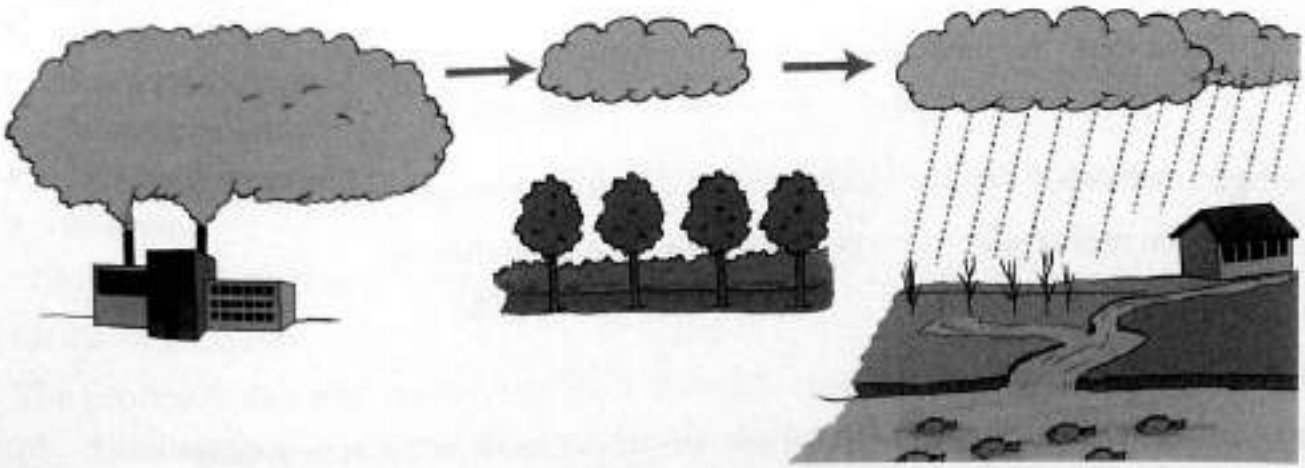
(b) Have adverse effects on human health.

(c) Cause greenhouse effect that lead to global warming.

(a) Acid rain

Acid rain is formed when large amounts of oxides of carbon, sulphur and nitrogen are emitted from factories and vehicle exhausts. The gases dissolve in rain droplets that fall as rain. Acid rain damages plants and kills acquatic life.

When gases, such as carbon dioxide, sulphur dioxide and nitrogen dioxide, are Emitted into the atmosphere, they dissolve in the water droplets when it rains to form acidic solutions that fall as acid rain.



Fi. 8.1 How acid rains form and its effects

As we already know, acid has a burning effect and is corrosive. Therefore, acid rain can damage both terrestrial and aquatic plants and it can also kill fish. It also causes corrosion of buildings made from limestone and corrugated iron sheets.

Other effects of the acid rain on the environment

(i) On human health

The pollutants that cause acid rain for example sulphur dioxide when inhaled causes heart and lung disorders such as asthma and bronchitis.

(ii) On useful materials

Acid rain contributes to corrosion of metals, destruction of buildings made from limestone among others. These effects seriously reduce the value of buildings, bridges, statues and even vehicles and other automobiles.

(iii) Aluminium compounds in the soil

Acid rain makes insoluble aluminium compounds in the soil soluble. Plants absorb the aluminium ions through their roots. Due to this absorption, the plant become weakened and die.

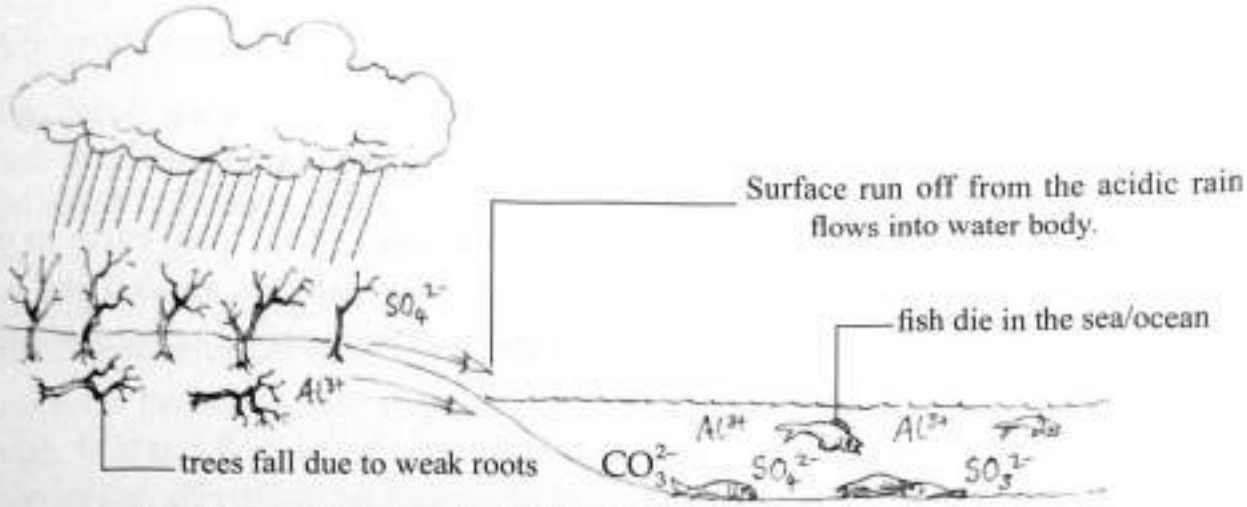


Fig 8.2: Effects of increased solubility of aluminium compounds due to acid rain

(b) Adverse effects on human health

Sulphur dioxide and other air pollutants can cause alveoli of the lungs to clump together. This reduces the surface area for gaseous exchange. When this happens body tissues are denied of oxygen and this may cause death.

Also, some air pollutants cause constriction in the branches of bronchial tubes. This reduces the rate at which air is exchanged between the alveoli and outside the lungs.

(c) Greenhouse effect

The earth gets its heat from the sun. Most of the heat that reaches the earth is reflected back to the outer space. The atmosphere around the earth acts as an **insulating layer**, absorbing in some of this heat. This is how the earth is able to maintain its temperature. However, some gases like carbon dioxide and methane are very good at trapping or insulating most of this heat. The burning of fossil fuels worldwide produces and releases large amounts of carbon dioxide into the atmosphere. This carbon dioxide would normally be used up in the process of photosynthesis by trees that make up large rainforests. Due to deforestation in many parts of the world, less photosynthesis takes place.

The increased concentration of carbon dioxide and methane traps most of the radiant energy of the sun in the same way a **greenhouse** does hence the term **greenhouse effect.**

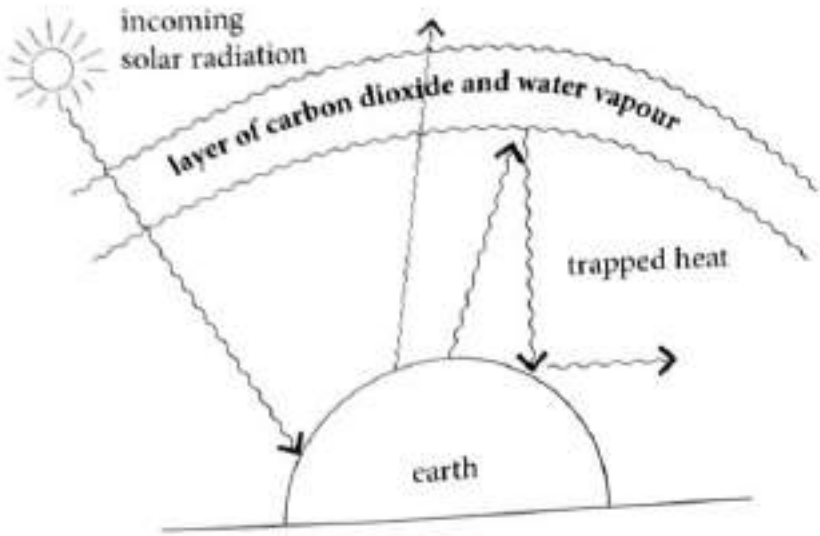


Fig. 8.3: The green house effect

This raises global temperatures by slowing down the loss of heat from earth to outer space. Increased global temperatures can affect weather patterns around the earth. For example, it may result in the polar ice caps melting. Such activity raises sea levels leading to floods in low-lying coastal regions.

Experiment 8.1

Aim: *To demonstrate the working of a green house.*

Apparatus and reagents

- A large glass beaker or glass box
- Two small plastic beakers full of water
- Thermometer

Procedure

1. Place a small plastic beaker full of water under a big glass beaker or a glass box. Place another in the open as shown in the figure below.

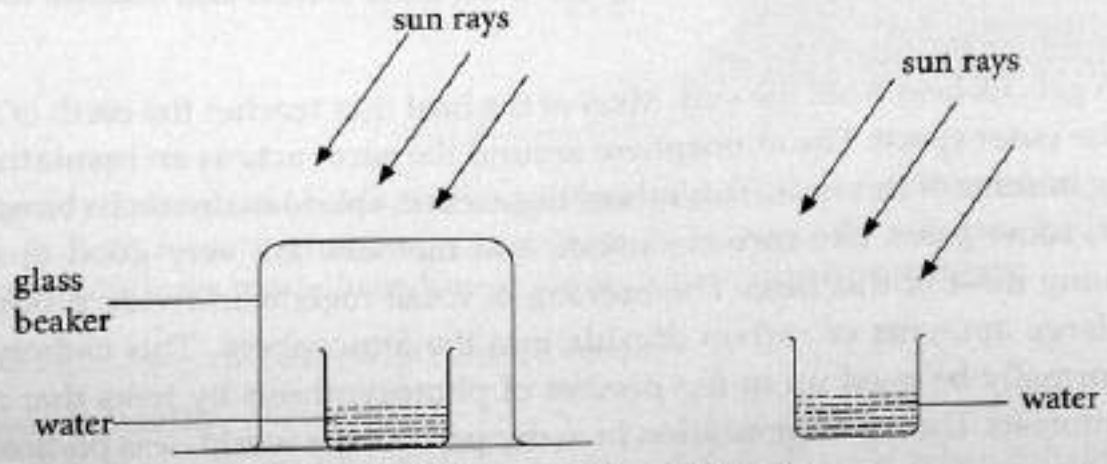


Fig. 8.4: Demonstrating greenhouse

2. Leave the set-up in the sun for six hours.
3. Measure the temperature of the water in the two beakers.
   Which one has a higher temperature? Why?

The temperature of the water inside the glass beaker or box is higher. This is because the heat from the sun passes through the glass and is largely trapped within the glass and does not escape back to the atmosphere. This is the heat which warms the water inside the glass box to higher temperatures than the water in the uncovered beaker.

8.3 Air quality standards

Studies have shown that air quality in Malawi is still good but future air quality related problems should not be ignored. These include impact on human health, global warming and ozone depletion.

Therefore, there is need for regular checking of the quality of air by use of certain standards. The air quality standard in use at the moment are based on the World Health Orgnisation air quality standards for air pollutants.

In line with World Health Organisation WHO guidelines, the Malawi Government through Malawi Bureau of standards (MBS) published in the year 2005 guidelines on maximum allowable emissions of both gaseons and partuculate polluntant in a document called Malawi Standard MS 737:2005.

The guidelines addition propose mechanism to reduce emissions from statutory and mobile sources. It also gives types of mitigation measures and best available technique (BATs) that can be used especially in industrial emissions scenarios.

Table 8.1 and 8.2 highlights these specifications for selected pollutants.

Table 8.1: Ambient air quality standards limits in Malawi

Pollutant	Maximum concentration in ambient air	Averaging period
Suspended pariculate matter	25 μg/m³	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	0.12 ppm	1 hour
	0.3 ppm	1 year
Ozone	0.12 ppm	1 hour
Lead,μg/m³	0.50 μg/m³	1 year

Photochemical oxidants	0.10 ppm	1 hour
	0.08 ppm	4 hour

Table 8.2: Air quality standard for motor vehicles in Malawi

Pollutant	Maximum allowable concentration	Duration	Typical high concentration for specific pollutant, ppm
Hydrocarbons	160 μ/gm³	3 hours	6
Carbon Momoxide	10 μ/gm³	8 hours	40
Nitrogen dioxide	100 μ/gm³	1 year	0.1
Oxidants	160 μ/gm³	1 hour	0.6

ppm – means parts per million

μg/m³ – means concentration in micrograms per unit volume

8.4 Global warming

Global warming is the continuing rise in the average temperature of the earth due to cumulative effect of the greenhouse gases in the atmosphere. Refer to section 8.3 for more details on greenhouse effect.

Effects of global warming

Global warming has numerous effects on climatic conditions. Some of these effects are:

(a) **Melting of mountain glaciers and polar ice.** – This is evident in the Tundra climatic region. It was once covered with thick permafrost but the ice has melted with rising surface temperatures leading to flooding.

(b) **Sub-merging of islands** and coastlines due to rising sea level. High temperatures mean sea ice and polar ice sheets will tend to melt. This increases the amount of water in the world seas and oceans.

(c) **Expansion of deserts** – As the climate warms, drought conditions may increase. This leads to shrinking of water bodies and eventually desertification comes in

(d) Formation of **hurricanes** and **typhoons** – studies have shown that there is a relationship between increase in ocean and the occurance of hurricanes and typhoons..

Ways of reducing the effects of global warming

The effects of global warming can be reduced through:

1. Use of **renewable sources** of energy such as solar energy, wind power and hydro electric power.

2. **Afforestation** – planting of trees in lands where none existed before for example in fallow lands and in reclaimed lands.This will reduce carbon dioxide levels in the atmosphere.
3. Factories that produce carbon dioxide and sulphur dioxide can be fitted with chemicals called **acid scrubbers**. These chemicals absorb acidic gases from fumes before they are released into the atmosphere.
4. Automobiles can be fitted with **catalytic converters** to clean up exhaust gases before they are released into the atmosphere.
5. **Reafforestation** – planting of trees to replace the cut down ones and controlled felling of tropical forests can help maintain carbon dioxide at moderate levels in the atmosphere.

Self assessment exercise 8.1

1. Describe two effects of acid rain on the environment.
2. Explain how acid rain can be prevented.
3. Define the term global warming.
4. Briefly explain ways through which global warming can be minimised.
5. Explain how global warming leads to flooding and submerging of islands.

8.5 Ozone layer

Ozone layer refers to a region of the earth's stratosphere that absorbs most of the sun's ultraviolet radiation. Stratosphere contains high concentration of ozone gas in relation to other parts of the atmosphere. It is mainly found in the lower parts of the stratosphere from approximately 20km to 30km above the earth. However, the levels of ozone in the stratosphere varies seasonally and geographically.

Importance of the ozone layer

The ozone layer absorbs 93% to 99% of the Sun's high frequency ultraviolet (uv) light. Ultraviolet light is potentially damaging to life on earth. The exposure to uv rays may causes:

- Skin cancer.
- Damage to crops.
- Harm to aquatic life.

It is therefore important that the levels of ozone layer in the atmospere is not interfered with.

Depletion of the ozone layer

Substances that destroy the ozone layer are mainly gases from the chlorofluorocarbons (CFCs). Things that release CFCs include aerosols, coolants and foams.

Fig. 8.5: Spraying mosquitoes with CFC insecticides contributes to depletion of ozone layer

When CFCs get to the atmosphere, the chlorine atom (Cl) is released from the CFC molecule. The chlorine atom split the ozone to oxygen atoms and oxygen molecules. The free oxygen atom them combines with ozone molecule to form another oxygen molecule i.e. $O_3(g) \longrightarrow O_2(g) + [O]$

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

This results in depletion of the ozone layer

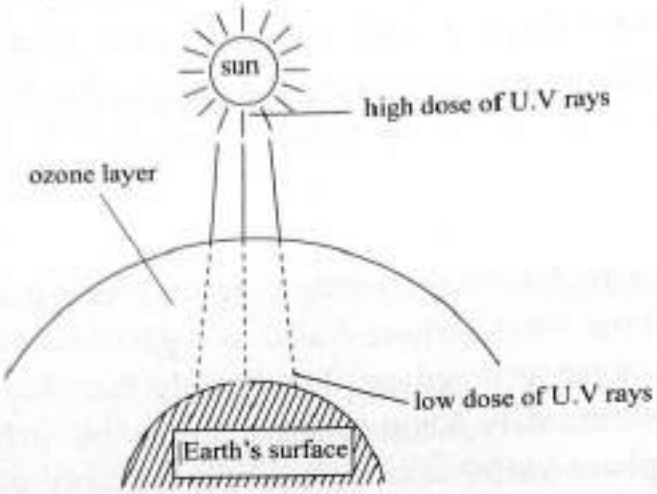


Fig 8.6(a): Undepleted ozone layer

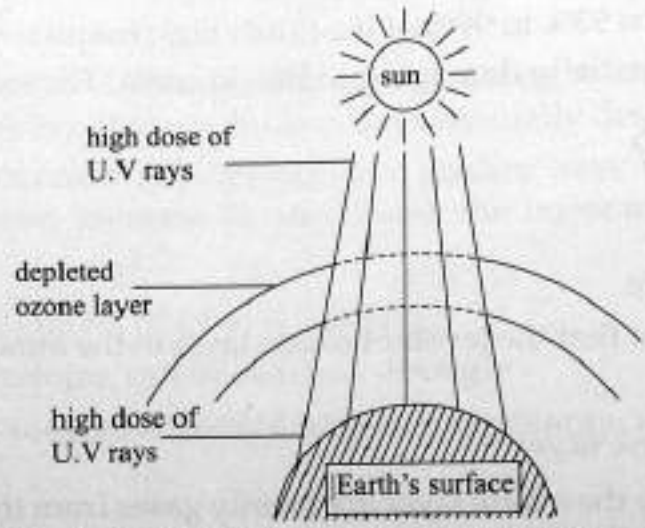


Fig. 8.6(b): Depleted ozone layer allowing uv rays to penetrate to Earth's surface

Effects of the ozone layer depletion

Thinner sections of the ozone layer allow more harmful ultraviolet radiation to penetrate to the earth. This may increase skin cancer cases, damage to crops, cases of cataracts i.e. an eye defect in which the lens is opaque among others.

Note: A high rate of thinning of the ozone layer has been noted over Antarctica and also over the Northern Hemisphere.

Ways of protecting the ozone layer

The ozone layer can be protected by adopting the following practices:

1. Adopt the use of non-fluorocarbon refrigerants such as ammonia, propane and butane.
2. Use of CFC-free aerosol sprays.
3. Replacing stock refrigerants with less harmful hydrofluorocarbons (HCFCs).
4. Recycling empty household aerosol containers and use up of entire product purchased.

Self assessment exercise 8.2

1. What is ozone layer?
2. What is the importance of ozone layer?
3. What do you understand by the term ozone layer depletion?
4. Describe the effects of ozone layer depletion.
5. State ways through which ozone layer depletion can be controlled.

Revision exercise 8

1. What are greenhouse gases?
2. Describe various effects of global warming.
3. Explain how CFCs contribute to the depletion of the ozone layer.
4. Explain the effects of burning fossil fuels and deforestation to the atmosphere.

THE MOUNTAINEERS
CHIPASULA SEC. SCHOOL
PRIVATE BAG 71
LILONGWE

Topic 9 Waste management

Success criteria

After studying this topic, you should be able to:

(a) Classify different types of wastes.
(b) State major sources of wastes.
(c) Describe environmentally friendly ways of treating and disposing wastes.
(d) State ways of minimising wastes.
(e) Describe the social and economic importance of recycling metals and plastics.

Introduction

Waste management is the collection, transport, processing and disposal of wastes. It also involves controlling and monitoring of waste materials. The term **waste** usually relates to unwanted or undesired materials produced by human activities. The process of managing wastes is generally undertaken to reduce its effect on health and the environment. Waste management practices differ from one place to another and for different wastes.

9.1 Classification of wastes

Research Activity

Find out from the library or Internet the various categories of wastes and how they should be managed. Write short notes then share with your friends.

Wastes can be classified based on three factors:

- Physical state
- Level of danger they exhibit
- Biodegradability

(a) Based on physical state, wastes can be categorised as:

- Solid wastes
- Liquid wastes
- Gas wastes
- Sludge wastes

(b) Based on the level of danger they exhibit, wastes can be grouped as:

- Hazardous wastes
- Non- hazardous wastes

Hazardous wastes are either toxic, flammable or explosive. They require special handling and proper procedures during their disposal to avoid health risks and pollution.

(c) Based on biodegradability, types of wastes include:

- **Biodegradable wastes** – These are wastes that can be broken down into harmless substances by micro-organisms. They include animal and plant remains.
- **Non- biodegradable wastes** – These kinds of wastes cannot be broken down into useful components by micro-organisms. Examples of such wastes are plastics polythene and glass.

9.2 Sources of wastes

The major sources of wastes include:

- Domestic sources
- Industrial wastes
- Wastes from farms
- Laboratories wastes

(a) **Domestic sources** – These are wastes from sources such as food leftovers, plastics, metals, untreated sewage, expired or unused medicines and their containers etc.

(b) **Industrial sources** – These include plastics, metals, used oil, toxic gases and other effluents.

(c) **Agricultural sources** – these include plant remains and excess agricultural chemicals such as fertilisers and herbicides.

(d) **Medical sources** – These are wastes from medical centres and hospitals. They include plastics, glass ware, expired drugs, linen, used hypodermic syringes and needles, cotton wool and detergents among others.

(e) **Laboratory sources** – These are waste from laboratories. They include plastics, broken glass wares, expired chemicals, contaminated chemicals, detergents and leaked gases.

Self assessment exercise 9.1

1. Name the different types of wastes based on:
   (a) Physical state
   (b) Degradability
2. List two sources of wastes giving examples in each case.

Effects of wastes on the environment

Research Activity

Research about the effects of wastes in the environment. Write a report and share with your class members. Common effects of wastes include:

1. Some wastes may consist of heavy metals such as mercury which are toxic. Mercury has a cumulative effect i.e. it is transferred along the food chain and can cause death.
2. Lead compounds such as tetraethyl lead ($Pb\ (C_2H_5)_4$) cause serious atmospheric pollution. For a long time, it was used in preventing premature explosion of petrol and air mixture in the vehicle engines. It acted as **antiknocking** agents. This caused alot of air pollution. Lead causes brain damage in young children as well as adults. Most countries have banned the use leaded petrol as a result.
3. Industrial effluents and plastic wastes cause a lot of soil and water pollution. These wastes **degrade** water making it unfit for human consumption.
4. Uncontrolled dumping of toxic wastes for example expired medicines, laboratory chemicals among other substances is common in developing countries. These are sometimes picked by children and people scavanging land fill for valuables and can occasionally be re-sold in the market. These wastes often carry infectious disease-causing micro-organisms which can easily be transmitted by humans and animals scavenging on the dump sites.
5. Burning solid waste from medical sources and decomposing litter often produce toxic gases. These gases are usually accompanied by irritating odours. Such activities also increase carbon dioxide in the atmosphere. Toxic gases increase respiratory infections.
6. In many rural areas, collection and disposal of solid wastes is not taken seriously. Littering by solid wastes is the order of the day. Such wastes are usually human refuse and animal wastes. They end up in streams, rivers and lakes thus causing contamination. Under such circumstances, water borne diseases such as cholera and typhoid are easily transmitted.

9.3 Treatment and disposal of solid wastes

Solid waste is any worthless, unwanted and discarded material that is either liquid nor gas. Solid wastes include food leftovers, plastics, metals, expired drugs and other chemicals. It also includes raked leaves, old newspapers, non-returnable cans, broken bottles and glass, crop residues, animal manure, sewage sludge, worn out furniture, abandoned automobiles and their parts, food processing wastes, mining wastes and construction wastes.

Waste disposal means getting rid of substances that we no longer need in our environment. Activities related to waste disposal are usually done in accordance to local environmental guidelines.

Methods of solid waste disposal:

Some methods that can be used to dispose of waste include: burning, burying, recycling or re-using, use as feeds and as compost manure.

(a) Burning

Wastes such as papers, plastics, carton boxes, empty milk packets, bread wrappers polythene bags among others are usually burnt to ashes.

> *Remember!*
>
> Burning these wastes causes air pollution. Therefore, this method of waste disposal is discouraged.

(b) Burying

This is used to dispose inorganic wastes such as empty containers of insecticides, used torch batteries, broken bottles, pots, glass, plates, cups, and spoons. Organic wastes such as animal carcasses are also buried.

> *Remember!*
>
> We should desist from burying non-bio-degradable wastes as this causes soil pollution.

(c)Animal feeds

Organic wastes such as vegetables, fruits, potato and banana peelings, cabbage and spinach leaves are usually fed to goats, sheep, cattle, poultry and pigs. Bones are dried and ground into powder which is then used as dog food. Food leftovers which are not fit for human consumption can also be fed to animals.

(d) Compost manure

After separating inorganic wastes from organic wastes, a pit that is one metre deep is dug where the organic wastes is dumped to decompose. To keep away flies, reduce smell and other health hazards, a thin layer of wood ashe and soil is used to cover the decomposing matter. After the waste has decomposed, the farmer uses it as compost manure to add fertility to soil.

(e) Recycling or re-use

Recycling is the reprocessing of materials that would otherwise be considered waste. It aims at reducing the cost of producing a new product and reducing pollution by wastes. For example old newspapers, magazine, cartons and packaging materials are sold to vendors who then sell them to recycling industries. These industries convert the waste papers into tissue. The same applies to plastics which can be used to make pipes and other household utilities.

Re-using on the other hand refers to using something for what it was not initially intended for. For example, we can re-use plastic bottles and cans as containers for storing water in our homes.

(f) Incineration

This is the process of destroying medical wastes by the use of fire. The refuse is burnt to ashes in controlled chambers known as **incinerators.** Examples of wastes that are incinerated are medical wastes.

Fig. 9.1: Incineration process

Solid waste processing

The process of proper solid wastes disposal involves:

- Solid waste generation
- Solid waste collection
- Solid waste sorting
- Solid waste transportation.

(a) Solid waste generation

Every day, homes, schools, offices and industries generate wastes. In a well organized system, the garbage bins are provided by the local authorities for storage of wastes. These bins are emptied frequently depending on the garbage generated.

(b) Solid waste collection

Garbage collection trucks go round the residential areas as well as urban centres to collect wastes. Where garbage collection trucks cannot penetrate to collect wastes, push carts are used. After collection, the garbage is taken to a dump site which has been set aside by the local authorities.

(c) Solid waste sorting

Garbage sorting is far much better done at the source. People in residential areas who are the highest generators of solid wastes should separate wastes such as newspapers and packaging materials such as cartons, unbroken bottles, plastic containers and polythene. Some of these wastes such as bottles and plastic containers can be sold to dealers in recycling industries. If garbage sorting is done at the dump site, pieces of

timber and wood recovered are used again as fuel. Also, biodegradable wastes can be taken to compost pits to make manure.

(d) Solid waste transportation

The garbage which remains after sorting and can neither be used again nor recycled is loaded into the garbage trucks and taken to disposal sites. At the disposal site, solid wastes can be treated by the use of the following methods: open dumps, landfills, sanitary landfills and secured landfills.

(i) Open dump

This is a land disposal site where solid wastes and liquid wastes are deposited and left uncovered often with little or no regard to control of scavengers, air pollution and water pollution. This method is strongly discouraged as it is not good for the environment.

It is considered as one of the most harmful methods of disposing solid wastes. This is due to its environmental and health hazards.

(ii) Landfill

This is also land disposal site located with little regard for possible environmental pollution of ground or surface water due to runoff and leaching. It is used to bury solid wastes and sludges. It is better than open dump method though. This is because the layer of soil used to cover it reduces scavenging-spread of disease-causing micro-organisms and air pollution by irritating odour.

(iii) Sanitary landfill

It is a waste disposal site that takes care of water pollution from runoff and leaching. In this method, solid waste is spread in thin layers compacted and covered with a fresh layer of soil each day to minimise air pollution, water pollution and scavenging.

Note: Solid waste disposal in Malawian urban centres since 1996 has been mainly through sanitary landfill and open dump method.

(iv) Secure landfill

It is a site for disposing hazardous wastes (both solid and liquid). The wastes are normally stored in containers and then buried. The site is restricted and continuously monitored. It is located above an impervious geological strata in order to prevent any liquid waste from seeping into the underground water.

Note: In developing countries, solid waste dump sites are most of the time poorly managed. The septic tanks are rarely emptied. When liquid waste from these sources is discharged onto the waste dumping site, it can find its way to underground water sources.

9.4 Treatment and disposal of liquid wastes

Liquid wastes are mainly from the toilets, bathrooms and kitchen. These wastes are directed to the sewage system through the underground pipes. The pipes then drains the liquid wastes either to the ocean or to sewage treatment plant. Where there is no well established sewage system, septic tanks are used.

The liquid wastes are channeled to the septic tank where solid materials undergo biological degradation. When the septic tank is full, the sewage is sucked and disposed or treated to provide clean water and fertilizers.

Ways of minimising wastes

Generally, minimising wastes involves applying three concepts: **recover, recycle** and **re-use.** This is commonly known as the 3Rs.

One way of minimising wastes is **recycling**. It involves reclamation of the waste material and reprocessing it. Recycled materials are usually obtained from resources which get depleted with time for example oil. To conserve our limited resources recycling becomes an important process when handling wastes. Also, scrap metals from used copper and tin products are usually recycled so as to preserve their deposits for future generations.

Materials to be recycled are either brought to a collecting centre or picked up from one spot, then stored, cleaned and reprocessed into new materials bound for re-use.

Some of the waste materials that can be recycled include: tin cans, cardboard, engine oil, glass, paper, plastic bottles, scrap metals, timber and textiles.

Remember

Recently, the government of Malawi issued a ban on production, importation and use of some plastic carrier bags, flatbags, alcohol satchets and plastic papers used in secondary packaging to minimise plastic paper wastes within the country. Ensure that you adhere to this law to help conserve our environment.

Benefits of recycling

Recycling of wastes offers numerous social and economic benefits. The following are some of the benefits:

1. It changes waste materials into new useful materials – it is a waste disposal mechanism.
2. Through recycling, consumption of fresh raw materials is reduced. This helps in conserving non-renewable resources.
3. Recycling saves energy. Reduced amount of energy is needed in recycling than in the fresh production of materials.
4. Reduces air and water pollution by reducing the need for conventional waste disposal mechanisms such as burning and open dumping.
5. It provides jobs to the unemployed i.e. employment opportunities are provided at the gargabe collection sites, and in the recycling companies.

Revision exercise 9

1. Distinguish between waste and waste disposal.
2. With respect to physical state, name the four categories of wastes.
3. State the methods of solid waste disposal.
4. Explain five effects of poor waste disposal to the environment.
5. Describe the social and economic benefits of recycling wastes.
6. State various examples of agricultural and laboratory wastes.
7. What is the meaning of 3Rs in waste management.

Model Test Paper

1. With reference to an energy level diagram, explain how a catalyst speeds up the rate of a chemical reaction.
2. The table below shows results of three experiments on the reaction of excess sulphuric acid with 1g of marble chips done under different conditions. In each experiment, the volume of gas produced was recorded at different time intervals.

Experiment	Form of calcium carbonate	Volume of gas produced
A	Powdered marble chips	
B	Small marble chips	
C	Large marble chips	

   (a) Sketch and label the three curves that could be obtained from such results.

   (b) Why does the reaction stop after sometime?

3. (a) (i) What is the oxidation state of chlorine in HCl?

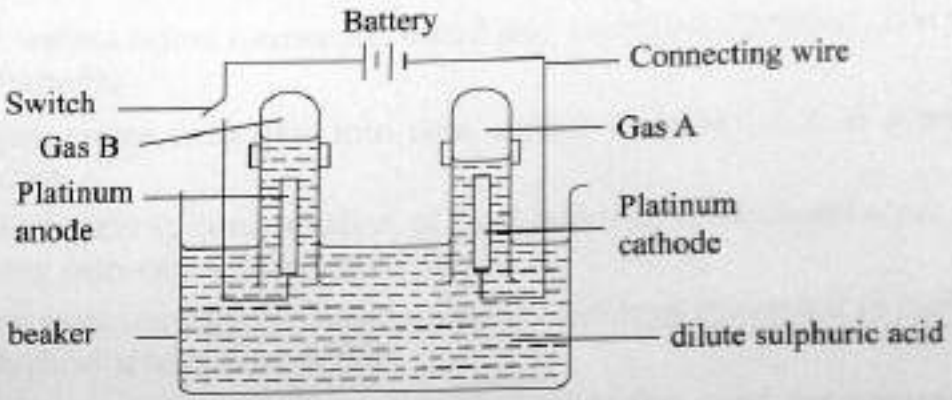
   (ii) What is the oxidation state of chlorine in ClO_3^- ?

   (b) Identify the reduced spicies in the equation below.

   $$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$$

   (c) What is the oxidation number of Mg, H^+, H_2, Mg^{2+} ?

4. Carbon dioxide can be dissolved in water under pressure to make an acidic solution. What is an acidic solution?
5. The diagram below represents a set- up that was used to study the effect of electric current on dilute sulphuric acid.



   (a) State and explain the observation made when each experiment was started.

   (b) State and explain the observation made when the experiment was done with copper sulphate solution as the electrolyte instead of dilute sulphuric acid.

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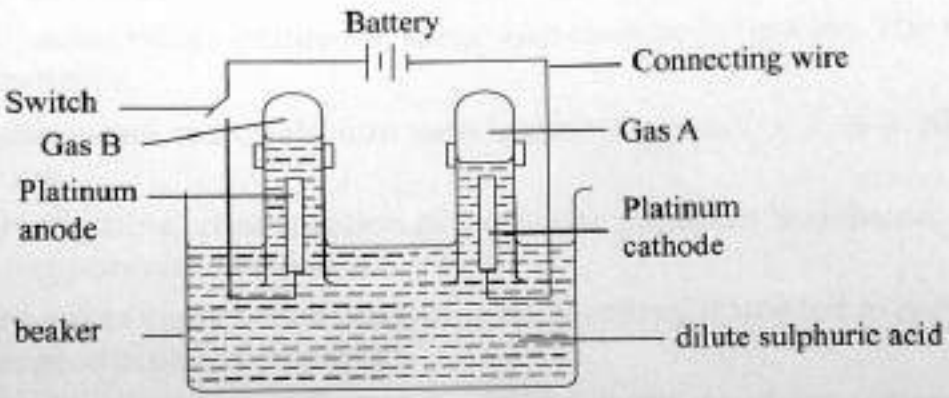


   (a) State and explain the observation made when each experiment was started.

   (b) State and explain the observation made when the experiment was done with copper sulphate solution as the electrolyte instead of dilute sulphuric acid.

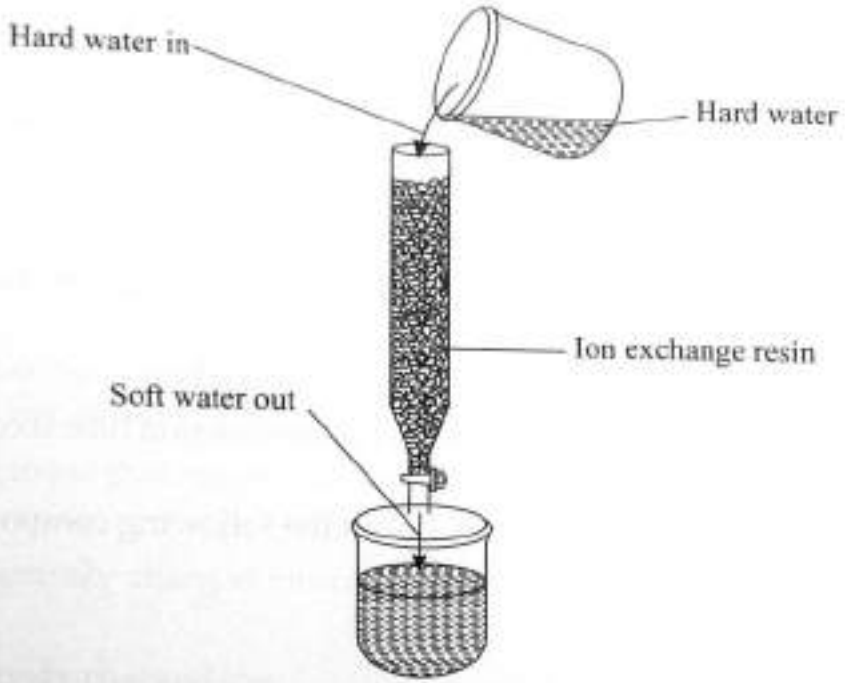
6. A student added very dilute sulphuric acid to four metals and recorded the observations shown in the table below.

Experiment	Metals	Gas given off
1	Sodium	Yes
2	Iron	No
3	Carbon	Yes
4	Copper	No

Which experiment does not match with the correct observation? Explain.

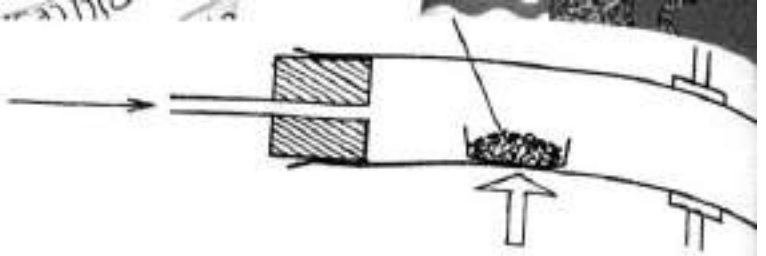
7. (a) What is the meaning of water hardness?
   (b) Describe two methods used to remove water hardness.

8. The column below was used to soften hard water.



   (a) Explain how hard water is softened as it passes through the column.
   (b) After sometime, the material in the column is not able to soften hard water. How can the material be re-activated?
   (c) Give one advantage and one disadvantage of hard water.

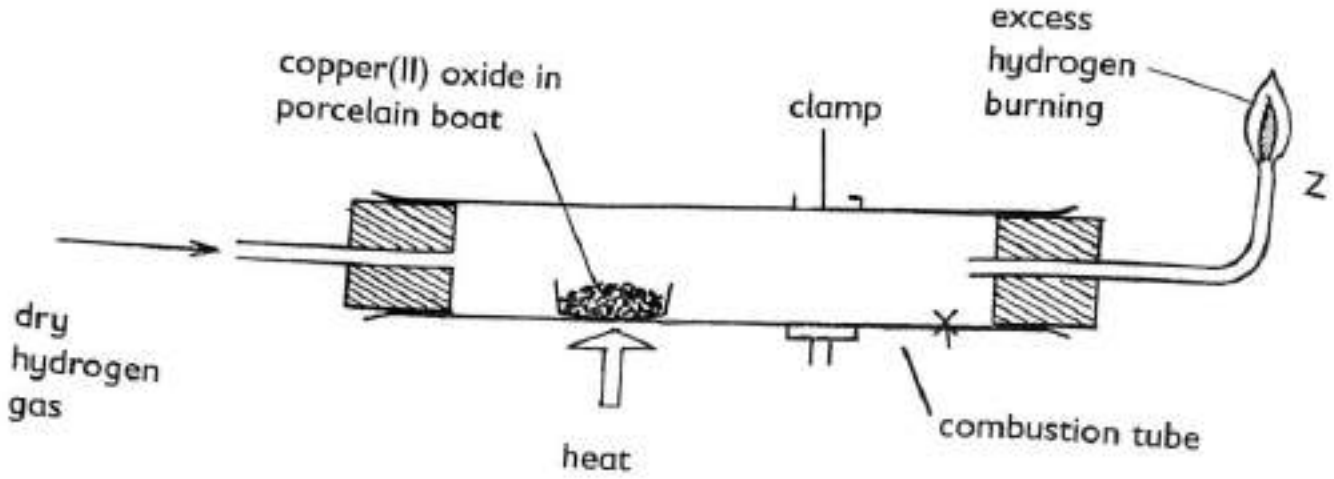
9. Carbon dioxide, methane, nitrogen monoxide and trichloroflouromethane are greenhouse gases.
   (a) State the effect of an increased level of these gases to the environment.
   (b) Give one source of each of the following gases in the environment.



(i) Nitrogen monoxide

(ii) Trichloroflouromethane

10. Draw the structural formula of:
    (a) Ethanol
    (b) Propanoic acid
11. Describe the effect of branching on melting and boiling points of alkanes.
12. In an experiment, dry hydrogen gas was passed over heated lead oxide as shown in the diagram below.



    (a) State and explain the observations that were made in the combustion tube. Use chemical equations.
    (b) What is the name of this type of reaction? Explain.
    (c) What observation would be made in the combustion tube if copper oxide was used instead of lead oxide?
13. Determine the oxidation state of sulphur in the following compounds:
    (a) H_2S
    (b) Na_2SO_4
    (c) H_2SO_4
    (d) H_2SO_3
14. (a) What is meant by isomerism?
    (b) Draw and name the isomers of pentane and butane.
15. Describe three ways of reducing greenhouse effect.
16. How does the following natural disasters affect the atmosphere?
    (a) Volcanic eruptions
    (b) Forest fires

Glossary

Acid rain: Acid rain is as a result of dissolved gases such as carbon dioxide, sulphur dioxide, oxides of nitrogen and hydrogen chloride produced from industries and other human activities.

Acid: A substance which contains hydrogen ions which can be displaced by a metal or an ammonium radical.

Acidic oxides: Oxides of non- metals that react with water to form acidic solutions.

Addition polymerisation: The process that occurs when unsaturated monomers combine in a chain forming a polymer.

Alkali: Any base that is soluble in water.

Alkanols: Organic compounds made of a hydrocarbon group and one or more hydroxyl (OH) groups e.g. ethanol (C_2H_5OH). They are also known as alcohols.

Alloy: A mixture which is made up of two or more metals. It may also contain non-metals. Examples are brass, bronze and steel.

Amphoteric oxide: A metal oxide which reacts both as an acid and a base producing salts in both cases. Examples are aluminium oxide, lead (II) oxide and zinc oxide.

Anion: A negatively charged ion.

Anode: The positively charged electrode. It is connected to the positive terminal of the battery.

Base: A substance that reacts with an acid to form salt and water only. For example, metal oxides or hydroxides.

Basic oxide: An oxide that reacts with an acid to form salt and water only, for example metal oxides.

Cathode: A negatively charged electrode. It is connected to the negative terminal of the battery.

Cation: A positively charged ion.

Condensation polymerisation: A reaction in which a simple molecule (for example water or ammonia) is produced when reactant molecules join together to form a polymer.

Conductivity: A measure of the ability of a material to allow electricity to pass through it. It also includes ability to allow heat energy to pass through it.

Electric current: A flow of electrons from one point to another.

Electrode: A metal or carbon rod (graphite) used to conduct electricity into or out of an electrolyte.

Electrolysis: Decomposition of a compound in solution or molten form when an electric current is passed through it.

Electrolyte: A substance which conducts electric current in solution or in molten form and in the process it gets decomposed.

Electroplating: A process in which a thin layer of metal is deposited on the surface of another metal by electrolysis.

Functional group: The atom or group of atoms responsible for the characteristic chemical reactions of an organic compound.

Hydrated: Having water of crystallisation e.g. hydrated copper sulphate

Ion – a charged atom or group of atoms that are charged.

Neutral oxide – An oxide of a non-metal which is neither acidic nor basic e.g. carbon monoxide.

Neutralisation: A chemical reaction that occurs between a base and an acid to produce salt and water only.

Non-electrolyte: A substance which does not conduct electricity in solution or molten form, for example, benzene.

Oxidation number or state: It is the charge that an atom of an element has when in a molecule or when existing separately as an ion.

Oxidation: A reaction involving loss of electrons from an atom, molecule or ion or addition of oxygen or loss of hydrogen atom.

Oxidising agent: The atom, radical or ion which takes up electrons during a chemical reaction.

Permanent hardness: Hardness caused by a sulphates and chlorides of calcium and magnesium and can only be removed by addition of a chemical softener like sodium carbonate and not boiling.

pH: The measure of degree of acidity or alkalinity of a solution.

Polymer: A long chain molecule built up of a number of repeating units called monomers.

Polymerisation: The chemical reaction in which molecules (monomers) join together to form a long – chain (polymer).

Redox reaction: Short form for reduction-oxidation reaction. It is a reaction whereby both oxidation and reduction processes occur.

Reducing agent: The atom, radical or ion which gives out electrons during a chemical reaction.

Reduction: A reaction involving gain of electrons by an atom, molecule or ion or loss of oxygen atom or gain of hydrogen atom.

Temporary hardness: Hardness of water caused by the hydrogen carbonate or calcium and magnesium ions which can be removed by boiling the water.

Thermoplastics: Plastics which soften when heated. They can be remolded as well.

Thermosetting plastics: Plastics which do not soften when heated. They cannot be remolded.

Universal indicator: A mixture of indicators with various colours according to the level of acidity or alkalinity of the solution in which it is placed.

Valence electron: An electron that can participate in bond formation with other atoms.

Valency: A term that refers to the combining power of an element.

Waste management: Collection, transport, processing, disposal, monitoring and control of waste materials.

Water of crystallisation – Fixed amount of water in a crystal, for example, ($CuSO_4.5H_2O$) has five molecules of water of crystallisation.

Water softener: A chemical substance which is added to hard water to remove hardness.

Appendices

Appendix I

Atomic numbers and relative atomic masses of some common elements

Element	Symbol	Atomic number	Relative atomic mass
Aluminium	Al	13	27
Argon	Ar	18	40
Barium	Ba	56	127
Beryllium	Be	4	9
Boron	B	5	11
Bromine	Br	35	80
Calcium	Ca	20	40
Carbon	C	6	12
Chlorine	Cl	17	35.5
Chromium	Cr	24	52
Copper	Cu	24	52
Flourine	F	9	19
Gold	Au	79	197
Helium	He	2	4
Hydrogen	H	1	1
Iodine	I	53	127
Iron	Fe	26	56
Krypton	Kr	36	84
Lead	Pb	82	207
Lithium	Li	3	7
Magnesium	Mg	12	24
Manganese	Mn	25	55
Mercury	Hg	80	200
Neon	Ne	10	20
Nickel	Ni	28	59
Nitrogen	N	7	14
Oxygen	O	8	16

Phosphorous	P	15	31
Potassium	K	19	39
Silicon	Si	14	28
Silver	Ag	47	108
Sodium	Na	11	23
Sulphur	S	16	32
Tin	Sn	50	119
Vanadium	V	23	51
Zinc	Zn	30	65

Appendix II

Valencies of some of the elements in the Periodic Table

Name of metal	**Symbol**	**Valency**
Zinc	Zn	2
Iron	Fe	2 or 3
Tin	Sn	4
Lead	Pb	2
Copper	Cu	1 or 2
Silver	Ag	1
Barium	Ba	2
Phosphorus	P	3 or 5

Appendix III

Valencies of some common radicals

Valency 1		**Valency 2**		**Valency 3**	
Radical	*Formula*	*Radical*	*Formula*	*Radical*	*Formula*
Ammonium	NH_4^+	Carbinate	CO_3^{2-}	Phosphate	PO_4^{3-}
Hydroxide	OH^-	Sulphate	SO_4^{2-}		
Nitrate	NO_3^-	Sulphite	SO_3^{2-}		
Chloride	Cl^-				

Hydrogen carbonate	HCO_3^-				
Hydrogen sulphide	HSO_4^-				

Appendix IV

Valencies of some elements in some compounds

Compound	Element	Valency
Copper(I) oxide	Copper	1
Copper(II) oxide	Copper	2
Iron(II) sulphate	Iron	2
Iron(III) chloride	Iron	3
Sulphur(IV) oxide	Sulphur	4
Sulphur(VI) oxide	Sulphur	6
Carbon(IV) oxide	Carbon	4
Carbon(II) oxide	Carbon	2

Appendix IV: The Periodic Table of Elements

x
Z
y
Name

x is the Relative Atomic Mass (RAM)
Z is the symbol of the element
y is the atomic number of the element

Periods	Groups I	II											III	IV	V	VI	VII	VIII
1	1 H 1 Hydrogen																	4 He 2 Helium
2	7 Li 3 Lithium	9 Be 4 Beryllium											11 B 5 Boron	12 C 6 Carbon	14 N 7 Nitrogen	16 O 8 Oxygen	19 F 9 Fluorine	20 Ne 10 Neon
3	23 Na 11 Sodium	24 Mg 12 Magnesium	Transition metals										27 Al 13 Aluminium	28 Si 14 Silicon	31 P 15 Phosphorus	32 S 16 Sulphur	35.5 Cl 17 Chlorine	40 Ar 18 Argon
4	39 K 19 Potassium	40 Ca 20 Calcium	45 Sc 21 Scandium	48 Ti 22 Titanium	51 V 23 Vanadium	52 Cr 24 Chromium	55 Mn 25 Manganese	56 Fe 26 Iron	58.9 Co 27 Cobalt	59 Ni 28 Nickel	64 Cu 29 Copper	65 Zn 30 Zinc	70 Ga 31 Gallium	73 Ge 32 Germanium	75 As 33 Arsenic	79 Se 34 Selenium	80 Br 35 Bromine	84 Kr 36 Krypton
5	86 Rb 37 Rubidium	88 Sr 38 Strontium	89 Y 39 Yttrium	91 Zr 40 Zirconium	93 Nb 41 Niobium	96 Mo 42 Molybdenum	99 Tc 43 Technetium	101 Ru 44 Ruthenium	103 Rh 45 Rhodium	106 Pd 46 palladium	108 Ag 47 Silver	112 Cd 48 Cadmium	115 In 49 Indium	119 Sn 50 Tin	122 Sb 51 Antimony	128 Te 52 Tellurium	127 I 53 Iodine	131 Xe 54 Xenon
6	133 Cs 55 Caesium	137 Ba 56 Barium	175 Lu 71 Lutetium	178 Hf 72 Hafnium	181 Ta 73 Tantalum	184 W 74 Tungsten	186 Re 75 Rhenium	190 Os 76 Osmium	192 Ir 77 Iridium	195 Pt 78 Platinum	197 Au 79 Gold	201 Hg 80 Mercury	204 Ti 81 Thallium	207 Pb 82 Lead	209 Bi 83 Bismuth	210 Po 84 Polonium	210 At 85 Astatine	222 Rn 86 Radon
7	223 Fr 87 Francium	226 Ra 88 Radium	257 Lr 103 Lawrencium	261 Rf 104 Rutherfordium	262 Db 105 Dubnium	266 Sg 106 Seaborgium	264 Bh 107 Bohrium	269 Hs 108 Hassium	268 Mt 109 Meitnerium	281 Ds 110 Darmstadtium	280 Rg 111 Roentgenium	285 Cn 112 Copernicium	284 Uut 113 ununtrium	289 Fl 114 Flerovium	288 Uup 115 Ununpentium	292 Lv 116 Livermorium	294 Uus 117 Ununseptium	294 Uuo 118 ununoctium

Lanthanide elements	257 La 57 Lanthanum	140 Ce 58 Cerium	141 Pr 59 Praseodymium	144 Nd 60 Neodymium	147 Pm 61 Promethium	150 Sm 62 Samarium	152 Eu 63 Europium	157 Gd 64 Gadolinium	159 Tb 65 Terbium	163 Dy 66 Dysprosium	165 Ho 67 Holmium	167 Er 68 Erbium	169 Tm 69 Thulium	173 Yb 70 Ytterbium	
Actinide elements	227 Ac 89 Actinium	232 Th 91 Thorium	231 Pa 91 Protactinium	238 U 92 Uranium	237 Np 93 Neptunium	242 Pu 94 Plutonium	243 Am 95 Americium	247 Cm 96 Curium	249 Bk 97 Berkelium	251 Cf 98 Californium	254 Es 99 Einsteinium	253 Fm 100 Fermium	256 Md 101 Mendelevium	254 No 102 Nobelium	

KEY

Metals

Non-metals

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