Chemistry

For Rwanda Schools

Student's Book

Senior 3

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UNIT 1

Carbon and its inorganic compounds

Key unit competency

To be able to relate the properties of carbon and its compounds to its uses and describe how the compounds of carbon are prepared.

Learning objectives

By the end of this unit, I should be able to:

- Name the allotropic forms of carbon and relate its properties to its uses.
- Explain the properties of carbon and its compounds.
- Prepare, collect and test for carbon dioxide gas.
- Prepare, collect and test carbonates of different metals.
- Explain the impact of carbon compounds on the environment.
- Explain the carbon cycle.
- Appreciate the importance of natural resources.
- Develop self – confidence during presentations and a culture of working cooperation together in groups.

Mind teaser

Study the pictures below. Identify the things in the pictures.

Fig 1.1

Can you tell from what element the items or some of their components in the pictures are made from?

Introduction

The items shown in the pictures, fig 1.1 are all made from or contain carbon element. Carbon is one of the most important elements. Carbon exists both naturally as **diamond** and **graphite** and in combined forms in many compounds. Coal, oil, natural gas, limestone and other metal carbonates are all compounds of carbon. Charcoal and the positive electrode in dry cells are all composed of carbon.

1.1 Definition of allotropy

Activity 1.1

Search about meaning of allotropy and present your findings to the class. Use the following questions as a hint.

1. What is allotropy? Get the meaning from the dictionary, textbooks or the internet.
2. Which elements exhibit allotropy?
3. Name two allotropes of carbon that you have identified from this activity.

The facts

Atoms in certain solids often may arrange themselves in different patterns. When this occurs it results in different forms of that element called **allotropes**. Allotropy is therefore the existence of an element in two or more forms. Elements that exhibit allotropy include carbon, sulphur and phosphorus. Allotropes have same chemical properties but different physical properties.

1.2 Allotropes of carbon and their physical properties

In Senior 2 under covalent bonding, you learnt about giant covalent structures, their properties and uses. In this subtopic, you are going to learn how covalent bonding brings about the allotropic forms of carbon.

Physical properties of carbon allotropes and their uses

Activity 1.2

Materials

Gloves, dry cells, ball-pein hammer and a protective sheet of paper.

Procedure

1. Spread the protective sheet of paper on the floor.
2. Wear gloves and use the ball-pein hammer to crush the dry cell.

3. Remove the black rod at the centre.
4. Try breaking it into smaller pieces and feel it between your fingers.

Study questions

1. What is the colour of charcoal, soot and the rod in the dry cell?
2. What is the texture of the charcoal and soot?
3. How are soot and charcoal formed?
4. Which allotrope of carbon is used to make the rod in the dry cell?

Discussion corner!

1. Answer the study questions above.
2. In groups, discuss the concept of allotropy; consider graphite in dry cells.
3. Prepare a report and present it to the rest of the class.

Self-evaluation Test 1.1

1. What is allotropy?
2. Name two elements that exhibit allotropy.
3. Name the allotropes of elements named in question 2 above.
4. Complete the table below in relation to charcoal.

Colour	**Texture**

5. Fill in the missing words

   a. ____________ is the allotrope of carbon used in dry cells.

   b. Soot and charcoal are formed when __________________.

6. What is the role of the rod in a dry cell?

Crystalline forms of carbon

Research activity

1. Using textbooks and from the internet, research on the structure, the physical properties and uses of diamond and graphite.
2. Write a report of your findings and do a class presentation.
3. Compare your findings with the ones given below.

The facts

There are two main crystalline forms of carbon: **diamond** and **graphite.** Non-crystalline forms of carbon also exist. They are called **amorphous carbon.**

(a) Diamond

Diamond occurs naturally in many countries. Thus diamond is a natural resource.

Structure of diamond

In diamond, each carbon atom is covalently bonded to four other carbon atoms giving a regular tetrahedron shape as shown in fig 1.2. This results in a rigid closely interlocked three dimensional structure. The atoms form a giant atomic structure. All the four valence electrons in diamond are used in the bonding, which greatly contributes to the physical properties of diamond.

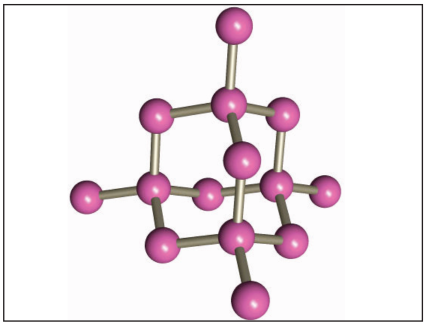


Fig 1.2: Structure of diamond

Physical properties of diamond

- It is the hardest known naturally occurring substance. This is because of the strong covalent bonds that hold the atoms of carbon together.
- When carefully and specially cut, it is colourless crystalline and transparent with a dazzling brilliant lustre. The lustre is caused by its high refractive index.
- Diamond has a higher density of 3.5g/cm^3 compared to that of graphite (2.25g/cm^3). This is because the arrangement of atoms in a diamond crystal allows for more atoms to be packed per given space compared to graphite.
- It has a high melting and boiling points. It melts at 4200°C. This is because all the carbon atoms are bonded by very strong covalent bonds that require a lot of heat energy to break.
- Diamond does not conduct electricity. This is because it has no delocalized electrons.
- Diamond is a good conductor of heat because of its strong covalent bonding.

Uses of diamond

Can you remember the uses of diamond you learnt in Senior 2?

The physical properties of diamond makes it find important applications in the following areas.

1. Because of its hardness, it is used for making drill tips used when drilling or cutting metals. It is also used to make glass cutters.
2. Used in making jewellery due to its beautiful sparkling radiance and lustre.

Fig 1.3: Diamond chain

(b) Graphite

Graphite is a dark grey shiny crystalline solid. It is the most abundant allotrope of carbon. Graphite occurs naturally but a superior form of it is made by heating anthracite (a variety of coal) in an electric furnace.

Structure of graphite

In graphite, each carbon atom is covalently bonded to three other atoms while the fourth electron is delocalised; i.e it is not attached to any particular atom but belongs to the entire structure. These delocalised electrons are free to move through the graphite structure making graphite a good conductor of electricity. Graphite exists in hexagonal layers of carbon atoms. The layers are held together by weak Van der Waal's forces of attraction as shown in fig 1.4. This makes them slide easily over each other.

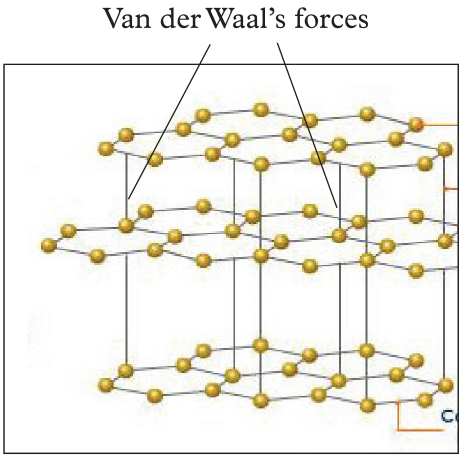


Fig 1.4 :Structure of graphite

Physical properties of graphite

- Graphite is a soft, black and shiny material with a greasy feel.
- Graphite easily flakes off.
- Has relatively high melting and boiling points due to the strong covalent bonds joining its atoms together.
- It has a density of 2.25g/cm^3.
- Graphite is a good conductor of electricity due to the presence of delocalised electrons in its structure.
- It is opaque. Numerous parallel layers arranged on top of one another blocks light from penetrating through.

Uses of graphite

The properties of graphite determine its uses. Some uses of graphite include:

1. It is used as electrodes in dry cells and fuel cells. This is because it is a good conductor of electricity.
2. Graphite is used as a carbon raiser in the production of steel. It gives steel its strengthening characteristics.
3. It is used in advanced high-friction applications such as car brakes and clutches because of its high thermal and electrical conductivity.
4. It is used to make pencil 'leads' when mixed with clay. This is due to the sliding of its layers which enable it to slide on paper when writing.

5. At high temperatures, graphite can be used as a lubricant in place of grease or oil. This is because graphite has a high boiling point.
6. In production of paints and shoe polish, the powdered form of lump graphite is used due to its natural water-repellent property. It is the best choice for giving a protective coating on wood or shoe leather.
7. It is used to make furnace linings, brushes for electric motors and generators.
8. It is used as a moderator in nuclear reactors.

Fig 1.5: Graphite lubricant

Non-crystalline forms of carbon

Activity 1.3

Materials

Pieces of charcoal, soot, gloves, ball-pein hammer and a sheet of paper.

Procedure

1. Spread the protective sheet of paper on a work bench.
2. Place the pieces of charcoal and soot on the bench.
3. In pairs, feel the charcoal and soot between your fingers.
   - Observe what happens to your fingers.
4. Crush the pieces of charcoal into smaller pieces and see what happens.

Non-crystalline forms of carbon are formed by decomposing other substances like wood using heat. Charcoal and soot are examples of amorphous carbon.

Charcoal

There are two types of charcoal; wood and animal charcoal. Wood charcoal is obtained when wood is strongly heated in absence of air. This process is called **destructive distillation** of wood. Remember that large-scale and uncontrolled charcoal-making leads to destruction of forests.

My environment, my life!

Avoid cutting down trees to make charcoal. Use alternative sources of fuel like biogas that is environmentally friendly.

Animal charcoal is formed when bones are subjected to destructive distillation. Animal charcoal is only about ten percent carbon with the rest being mainly calcium phosphate.

Fig 1.6: Animal bones

Other sources of amorphous carbon include:

- Burning petroleum products in a limited supply of air to form **lamp black**.
- Heated sugar in limited supply of air to form sugar charcoal. This can also be formed by dehydrating sugar using concentrated sulphuric acid.
- Heating coal in the absence of air to obtain coke.
- Soot is formed when there is incomplete combustion of fuels. It is found e.g in chimneys of houses and lantern lamp.

Physical properties of charcoal

Activity 1.4

Materials

Pieces of charcoal, two 100 cm^3 beakers, solution of a dye, filter paper, filter funnel, Bunsen burner.

Procedure

1. Put about 50 cm^3 of the dye solution into a beaker.
2. Place some pieces of charcoal into the beaker and heat for about ten minutes.
3. Filter the contents of the beaker and observe any changes in the filtrate.

Study questions

1. What changes occurred in the solution after it was heated with charcoal?
2. Explain the changes that took place.

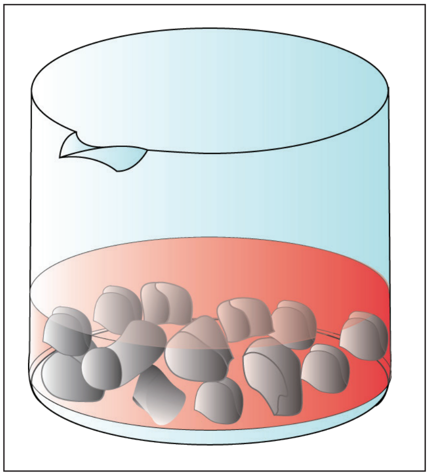


Fig 1.7 : Adsorption by charcoal

I have discovered that…

The intensity of the dye in solution decreased at the end of the experiment. Particles of the dye stick onto the surface of charcoal.

The facts

Charcoal adsorbs particles in a dye reducing the intensity of the colour of the dye. The particles of adsorbed substance stick onto the surface of the charcoal. This property is used in domestic water treatment. Some physical properties of charcoal include:

- Charcoal is a black porous solid.
- It is soft and has a low density.
- It can adsorb large volumes of gases and solids.

Uses of charcoal

1. Charcoal is used as a source of fuel mostly in developing countries.
2. Activated charcoal (charcoal that is finely powdered) is used to remove smelly gases in slaughter houses, gas manufacturing plants, large air conditioning systems and airports. This is because activated charcoal has a large surface area and hence can adsorb large volumes of gases.

Fig 1.8: Pieces of charcoal

3. Lamp black is used in making black ink, paints, carbon paper and as an ingredient in the rubber tyres.

Self-evaluation Test 1.2

1. While cooking, if the bottom of the vessel is getting blackened on the outside, it means_______________.
2. Name one non-crystalline form of carbon that you know.
3. Match the following statements (i.e A and B) using a line.

A	B
Amorphous carbon	Diamond
A good conductor of electricity	Made of hexagonal layers joined by weak Van der Waals forces
Graphite	Graphite
Uses all electrons for bonding during its formation	Non-crystalline form of carbon

4. Explain the following:
   (a) Diamond is used in making tips of drills while graphite is used to make pencil leads.
   (b) Graphite conducts electricity while diamond does not.
5. Charcoal is a cheap but expensive source of fuel. Explain.
6. How does excessive use of charcoal as a source of fuel affect our environment?
7. Explain why graphite is used as a lubricant whereas diamond is not.

1.3 Chemical properties of carbon

a) Reaction with oxygen

Activity 1.5

Apparatus and reagents

Deflagrating spoon, Bunsen burner, gas jar of oxygen, calcium hydroxide solution.

Procedure

1. Place a piece of charcoal in a deflagrating spoon and heat it strongly using a Bunsen burner flame until the charcoal glows red-hot.
2. Lower the red-hot charcoal into a gas jar of oxygen.
   - What do you observe?
3. After sometime, pass the product in the jar through a solution of calcium hydroxide and observe what happens.

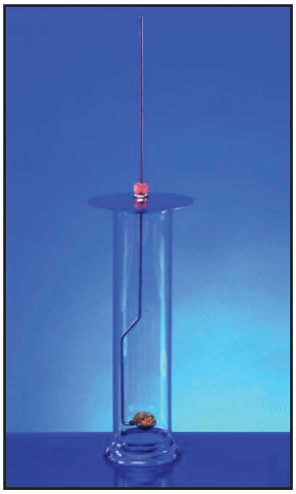


Fig 1.9 Burning charcoal in oxygen

Study questions

1. Why did burning stop after sometime?
2. What caused the white precipitate when the gas formed in the jar was passed through the calcium hydroxide solution

Discussion corner!

Discuss the observations you have made and the study questions in activity 1.5 with a classmate and share your conclusions with the rest of the class.

I have discovered that…

Charcoal continues to burn in the gas jar even after heating is stopped. The calcium hydroxide solution turned milky due to the carbon dioxide gas formed when charcoal burnt in the gas jar of oxygen.

The facts

Carbon burns in enough supply of oxygen to form carbon dioxide gas. The reaction is exothermic and hence burning continues even after heating has been stopped.

Carbon + oxygen → carbon dioxide + heat

$C(s) + O_2(g) \rightarrow CO_2(g) + \text{heat}$

When carbon dioxide was passed through calcium hydroxide, it reacted to form calcium carbonate thus white precipitate is seen. The white precipitate would disappear if carbon dioxide gas is bubbled in excess due to formation of calcium hydrogen carbonate.

Calcium hydroxide + carbon dioxide ⟶ Calcium carbonate + water

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

Calcium carbonate + carbon dioxide + water ⟶ Calcium hydrogen carbonate

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

Carbon reacts with limited oxygen supply to form carbon monoxide.

Carbon + oxygen ⟶ carbon monoxide + heat

$2C(s) + O_2(g) \longrightarrow 2CO(g) + \text{heat}$

Health check

Avoid using charcoal stove in a closed room because the built up of carbon monoxide to toxic levels causes suffocation leading to death!

b) Reaction with carbon dioxide

Activity 1.6

Materials

Charcoal, source of carbon dioxide, source of heat, combustion tube.

Procedure

1. Set up the apparatus as shown below.

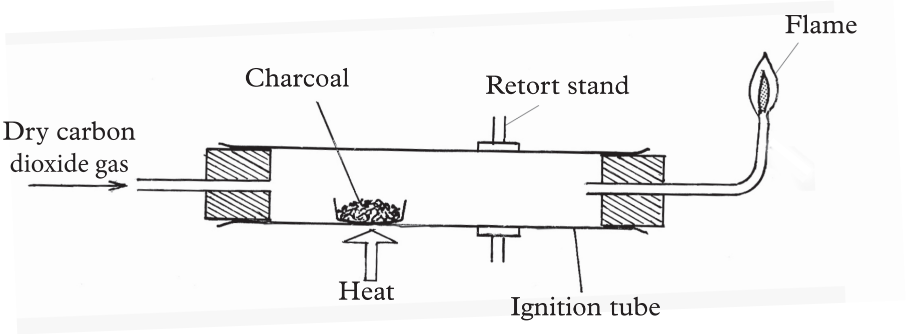


Fig 1.10: Experimental set-up for the reaction between charcoal and carbon dioxide

2. Heat the charcoal until it is red-hot.
3. Pass dry carbon dioxide gas through the heated charcoal.
4. Ignite the gas coming out through the other side of the combustion tube as shown above.

Precaution: Avoid inhaling the gas produced. It is poisonous!

Study questions

1. What observations did you make in the ignition tube?
2. Write a balanced chemical equation for the reaction that took place in the combustion tube.

Discussion corner!

In pairs, discuss the results of the experiment and the study questions above. Write a report and present to the other class members.

I have discovered that…

When carbon dioxide is passed through heated charcoal, the charcoal diminishes. A gas is produced that burns with a blue flame.

The facts

Carbon dioxide reacts with heated charcoal to form carbon monoxide gas. Carbon monoxide in turn burns with blue flame.

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

c) Reaction with iron (III) oxide

Activity 1.7

Materials and reagents

Crucible, iron (III) oxide, source of heat, powdered charcoal, tripod stand, wire gauze.

Procedure

1. Mix charcoal powder with iron (III) oxide in a crucible.
   - Note the initial colour of the mixture.
2. Apply heat on the crucible until it is red-hot.
   - Note the colour of the final product.

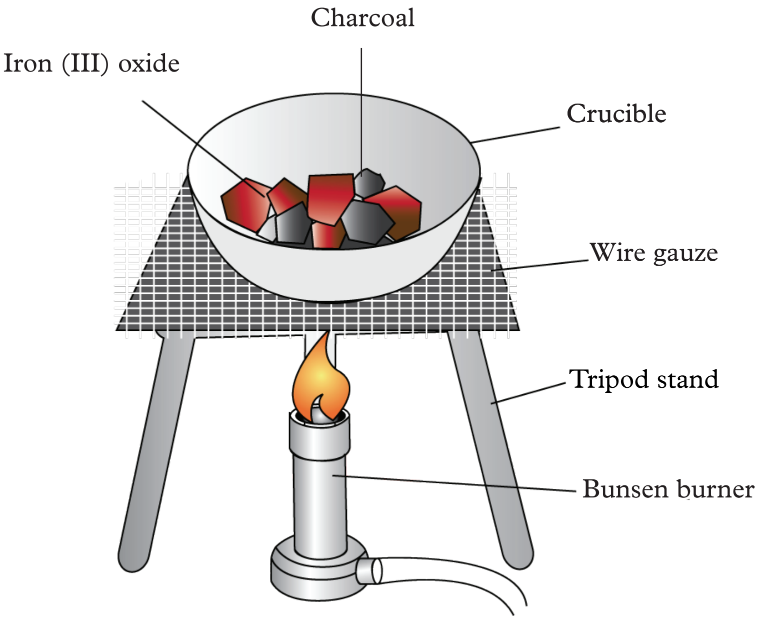


Fig 1.11 Set up for the reaction between carbon and iron (III) oxide

Study questions

1. Describe the colour change from that of the inside mixture and that of the final product. Give reasons for these changes.

Colour of the mixture at the beginning. ____________________	Colour of the mixture after heating. ____________________

2. Write a chemical equation for the reaction that took place.

Discussion corner!

In pairs, discuss the observations made and answers to the study questions above and write a report.

I have discovered that…

When the mixture of iron (III) oxide and powdered charcoal is heated, a red glow is produced. The glow continues spreading throughout the mixture even when heating is stopped. The colour of the mixture turns from red-brown before heating to grey after heating.

The facts

Carbon in the charcoal reacts with iron (III) oxide to form iron metal and carbon dioxide. This is a redox reaction where carbon reduces iron from its oxide and is itself oxidised to carbon dioxide. The reaction is exothermic and hence continues even after heating is stopped.

Iron (III) oxide + carbon ⟶ Iron + carbon dioxide

$$\underset{\text{(Red-brown)}}{2Fe_2O_3(s)} + 3C(s) \longrightarrow \underset{\text{(Grey)}}{4Fe(s)} + 3CO_2\,(g)$$

Work to Do

Name other metal oxides reduced by carbon to their respective metals

d) Reaction of carbon with concentrated sulphuric acid and nitric acid

Activity 1.8

Apparatus and reagents

Two test tubes, concentrated sulphuric acid, cork, delivery tube, retort stand, concentrated nitric acid, powdered charcoal, spatula, bunsen burner, tripod stand, calcium hydroxide solution (lime water).

Procedure

1. Put a spatula full of charcoal powder in a test tube.
2. Add about 10 cm^3 of concentrated sulphuric acid, warm and observe.
3. Repeat the experiment using concentrated nitric acid instead of sulphuric acid.

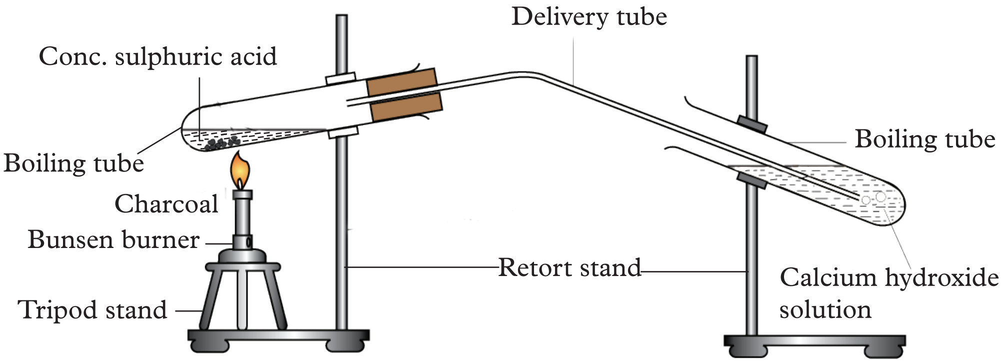


Fig 1.12 Reaction of carbon with concentrated acids

Study questions

1. What observations did you make on reacting each acid with carbon?
2. Write balanced chemical equations for the reactions that took place.

Discussion corner!

1. In pairs, discuss the observations made.
2. Come up with an explanation for your observations. Present to the rest of the class.

I have discovered that…

When carbon reacts with concentrated nitric and sulphuric acids, bubbles of a gas are produced. The amount of charcoal powder in the test tube eventually decreases.

The facts

Carbon reacts with concentrated sulphuric acid to form carbon dioxide, sulphur dioxide and water. With concentrated nitric acid, it forms carbon dioxide, nitrogen dioxide and water as shown below.

Carbon + sulphuric acid ⟶ carbon dioxide + sulphur dioxide + water

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

Carbon + nitric acid ⟶ carbon dioxide + nitrogen dioxide + water

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

Note: These acids oxidise carbon to carbon dioxide. It is the carbon dioxide gas that turns lime water milky.

Self-evaluation Test 1.3

1. Carbon reacts with carbon dioxide according to the equation below.
   $CO_2\,(g) + C(s) \longrightarrow 2CO\,(g)$
   Identify the substance that loses oxygen and the one that gains oxygen from the reaction above.
2. Carbon reacts with oxygen according to the equations below:
   i. $C(s) + O_2\,(g) \longrightarrow CO_2\,(g) + \text{heat}$
   ii. $2C(s) + O_2\,(g) \longrightarrow 2CO(g) + \text{heat}$
   (a) What condition is required in (i) and (ii) above?
   (b) What does heat in the equations indicate about the two reactions?
3. Write balanced chemical equations for the reactions that take place when:
   i. Carbon reacts with concentrated nitric acid.
   ii. Carbon reacts with concentrated sulphuric acid.

1.4 Inorganic compounds of carbon

Carbon can exist as a free element or in combined state. Commonly known inorganic compounds of carbon include: oxides of carbon, carbonates and hydrogen carbonates.

a) Oxides of carbon

Research activity

1. Using textbooks in library or from the internet, research on the physical properties of carbon oxides.
2. Present your findings to the class.
3. Compare your findings with the ones given below.

The two oxides of carbon are carbon dioxide and carbon monoxide.

i. Carbon monoxide

This is a colourless and odourless gas. It is slightly denser than air. Carbon monoxide is produced by incomplete combustion (burning of carbon in limited oxygen supply) of various fuels. These include coal, wood, charcoal, fractions of oil and natural gas. It is a poisonous and very toxic gas. This is because when inhaled, it binds to haemoglobin blocking off oxygen. This can lead to death over prolonged exposure to the gas.

ii. Carbon dioxide

This is a colourless and odourless gas. Carbon dioxide is formed when fuels undergo complete combustion in sufficient supply of oxygen or air. It is a raw material required by plants for photosynthesis. Carbon dioxide is a greenhouse gas. It is produced in large quantities into the atmosphere through human activities in industries, motor vehicle emissions and burning of wastes. Excess carbon dioxide in the atmosphere contributes to global warming that has resulted to climate change. The process of photosynthesis by plants helps to reduce the amounts of carbon dioxide in the atmosphere. This explains why people are advised to practise afforestration.

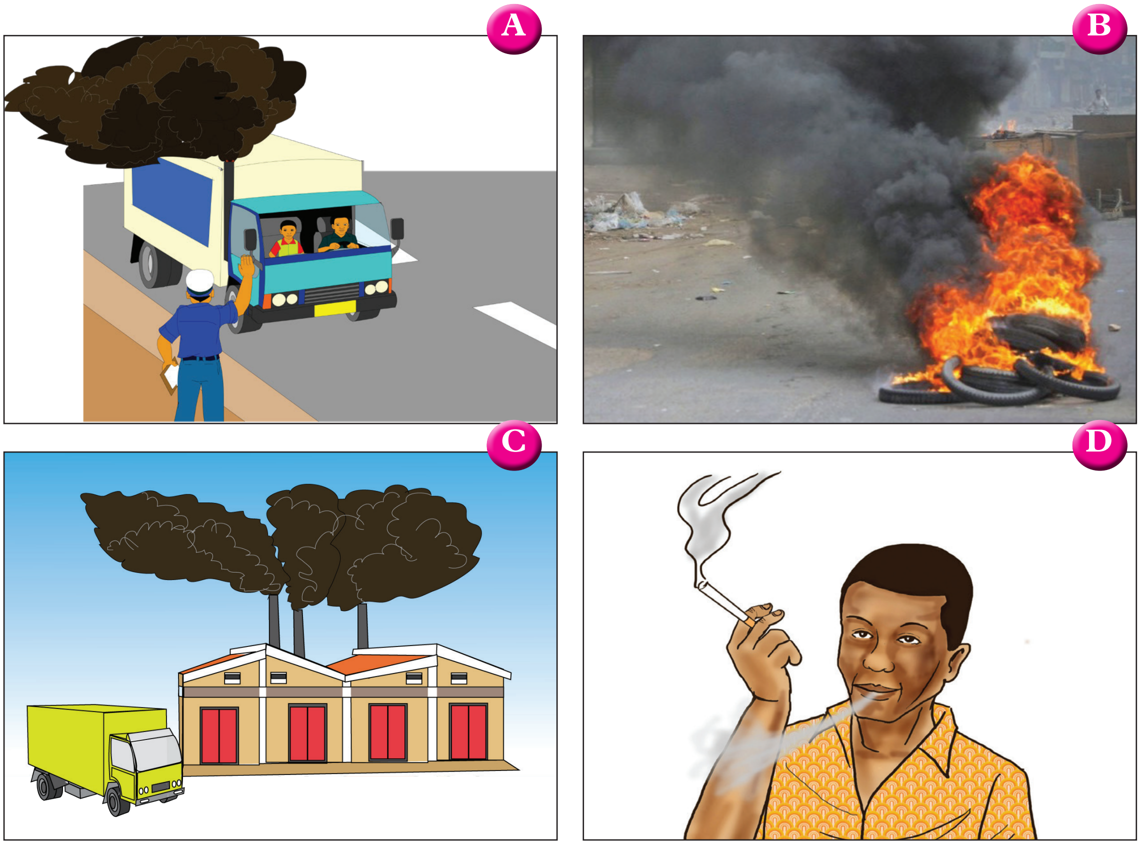


Fig 1.13: Activities that increase the amount of carbon dioxide in the atmosphere

> ***My environment, my life!***
>
> *We should protect our existing forests and plant more trees. This will minimise effects of global warming.*

Self-evaluation Test 1.4

1. Carbon monoxide is toxic. Explain.
2. a) What is global warming?
   b) Discuss the causes of global warming.
   c) What steps should be taken to minimise the effects of global warming?

b) Carbonates and hydrogen carbonates

A carbonate is a salt formed when all the hydrogen atoms in carbonic acid are replaced by a metal or an ammonium ion. Carbonic acid is formed when carbon dioxide gas dissolves in water. Examples of metallic carbonates include:

potassium carbonate, sodium carbonate, copper carbonate, zinc carbonate and magnesium carbonate.

On the other hand, a hydrogen carbonate (also called **bicarbonate**) is formed when the hydrogen atoms in carbonic acids are **partially** replaced by a metal. Some examples of metallic hydrogen carbonates are sodium hydrogen carbonate and potassium hydrogen carbonate.

1.5 Properties, preparation and testing for carbon dioxide, carbonates and hydrogen carbonates and their uses

a) Laboratory preparation of carbon dioxide

Activity 1.9

Apparatus and reagents

Flat-bottomed flask, dropping funnel, two wash bottles, delivery tubes, gas jars, cardboard covers, calcium carbonate (marble chips or limestone), dilute hydrochloric acid, distilled water, concentrated sulphuric acid.

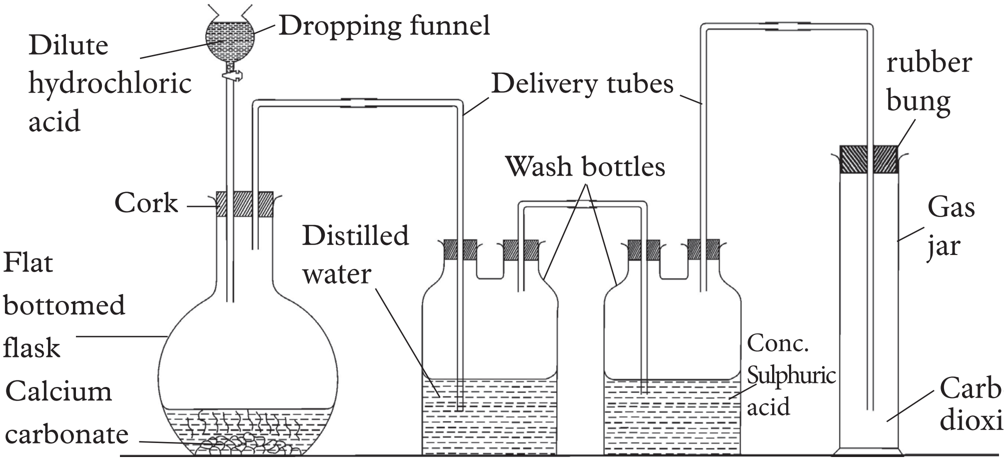


Fig 1.14: Laboratory preparation of carbon dioxide

Procedure

1. Arrange the apparatus as shown above.
2. To the limestone or marble chips, add dilute hydrochloric acid dropwise.
   - Note down your observation.
3. Collect the gas produced by downward delivery method (upward displacement of air) or over warm water.
4. Collect a few jars of the gas and preserve them for the next experiment.

Study questions

1. Why is the gas collected by downward delivery method or over warm water?
2. What is the role of water and concentrated sulphuric acid in the wash bottles?
3. Suppose dilute sulphuric acid is used instead of hydrochloric acid in the above

set up, what observation would you expect to make?

4. Write a chemical equation for the reaction that takes place.

Discussion corner!

Discuss the study questions in pairs and write a report. Present your report to the rest of the class.

I have discovered that…

When dilute hydrochloric acid comes into contact with marble chips, effervescence occurs. This means a gas is produced.

The facts

Dilute hydrochloric acid reacts with marble chips (calcium carbonate) to form calcium chloride, water and carbon dioxide gas.

Calcium carbonate + hydrochloric acid ⟶ calcium chloride + carbon dioxide + water

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

The gas is dried by passing it through concentrated sulphuric acid. It is then collected by downward delivery (upward displacement of air) as it is denser than air. Carbon dioxide can also be collected over warm water as it is only sparingly soluble in it. This is because the solubility of soluble gases in water decreases with increase in temperature.

Water in the first wash bottle is used to remove sprays of hydrochloric acid (hydrogen chloride gas). The concentrated sulphuric acid in the set up is used to dry the gas. Anhydrous calcium chloride packed in a U–tube can also be used to dry carbon dioxide.

All mineral acids react with metal carbonates to yield carbon dioxide gas. However, a reaction between dilute sulphuric acid and calcium carbonate is not feasible for preparation of carbon dioxide. This is because the reaction does not go to completion due to formation of an insoluble salt i.e. calcium sulphate that coats the rest of the carbonate preventing any further reaction.

b) Test , properties and uses of carbon dioxide

Activity 1.10

Apparatus and reagents

Nitric acid, gas-jars full of carbon dioxide, test-tubes, magnesium ribbon, a pair of tongs, candle, blue litmus paper, two water troughs, beehive shelves, calcium hydroxide solution, sodium hydroxide solution, dilute nitric acid, deflagrating spoon.

Procedure

1. Examine the gas jar full of carbon dioxide.
    - Note down the colour of the gas and its smell.
2. Pass carbon dioxide through a test-tube containing calcium hydroxide solution for a short time as shown below.

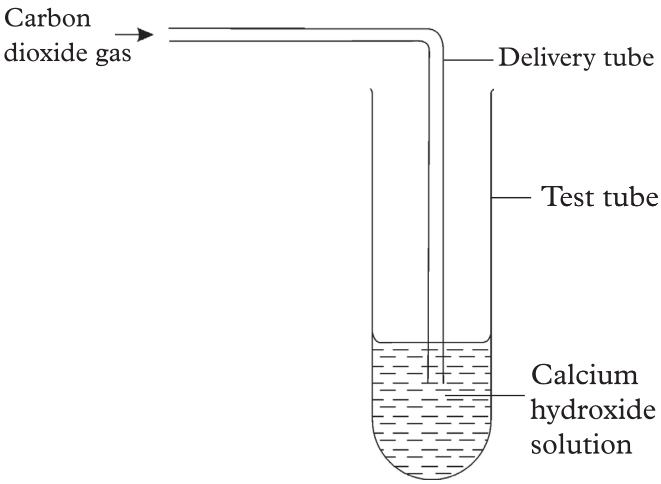


Fig1.15: Testing for carbon dioxide

- Note down your observations.

3. Continue passing more of the gas into the same test-tube until there is no further change.
    - Note down your observations.
4. Pass carbon dioxide into a test-tube containing distilled water and then dip a blue litmus paper into the resulting solution.
    - Note down the colour change on the blue litmus paper.
5. Invert a gas-jar full of carbon dioxide over a burning candle as shown below.
    - Note down what happens.

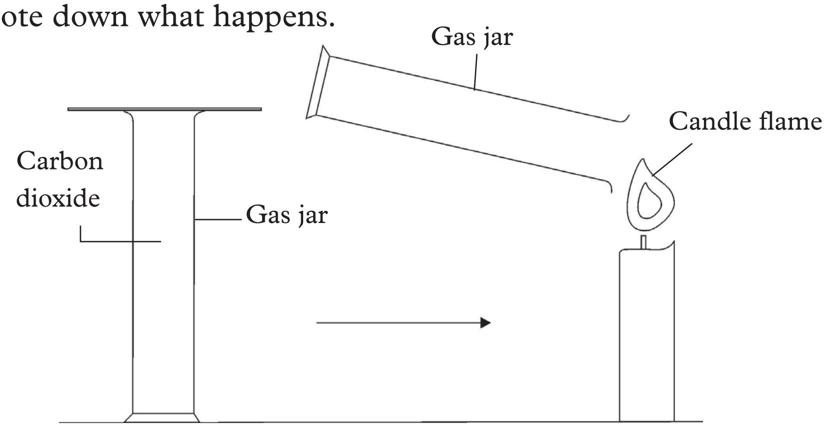


Fig 1.16: Testing whether carbon dioxide supports combustion

6. Plunge a piece of burning magnesium ribbon quickly in a gas-jar full of carbon dioxide as shown below.
   - Note down your observations.

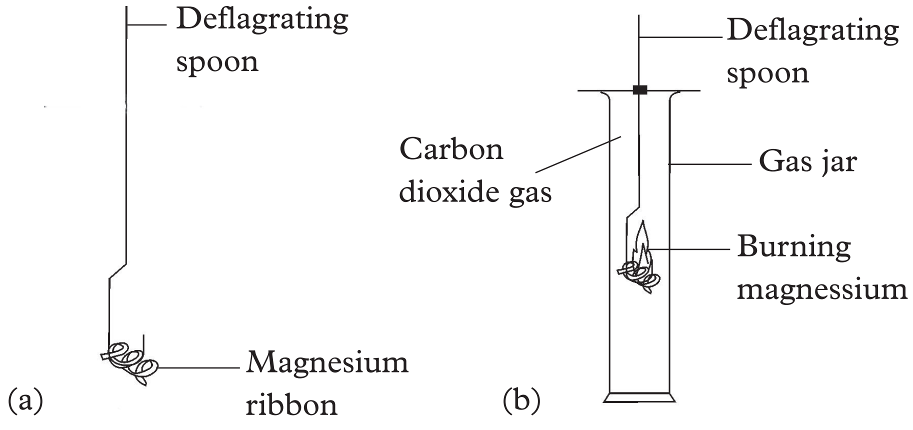


Fig1.17: Testing the effect of carbon dioxide on burning magnesium

7. Add about 50 cm^3 of dilute nitric acid to the mixture in the gas jar.
   - Note down your observations.
   - Write a chemical equation for the reaction.
8. Invert a gas jar full of carbon dioxide in a trough containing sodium hydroxide solution as shown below.
   - Shake the gas jar and wait for some time.

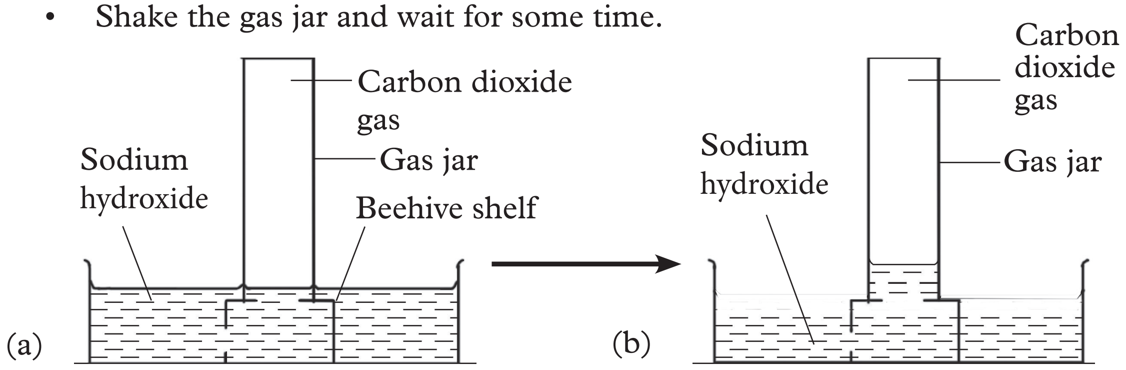


Fig 1.18: Dissolving carbon dioxide in sodium hydroxide solution

9. Repeat procedure (8) above using water instead of sodium hydroxide.
   - Note any water level changes inside the gas jar.
   - What happens to the level of sodium hydroxide solution?
   - Write chemical equations for the reactions that occur.

Study questions

1. Write balanced chemical equations for the following:
   (a) Dissolving little amount of carbon dioxide into calcium hydroxide solution.
   (b) Dissolving excess carbon dioxide gas into calcium hydroxide solution.

2. Explain why a solution of carbon dioxide in water turns blue litmus paper red.
3. Write a balanced chemical equation for the reaction between carbon dioxide and water.
4. Complete the chart below by finding the names of the compounds that correspond to A-D.

$$\boxed{CaCO_3} \xrightarrow{+\boxed{A}} \boxed{CaCl_2} + \boxed{H_2O} + \boxed{CO_2} \xrightarrow{+\boxed{B}} \boxed{CaCO_3} \xrightarrow{heat} \boxed{C} + \boxed{D}$$

5. Carbon dioxide does not support burning but magnesium continues to burn in its presence. Explain.

Discussion corner!

- Make groups of five and discuss the study questions above.
- Write your suggested answers in your note books and submit to the teacher for evaluation.

I have discovered that…

Carbon dioxide is a colourless gas that does not smell. When a little amount of carbon dioxide is dissolved in calcium hydroxide solution, a white precipitate forms. However, the precipitate dissolves on addition of excess carbon dioxide. A solution of carbon dioxide in water does not completely turn blue litmus paper red. Carbon dioxide extinguishes a burning candle. However, magnesium continues to burn in it. When a gas jar containing carbon dioxide is inverted over a beaker containing sodium hydroxide solution, the level of sodium hydroxide in the gas jar rises.

The facts

The following are some of the properties of carbon dioxide.

(a) Physical properties

- Carbon dioxide is a colourless and odourless gas.
- It is denser than air. That is why it can be collected by downward delivery.
- It is slightly soluble in water.

(b) Chemical properties

(i) Reaction of carbon dioxide with calcium hydroxide

Reaction of carbon dioxide with little amounts of calcium hydroxide leads to formation of a white precipitate. The precipitate is calcium carbonate.

Calcium hydroxide + carbon dioxide ⟶ calcium carbonate + water
$Ca(OH)_2$ (aq) + CO_2(g) ⟶ $CaCO_3$(s) + H_2O(l)

When excess carbon dioxide is passed through the white precipitate, the precipitate dissolves. This is due to the formation of soluble calcium hydrogen carbonate.

Calcium carbonate + carbon dioxide + water ⟶ calcium hydrogen carbonate
$CaCO_3$(s) + CO_2 (g) + H_2O (l) ⟶ Ca $(HCO_3)_2$ (aq)

(ii) Reaction of carbon dioxide with water

Carbon dioxide moderately reacts with water. The solution formed is slightly acidic i.e. weak carbonic acid is formed but in very small quantities.

Carbon dioxide + water ⟶ carbonic acid
CO_2 (g) + H_2O (l) ⟶ H_2CO_3 (aq)

The carbonic acid formed reacts with sodium hydroxide to form sodium carbonate and water only.

$$H_2CO_3 \text{ (aq)} + 2NaOH \text{ (aq)} \longrightarrow Na_2CO_3 \text{ (aq)} + 2H_2O \text{ (l)}$$

The sodium hydroxide level rises inside the gas jar after some minutes to occupy the space initially occupied by the reacted carbon dioxide.

Sodium hydroxide + carbon dioxide ⟶ sodium carbonate + water
$2NaOH$ (aq) + CO_2 (g) ⟶ Na_2CO_3 (aq) + H_2O (l)

Since carbonic acid is dibasic, it can form both normal and an acidic salt. When carbon dioxide is passed for a long time in concentrated solution of sodium hydroxide, a further reaction occurs forming sodium hydrogen carbonate some of which precipitates out.

Sodium carbonate + carbon dioxide + water ⟶ sodium hydrogen carbonate

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

Note: Potassium hydroxide solution reacts in a similar way but its hydrogen carbonate is more soluble and does not precipitate. That is why potassium hydroxide solution is preferred as an absorbent for carbon dioxide.

(iii) Carbon dioxide and combustion

Carbon dioxide extinguishes a burning candle showing that it does not support combustion.

(iv) Reaction of carbon dioxide with magnesium

A burning magnesium ribbon continues to burn inside a gas jar full of carbon dioxide. Black carbon particles form on the sides of the jar, together with white ash of magnesium oxide.

The addition of dilute nitric acid enables the black carbon particles to be seen more clearly because the acid reacts with magnesium oxide forming magnesium nitrate solution and water. Although carbon **dioxide** does not support combustion, magnesium continues to burn in it. The heat produced by the intense burning of magnesium is sufficient to decompose the carbon dioxide into constituent elements: carbon and oxygen. This oxygen produced supports the continued combustion to magnesium oxide.

Carbon dioxide $\xrightarrow{\text{heat}}$ carbon + oxygen

$$CO_2(g) \xrightarrow{\text{heat}} C(s) + O_2(g)$$

Magensium + oxygen $\longrightarrow$ magnesium oxide

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

The general equation for the reaction can then be written as follows:

Magnesium + carbon dioxide $\longrightarrow$ magnesium oxide + carbon

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

Magnesium being more reactive than carbon displaces it from oxygen.

Uses of carbon dioxide

Research activity

1. Using textbooks and from the internet, research on the uses of carbon dioxide then write down your findings.
2. Discuss your findings with a friend.
3. Compare your findings with the ones outlined below.

1. Carbon dioxide is used in the manufacture of sodium carbonate used in baking of cakes and bread among others. It is also an ingredient in some health salts to relieve constipation.

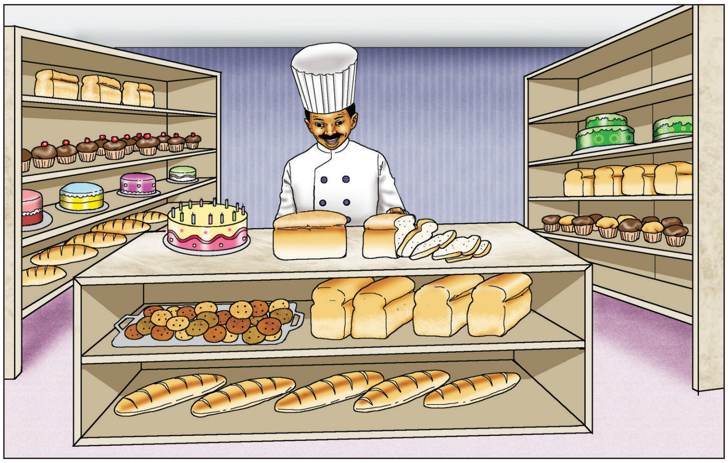


Fig 1.19: Bakery

2. **Carbonated drinks**
   Carbon dioxide is used as preservative in the production of mineral water and carbonated drinks like coca-cola (also called aerated or effervescence drinks). Soda-water is a solution of carbon dioxide in water under pressure. It is later sweetened, flavoured and sometimes coloured. The dissolved carbon dioxide gives it a pleasant taste. Figure 1.20 shows a girl opening a soda bottle. What do you think will happen to the contents of the bottle when the bottle top is removed?


Fig 1.20: A girl opening a soda

3. **As a refrigerant**
   Solid carbon dioxide commonly known as **dry ice** is a preferred refrigerant to ordinary ice (solid water) see figure 1.21. This is because it sublimes at room temperature forming gaseous carbon dioxide and therefore leaves no residue unlike ordinary ice.


Fig 1.21: Dry ice

4. **Fire extinguishers**
   Some fire extinguishers contain sodium hydrogen carbonate or sodium carbonate solution and sulphuric acid. When mixed by inversion or pressing the plunger, carbon dioxide is produced due to the reaction between carbonate with sulphuric acid. In Fig 1.22 a person is trying to put out fire by using carbon dioxide fire extinguisher.

   Sodium hydrogen carbonate + sulphuric acid ⟶ sodium sulphate + carbon dioxide + water

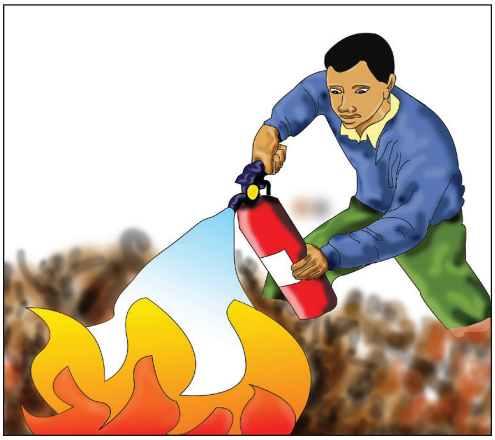


Fig 1.22: Using a fire extinguisher to put out a fire

$2NaHCO_3\ (aq) + H_2SO_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2CO_2\ (g) + 2H_2O(l)$

Sodium carbonate + sulphuric acid ⟶ sodium sulphate + carbon dioxide + water

$Na_2CO_3\ (aq) + H_2SO_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + CO_2\ (g) + H_2O\ (l)$

A high pressure is set which forces a froth of carbon dioxide through the jet. It is directed onto the burning substances.
In fire extinguishers that have carbon dioxide stored under pressure, the carbon dioxide is released through a nozzle.
Carbon dioxide does not support combustion and since it is denser than air, it forms a blanket on the burning substance thereby cutting off the supply of oxygen from the air.

5. Making rain during drought or in areas of little rain
   In making rain, dry ice (solid carbon dioxide) is spread in the clouds to accelerate the condensation process. Small aircrafts are sometimes used to spread dry ice in the sky.
6. It is also used to transfer heat energy from certain types of nuclear reactors.

Fig1.23: Aircraft dropping dry ice

Further activity

Check the type of fire extinguishers you have in your school. Inquire from your teacher how they are used.

Self-evaluation Test 1.5

1. State two physical properties of carbon dioxide gas.
2. Write equation to show how sodium hydroxide reacts with carbon dioxide.
3. Describe what happens when carbon dioxide is passed through a solution of calcium hydroxide (lime water). Use equations in your explanations.
4. Carbon dioxide was passed through burning magnesium. State what was likely to be observed. Write equation for the reaction.
5. The product in (4) was dissolved in water and a red litmus paper dipped into the solution formed.
   (a) State what was observed.
   (b) Write balanced equations for the reaction.

1.6 Preparation, properties and test for carbonates and hydrogen carbonates

a) Preparation of metal carbonates

Activity 1.11

Apparatus and reagents

Calcium chloride, lead (II) nitrate, sodium carbonate, filter funnel, filter paper, conical flask, beakers, stirring rod, spatula, measuring cylinder.

Procedure

1. Place 5 g of calcium chloride in a beaker and add to it 20 cm^3 of water. Stir until all the solid dissolves.
2. Put 5 g of sodium carbonate in another beaker and add to it 20 cm^3 of distilled water, stir the mixture to form a uniform solution.
3. Add the sodium carbonate solution to the calcium chloride solution and stir the mixture. Allow to settle.
   - Note down your observations.
4. Filter off the precipitate obtained and wash it with distilled water.
5. Dry the precipitate between filter papers.
6. Repeat the procedure steps 1-5 using lead (II) nitrate in place of calcium chloride.

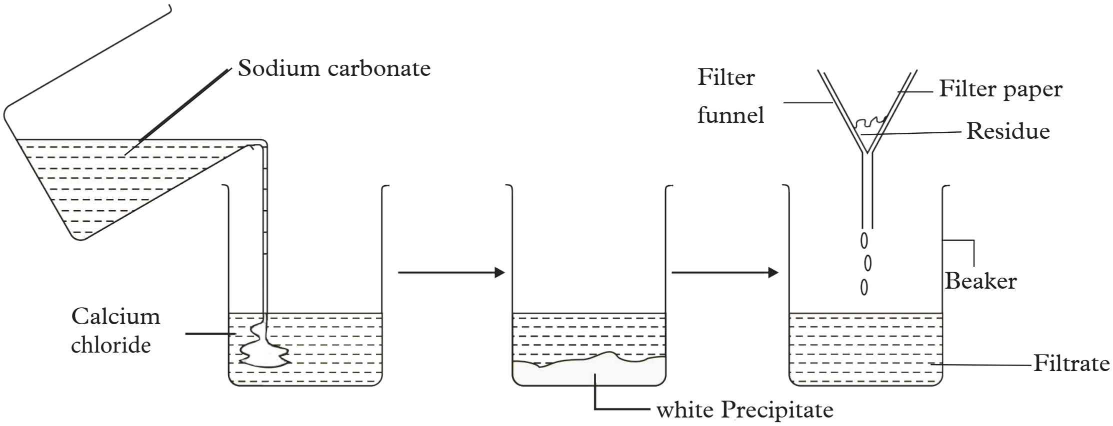


Fig 1.24: Preparation of calcium carbonate

Study questions

1. What observation did you make when sodium carbonate solution was added to :
   i. Calcium chloride solution?
   ii. Lead (II) nitrate solution?
2. Write balanced chemical equations for the reactions that occurred during the experiments.
3. Which other solution can be used in place of sodium carbonate solution?

Discussion corner!

- In your study groups, discuss the results of these two experiments and the study questions above.
- Write a joint report on your discussions and present it to the class.

I have discovered that…

When sodium carbonate solution is added to calcium chloride solution, a white precipitate forms. A similar precipitate is formed when lead (II) nitrate solution is mixed with sodium carbonate.

The facts

Sodium carbonate reacts with calcium chloride to form sodium chloride solution and insoluble calcium carbonate (white precipitate). Similarly, sodium carbonate reacts with lead (II) nitrate solution to form sodium nitrate solution and an insoluble lead (II) carbonate (white precipitate). The equations for the reactions are as follows:

$$CaCl_2(aq) + Na_2CO_3\,(aq) \longrightarrow CaCO_3(s) + 2NaCl\,(aq)$$

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

Filtering is done to separate the solid carbonate from the rest of the solution. The residue is then dried between filter papers or placed outside in the sun. Other carbonates such as magnesium carbonate and zinc carbonate can be prepared using the same method.

b) Preparation of metal hydrogen carbonate

Activity 1.12

Apparatus and reagents

Conical flask, carbon dioxide generator, delivery tubes, cold concentrated solution of sodium hydroxide, filter funnel, filter papers.

Procedure

1. Bubble carbon dioxide in excess through a cold concentrated solution of sodium hydroxide for 10 minutes.
2. Filter the resulting mixture.
3. Leave the residue undisturbed for sometime.
4. Collect the crystals formed.

Study questions

1. What do you observe when carbon dioxide is bubbled through sodium hydroxide solution for 10 minutes?
2. Write the equation for the reaction that occurs.

Discussion corner!

In groups of four discuss the study questions above and present your answers to the whole class.

I have discovered that…

When carbon dioxide is bubbled through sodium hydroxide for some time, a white precipitate is formed. After the residue is filtered and dried, crystals of a salt are formed.

The facts

In the laboratory, sodium hydrogen carbonate is prepared by passing carbon dioxide through cold concentrated solution of either sodium hydroxide or sodium carbonate. The equations for the reactions that take place are as follows:

$$2NaOH\,(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$
$$Na_2CO_3\,(aq) + H_2O\,(l) + CO_2\,(g) \longrightarrow 2NaHCO_3(aq)$$

The Solvay process

Since the 1890s, sodium carbonate has been produced using Solvay process. Solvay process is a continuous process whereby limestone ($CaCO_3$) is used to produce carbon dioxide, which then reacts with ammonia dissolved in brine (concentrated sodium chloride) to produce sodium carbonate.

The steps in solvay process are:

- Brine purification
- Sodium hydrogen carbonate formation
- Sodium carbonate formation
- Ammonia recovery

Environmental issues involved in production of sodium carbonate

- Solid wastes such as calcium chloride, sand, clay and unburnt calcium carbonate.
- Air pollution- ammonia released to the atmosphere is toxic.
- Thermal pollution- due to release of large quantities of heat.

Sodium carbonate has a number of uses but its most common use is in the production of glass.

Uses	**Process notes**
Glass making	A mixture of sodium carbonate, calcium carbonate and silicon dioxide is used for making window or bottle glass.
Water softening agent	Carbonate ion (CO_3^{2-}) from dissolved sodium carbonate can precipitate magnesium (Mg^{2+}) and calcium (Ca^{2+}) ions from hard water. The reaction leads to the formation of insoluble carbonates, removing them from water hence the water becomes soft. For this reason sodium carbonate is also known as washing soda.
Baking soda production	Baking soda or sodium bicarbonate ($NaHCO_3$) is used in food preparation and in fire extinguishers.
Wool processing	Sodium carbonate removes grease from wool and neutralises acidic solutions.

c) Chemical properties of carbonates and hydrogen carbonates

i) Action of heat on carbonates and hydrogen carbonates

Activity 1.13

Apparatus and reagents

Test-tubes, spatula, source of heat, calcium hydroxide solution, distilled water, red and blue litmus papers, copper (II) carbonate, magnesium carbonate, zinc carbonate, lead (II) carbonate, sodium carbonate, potassium carbonate, sodium hydrogen carbonate, calcium hydrogen carbonate, ammonium carbonate.

Procedure

1. Place a spatula full of the carbonate or hydrogen carbonate salts mentioned above in different test-tubes then heat.
2. Test any gas formed with calcium hydroxide solution and moist litmus papers (red and blue).

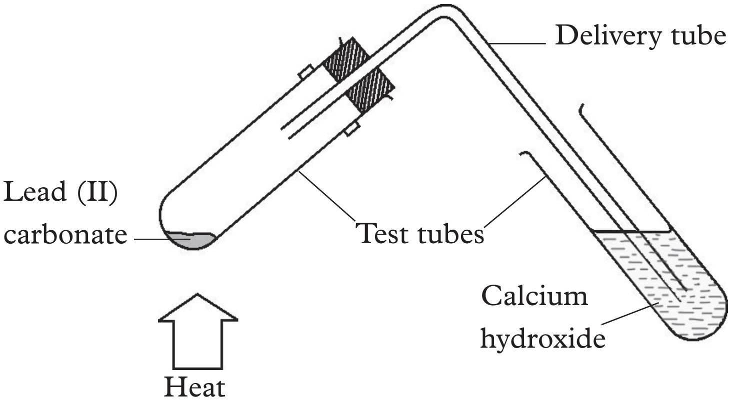


Fig1.25: Heating carbonates and hydrogen carbonates

- Note down the colour of the residue when hot and cold for each residue.

3. Record all your observations in a table like the one shown below.

Table 1.1: Results for action of heat on carbonates and hydrogen carbonates

Name of carbonate/ hydrogen carbonate	**Colour before heating**	**Colour after heating**	**Colour after cooling**	**Effect of gas evolved on calcium hydroxide solution**	**Effect of gas evolved on moist red and blue litmus papers**
Sodium carbonate					
Potassium carbonate					
Magnesium carbonate					
Zinc carbonate					
Lead carbonate					
Copper carbonate					
Ammonium carbonate					
Sodium hydrogen carbonate					
Calcium hydrogen carbonate					

Study questions

1. Write balanced chemical equations for all the reactions that took place.
2. Identify carbonates that behaved in a similar way.
3. Which carbonate evolved a gas that turned red litmus paper blue?
4. Which carbonate produced a residue that showed colour changes?
5. Which carbonates were not affected by heat?

Discussion corner!

- In your groups, discuss the results of the experiment.
- Classify the carbonates and hydrogen carbonates basing on similarity of their behaviour during and after heating.
- Write a report on your conclusions and present it to the teacher for evaluation.

I have discovered that…

Sodium carbonate and potassium carbonate were not affected by heat. However, when the carbonates of magnesium, zinc, lead, copper as well as sodium hydrogen carbonate and calcium hydrogen carbonate were heated they produced a gas that formed a white precipitate with calcium hydroxide solution. The gas also turned moist blue litmus paper red. Ammonium carbonate produced two gases, one turned blue litmus paper red while the other turned red litmus paper blue.

The facts

Some carbonates decompose on heating while others do not. Carbonates of magnesium, zinc, lead and copper decompose on heating to produce respective metal oxide and carbon dioxide gas. Carbon dioxide gas produced forms a white precipitate with calcium hydroxide solution. The gas also turns moist blue litmus paper red as it is acidic. The metal oxides formed are of different colours as shown in the equations below.

Copper (II) carbonate $\xrightarrow{\text{heat}}$ copper(II) oxide + carbon dioxide

$CuCO_3(s) \xrightarrow{\text{heat}} CuO(s) + CO_2(g)$

(green) (black)

Magnesium carbonate $\xrightarrow{\text{heat}}$ magnesium oxide + carbon dioxide

$MgCO_3(s) \xrightarrow{\text{heat}} MgO(s) + CO_2(g)$

(white) (white)

Zinc carbonate $\xrightarrow{\text{heat}}$ zinc oxide + carbon dioxide

$ZnCO_3(s) \xrightarrow{\text{heat}} ZnO(s) + CO_2(g)$

(white) (white)

Lead (II) carbonate $\xrightarrow{\text{heat}}$ lead(II) oxide + carbon dioxide

$PbCO_3(s) \xrightarrow{\text{heat}} PbO(s) + CO_2(g)$

(white when hot) (yellow on cooling)

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

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

Copper (II) oxide is black while oxides of magnesium and calcium are white. Zinc oxide is **yellow** when hot and **white** when cold. Lead (II) oxide is **red-brown** when hot and yellow when cold. Carbonates of sodium and potassium are not affected by heat.

Sodium hydrogen carbonate and calcium hydrogen carbonate decompose on heating to form a carbonate, water and carbon dioxide.

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

$$2NaHCO_3(s) \xrightarrow{\text{heat}} Na_2CO_3(s) + CO_2\ (g) + H_2O\ (l)$$

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

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

Note: When calcium hydrogen carbonate is strongly heated, calcium carbonate formed decomposes to form calcium oxide and carbon dioxide gas.

Ammonium carbonate decomposes on heating to form ammonia gas, water and carbon dioxide.

Ammonium carbonate $\xrightarrow{\text{heat}}$ ammonia + carbon dioxide + water

$$(NH_4)_2CO_3(s) \xrightarrow{\text{heat}} 2NH_3(g) + CO_2(g) + H_2O(l)$$

Ammonia and carbon dioxide gases are liberated at the same time. If moist red and blue litmus papers are put together at the mouth of the test-tube when ammonium carbonate is heated, ammonia gas is detected first (red litmus paper turns blue).

ii) Action of dilute acids on carbonates and hydrogen carbonates

Activity 1.14

Apparatus and reagents

Test tubes in a rack, spatula, calcium hydroxide solution, calcium carbonate, magnesium carbonate, zinc carbonate, lead (II) carbonate, copper carbonate, sodium carbonate, ammonium carbonate, sodium hydrogen carbonate, calcium hydrogen carbonate, dilute nitric acid, dilute sulphuric acid, dilute hydrochloric acid.

Procedure

1. In different test-tubes, put a spatulaful of the above carbonates and hydrogen carbonates and label them accordingly.

2. Add dilute hydrochloric acid into each test tube and test the gas evolved with calcium hydroxide solution.

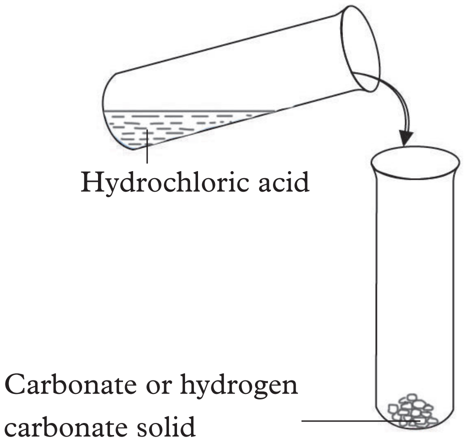


Fig.1.26: Reaction of carbonates or hydrogen carbonates with acid

- Note down the observation made in test tubes A and B.

3. Repeat the experiment using sulphuric acid and then dilute nitric acid on all the carbonates and hydrogen carbonates.
    - Note down your observations in both test tubes A (reaction of acid + carbon/hydrogen carbonate and B reaction of gas produced + hydrogen carbonate solution).
4. Record your observations in a table like the one shown below.

Table 1.2 Reaction of carbonate or hydrogen carbonates with acid

Carbonate/hydrogen carbonate	**Acid**	**Observation in test tube A**	**Observation in test tube B**
Sodium hydrogen carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		
Sodium carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		
Calcium hydrogen carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		

Copper (II) carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		
Magnesium carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		
Zinc carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		
Ammonium carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		
Lead (II) carbonate	Hydrochloric acid		
	Sulphuric acid		
	Nitric acid		

Study questions

1. Explain the observations made when dilute hydrochloric acid and sulphuric acid were each added to lead (II) carbonate.
2. What happened when dilute sulphuric acid was added to calcium carbonate solution?
3. Write balanced chemical equations for all the reactions that took place.

Discussion corner!

- In your study groups, discuss the results obtained in this experiment.
- Write your report on your findings and present it to the teacher for evaluation.

I have discovered that…

When an acid is added to any carbonate or a hydrogen carbonate, effervescence occurs, showing that a gas is produced. When the gas is dissolved in calcium hydroxide solution, a white precipitate is formed. However, on reacting lead (II) carbonate with either dilute sulphuric acid or hydrochloric acid, the reaction starts but stops after sometime. The same case happens when dilute sulphuric acid reacts with calcium carbonate.

The facts

Metal carbonates react with dilute mineral acids to produce a salt, water and carbon dioxide gas.

Metal carbonate/hydrogen carbonate + dilute acid ⟶ Salt + water + carbon dioxide

The reaction between dilute sulphuric acid and calcium carbonate does not proceed to completion. This is due to formation of an insoluble coat (salt) of calcium sulphate that forms the impervious layer preventing any further reaction. Similarly, the reaction between lead (II) carbonate and dilute sulphuric acid as well as hydrochloric acid does not proceed to completion. This is due to formation of insoluble salts of lead (II) sulphate and lead (II) chloride respectively.

iii) Solubility of carbonates and hydrogen carbonates in water

Activity 1.15

Apparatus and reagents

Test tubes, distilled water, spatula, sodium carbonate, potassium carbonate, ammonium carbonate, sodium hydrogen carbonate, calcium hydrogen carbonate, calcium carbonate, zinc carbonate, lead (II) carbonate, magnesium carbonate.

Procedure

1. Place 0.5g of sodium carbonate in a test tube and add 5cm^3 of distilled water.
2. Label the test tube and shake its contents thoroughly.
   - Note down your observations.
3. Repeat the experiment using the other carbonates and hydrogen carbonates provided and record your results in a table like the one shown below.

Table 1.3: Solubility of carbonates and hydrogen carbonates in water

Carbonate/hydrogen carbonate	**Solubility in water Use (√) for soluble or (×) for insoluble**
Sodium carbonate	
Potassium carbonate	
Ammonium carbonate	
Sodium carbonate	
Sodium hydrogen carbonate	
Calcium hydrogen carbonate	
Calcium carbonate	
Zinc carbonate	
Lead carbonate	
Magnesium carbonate	

Study questions

1. Which carbonates are:
   (a) soluble in water?
   (b) insoluble in water?
2. Which hydrogen carbonates are:
   (a) soluble in water?
   (b) insoluble in water?

Discussion corner!

- Discuss your results and the study questions above in your groups.
- Present your conclusions to the whole class.

I have discovered that…

Carbonates of sodium, potassium and ammonium dissolve in water. All other carbonates do not dissolve in water. Sodium hydrogen carbonate also dissolves in water.

The facts

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

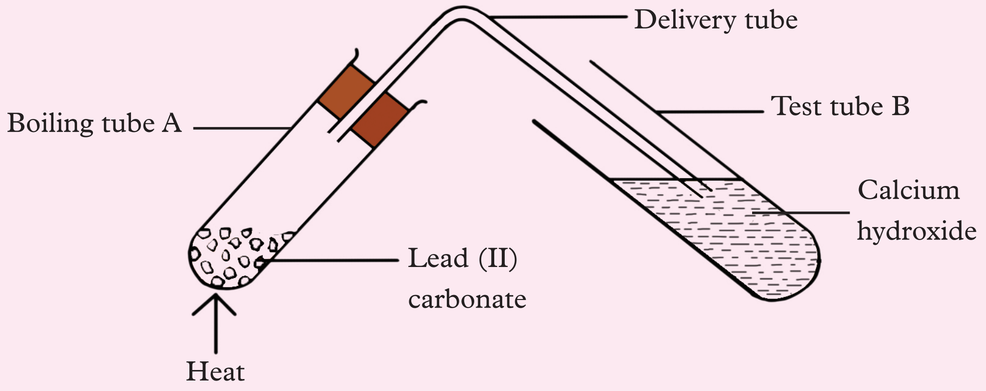
Research activity

Research and do a presentation on the uses of carbon compounds and their environmental issues under the following sub-headings.

1. Carbon dioxide in fire extinguishers.
2. Calcium carbonate in the manufacture of cement.
3. Toxicity of carbon monoxide.
4. Greenhouse effect and global warming.

Self-evaluation Test 1.6

1. 2 cm^3 of dilute nitric acid was added to 0.5g of zinc carbonate in a test tube. The gas produced was then dissolved in calcium hydroxide solution.
   (a) State the observations made.
   (b) Write stoichiometric equations for the reactions that occur.
   (c) Draw a set up of the apparatus that can be used to perform this experiment.
2. The following set up was used to investigate the effect of heat on lead (II) carbonate.



a) State the observations made in boiling tube A.
b) What observation is made in test tube B?
c) Write an equation for the reaction that occurs in test tube:
   i) A
   ii) B

3. Name two carbonates that are:
   a) Soluble in water.
   b) Insoluble in water.

4. A senior 3 student separately heated three solids A, B and C suspected to be carbonates. He tested the gases evolved (if any) with moist blue and red litmus papers. The following are the observations he made.

Carbonate	**Observation**
A	Blue litmus paper turned red No effect on red litmus paper
B	Red litmus paper turned blue Blue litmus paper turned red
C	No effect on litmus papers

   (a) Suggest the possible identities of carbonates A and C.
   (b) Explain the observations made with carbonate B.

5. a) A solution reacts with crushed egg shells to give a gas that turns lime water milky. Name the gas that is evolved and give a brief explanation for this observation.
   b) Describe a chemical test that you could carry out to confirm the gas evolved in 5(a) above.
   c) Name the gas evolved when dilute hydrochloric acid reacts with sodium hydrogen carbonate. How is it recognised?

6. Copy and complete the chemical equations below.
   a) $H_2SO_4(aq) + Na_2CO_3(aq) \longrightarrow$
   b) $HCl(aq) + Na_2CO_3(aq) \longrightarrow$
   c) $HNO_3(aq) + Na_2CO_3(aq) \longrightarrow$

d) $H_2SO_4(aq) + CaCO_3(aq) \longrightarrow$

e) $HCl(aq) + NaHCO_3(aq) \longrightarrow$

7. Describe what would happen if the following carbonates were strongly heated: zinc carbonate, copper (II) carbonate and lead (II) carbonate. Write equations for the reactions that would take place.

1.7 Environmental issues of carbon dioxide and carbon monoxide

Both carbon dioxide and carbon monoxide gases have adverse effects on the environment.

a) Environmental issues of carbon dioxide

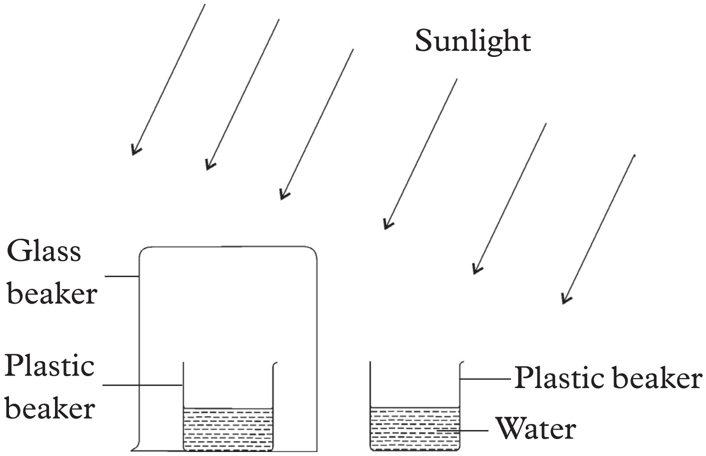
Apparatus and reagents

Activity 1.16

A large glass beaker or glass box, two small plastic beakers full of water, thermometer.

Procedure

1. Cover one of the small plastic beakers full of water with the big glass beaker or glass box as shown in Fig 1.27A below.
2. Leave the other small plastic beaker with water in the open air as shown in Fig 1.27B.
3. Put the two beakers under direct sunlight.



A *Fig 1.27* B

4. At the end of the day, measure the temperature of the water in the two beakers.
    - Record the temperatures.

Study question

In which case did you record a higher temperature? Why?

Discussion corner!

In pairs, discuss the results of the experiment and present your findings to the rest of the class.

I have discovered that…

Water in the beaker that was covered with a glass box had a higher temperature at the end of the day.

The facts

Some gases are able to trap heat in the atmosphere keeping the earth's surface warmer. Such gases are called **'greenhouse'** gases. Carbon dioxide is an example of a greenhouse gas. Others include methane, chlorofluorocarbons and the oxides of nitrogen. As the amount of carbon dioxide in the atmosphere increases, the amount of heat retained by the atmosphere increases and the Earth becomes hotter. This is called **greenhouse effect** and it has led to the increase of temperatures in the world a phenomenon known as **global warming**.

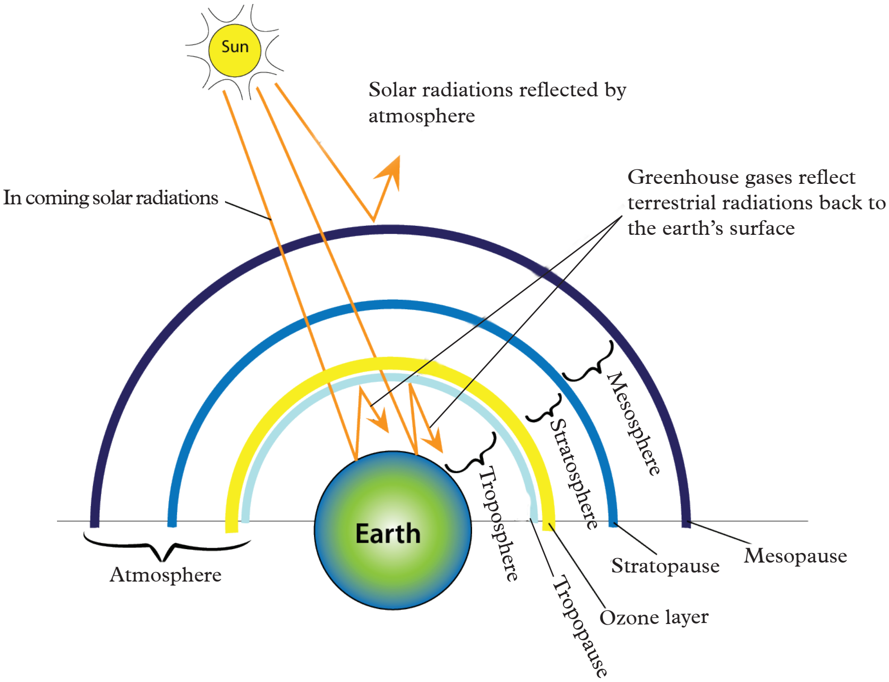


Fig 1.28: Greenhouse effect

Human activities such as burning of fossil fuels and clearing forest land for agriculture and settlement have been blamed for the increase in amounts of carbon dioxide in the atmosphere. This has led to global warming which has in turn has caused unpredictable harsh climatic conditions of several regions. It is believed that with time, global warming will cause melting of polar ice. This will lead to coastal flooding that is dangerous to the lives of human beings, animals and plants.

b) Environmental issues of carbon monoxide

Carbon monoxide is formed through incomplete combustion of charcoal, coal, natural gas (methane) or firewood in poorly ventilated places. It also forms in improperly ventilated rooms in industries.

Note: Carbon monoxide is not only a pollutant but also deadly, poisonous and toxic gas. Being colourless and odourless it is therefore not easy to detect. As mentioned earlier carbon monoxide competes with oxygen for the binding sites in haemoglobin (pigment that transports oxygen in the body). When carbon monoxide binds to haemoglobin molecules, little or no oxygen binds. This starves body tissues and cells of oxygen and may eventually cause death.

$$\text{Haemoglobin} + \text{carbon monoxide} \longrightarrow \text{carboxyhaemoglobin}$$

The compound **carboxyhaemoglobin** is much more stable than oxyhaemoglobin (compound formed when oxygen binds with haemoglobin). This interferes with the transport of oxygen to various cells and tissues of the body.

If noticed early, carbon monoxide poisoning can be reversed by giving the victim air enriched with oxygen and carbon dioxide. In severe cases of poisoning, a blood transfusion is essential.

Self-evaluation Test 1.7

1. Carbon monoxide is a poisonous gas. Explain.
2. What is global warming?
3. Name 3 greenhouse gases.
4. Which of the following is not a major contributor to the greenhouse effect?
   (A) Carbon dioxide
   (B) Carbon monoxide
   (C) Chlorofluorocarbons
   (D) Methane gas
   (E) Nitrous oxide
5. Which of the following is not likely a result of global warming?
   (A) Rising sea levels.
   (B) Increased agricultural productivity worldwide.
   (C) Increased storm frequency and intensity.
   (D) Increased coastal flooding.
6. Which gas exists in the highest concentrations in the Earth's atmosphere?

1.8 Hard and soft water

Activity 1.17

Apparatus and reagents

Distilled water, rain water, water from a borehole, bar soap, three basins, three pieces of dirty clothes.

Procedure

1. Put the various kinds of water from each of the three sources in different basins.
2. Using soap, wash the dirty clothes seperately in each basin.
   - Note down the water that forms lather easily.

Study questions

1. In which water did soap lather easily?

Discussion corner!

Discuss your findings in pairs and come up with a conclusion.

I have discovered that…

Soap lathers easily in distilled water and rain water than in water from a borehole.

The facts

Hard water is water that contains salts of calcium and magnesium principally as bicarbonates, chlorides and sulphates. Hard water does not readily form lather with soap. Soft water on the other hand readily lather with soap. Distilled water and rain water are examples of soft water while water from borehole is an example of hard water.

Water hardness is caused by dissolved calcium and magnesium salts. Hard water therefore contains calcium ions (Ca^{2+}) and /or magnesium ions (Mg^{2+}). Soft water on the other hand lacks these ions in solution.

Temporary and permanent water hardness

Research activity

1. Using textbooks, journals from the library and from internet, research on the causes and how to remove temporary and permanent water hardness.
2. Write a report on your findings and present it in class.
3. Compare your findings with the ones outlined below.

Water hardness can either be temporary or permanent. **Temporary water hardness** is caused by dissolved calcium hydrogen carbonate and magnesium hydrogen carbonate. **Permanent water hardness** on the other hand is caused by dissolved calcium sulphate and magnesium sulphate.

Techniques of removing water hardness

Activity 1.18

Apparatus and reagents

Water from a borehole or river, source of heat, boiling pot, calcium hydroxide, sodium carbonate, three basins and soap, pieces of dirty clothes.

Procedure

1. Divide the water from a borehole into three portions.
2. Put one portion in a boiling pot then heat and cool, transfer to basin labelled A.
3. Put the remaining portions into two basins, label them as B and C.
4. To basin B add two spoonful of calcium hydroxide and mix thoroughly.
5. To basin C add two spoonful of sodium carbonate and mix thoroughly.
6. Add some soap in basin A, B and C. Try washing the dirty clothes in each portion. What happens?

Study questions

1. In which water did soap lather easily?
2. Suggest the methods used above to remove water hardness.
3. In which method was permanent water hardness removed?

Discussion corner!

Discuss your findings in pairs and come up with a conclusion.

a) Removing temporary water hardness

This type of hardness can be removed using the following methods:

- Boiling
- Distillation
- Addition of calcium hydroxide

i) Boiling

As mentioned earlier, boiling removes the temporary water hardness by decomposing the dissolved calcium hydrogen carbonate or magnesium hydrogen carbonate to their respective carbonates i.e.

$$Mg(HCO_3)_2\ (aq) \xrightarrow{heat} MgCO_3(s) + CO_2(g) + H_2O(l)$$

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

ii) Distillation

Distillation removes all dissolved solids. This method is used to prepare the distilled water we use in our chemistry laboratory. However, this method of water softening is too expensive to be used on a large scale.

iii) Addition of calcium hydroxide

When calcium hydroxide solution is added to temporary hard water, precipitation of the calcium and magnesium carbonate occurs. Calcium and magnesium ions are therefore removed from solution:

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

$$Mg\ (HCO_3)_2(aq) + Ca(OH)_2(aq) \longrightarrow MgCO_3(s) + CaCO_3(s) + 2H_2O(l)$$

Note: It is important to add just enough of calcium hydroxide solution because when in excess, it would itself cause hardness.

b) Removing permanent water hardness

It is important to note that all methods used to remove permanent water hardness can as well remove temporary water hardness. Some of these methods include:

i) Addition of washing soda (sodium carbonate)

When sodium carbonate is added to hard water containing dissolved calcium sulphate and magnesium sulphate salts, their corresponding carbonates are precipitated.

$$CaSO_4(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + Na_2SO_4(aq)$$

$$MgSO_4(aq) + Na_2CO_3(aq) \longrightarrow MgCO_3(s) + Na_2SO_4(aq)$$

Notice that some sodium salts remain in solution but this does not affect the water in any way.

ii) Use of ion exchange resins

This involves passing the hard water through a column filled with a particular type of material **(sodium permutit)** which contains sodium ions. The calcium and magnesium ions are removed and remain in the column while the sodium ions in the column replace the calcium ions in the water.

Note: Sodium ions in solution do not cause water hardness.

$$Na_2X\,(aq) + Ca^{2+}\,(aq) \longrightarrow CaX(s) + 2Na^{+}(aq)$$

$$Na_2X\,(aq) + Mg^{2+}\,(aq) \longrightarrow MgX(s) + 2Na^{+}(aq)$$

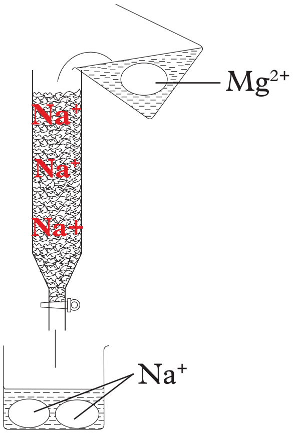


Fig 1.29: Permutit process

When all the sodium ions on the permutit column have been exchanged with calcium or magnesium ions i.e. resin now contains ions of calcium and magnesium, it is said to be **discharged**. The permutit, can be "regenerated" or "recharged" by passing concentrated sodium chloride solution (brine) through it. Another ion–exchange process takes place.

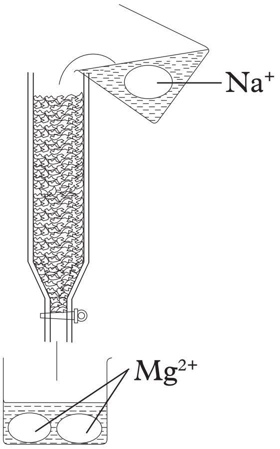


Fig1.30:Regenerating an ion exchange column

Research activity

1. Using textbooks and the internet, find out the advantages and disadvantages of hard water.
2. Discuss your findings with a friend.
3. Write a report and present it in class.
4. Compare the advantages and disadvantages of hard water with those of softwater.

Table 1.4 Advantages and disadvantages of hard water

Advantages	Disadvantages
The calcium and magnesium ions present in hard water are important for strong development of teeth and bones.	It wastes soap as it does not lather easily.
Helps in formation of an insoluble coating of lead in pipes made of lead. This prevents some lead ions from dissolving in the water which would results in lead poisoning.	Since it does not lather easily, it leads to formation of an insoluble substance called **scum** thus leaves clothes with unpleasant marks after washing.
Good for formation of animal shells. For example, snails obtain calcium from hard water.	It forms boiler scale inside electric kettles and hot water pipes. This requires a lot of heat energy to heat up hence expensive in terms of energy required.
Calcium and magnesium ions are important for plant growth. Hard water provides these.	Deposits of calcium and magnesium carbonates can reduce the capacity of pipes from boilers and also reduce the efficiency of boilers.

Self-evaluation Test 1.8

1. Differentiate between hard water and soft water.
2. Write balanced chemical equations to show how water hardness is removed using the following methods:
   (a). Boiling
   (b). Use of washing soda
3. Hard water is preferred for drinking while soft water is preferred for washing clothes. Explain.
4. The type of water hardness that can be removed by boiling is called____________.
5. Water which forms scum with soap is called ____________.

1.9 The carbon cycle

Discussion corner!

Study questions

1. Why do plants take in carbon dioxide and release oxygen during the day.
2. Why do animals take in oxygen and release carbon dioxide as a waste product.
3. When we burn wastes, what is produced?
4. What happens when you burn fossil fuels like coal, petroleum and diesel?

- Discuss the above study questions in groups.
- Summarise your findings in a report and present it to the rest of the class. Your report should include a cycle showing how carbon is recycled on Earth.

The facts

Some natural processes release carbon and its compounds into the atmosphere while other processes utilise this carbon and its compounds. **Carbon dioxide** is the compound through which carbon is recycled making it available to living things. The amount of carbon dioxide in the atmosphere remains fairly constant due to a delicate balance between processes that release the gas in to the atmosphere and those that remove the gas from the atmosphere. The circulation of carbon compounds in nature is called **carbon cycle**.

a) Processes that add carbon dioxide to the atmosphere

i. **Combustion:** This involves burning of petroleum, coal, wood, charcoal, wax or any other organic compound.

ii. **Respiration:** During this process, sugar (glucose) obtained from food taken by animals is converted to release energy needed by the body while carbon dioxide and water are given out as by products. Carbon dioxide is exhaled to the atmosphere.

iii. **Decay of plants and animals:** Any carbon (in plants and animals) is converted to carbon dioxide during the decaying process if sufficient oxygen supply is available.

iv. **Making of beer and wine:** Yeast converts sugar to ethanol and carbon dioxide. This process is called **fermentation**. It is summarised using the equation below.

$$\text{Sugar} \xrightarrow{\text{yeast}} \text{ethanol} + \text{carbon dioxide}$$

$$C_6H_{12}O_6\,(aq) \xrightarrow{\text{yeast}} 2C_2H_5OH\,(aq) + 2CO_2\,(g)$$

The carbon dioxide formed in this process finally finds its way into the atmosphere.

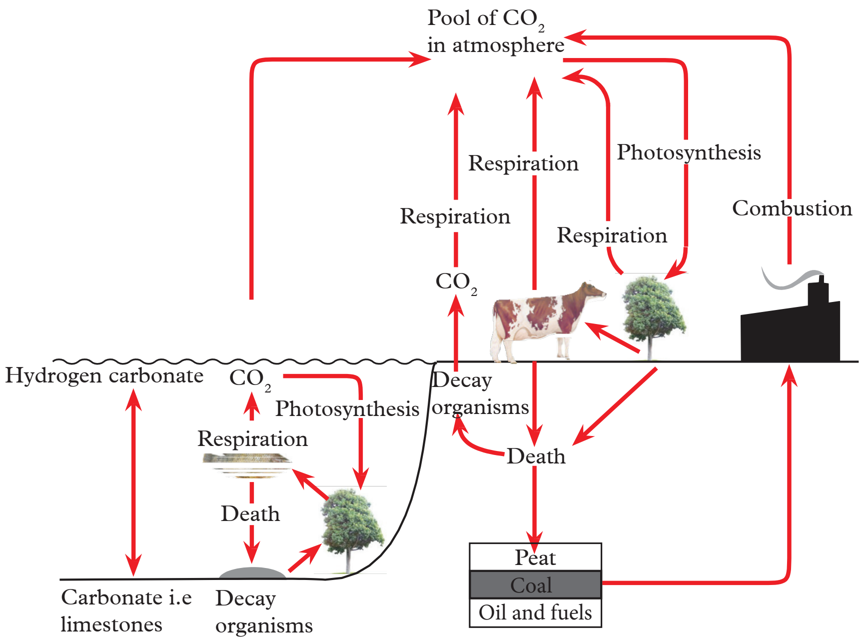


Fig1.31: The carbon cycle

b) Processes that remove carbon dioxide from the atmosphere

These include:

i. **Dissolution in water:** Carbon dioxide dissolves in rivers and lakes to form weak carbonic acid. The carbonic acid reacts with calcium carbonate and magnesium carbonate to form calcium hydrogen carbonate and magnesium hydrogen carbonate respectively.

Carbon dioxide + water ⟶ carbonic acid

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

Calcium carbonate + water + carbon dioxide ⟶ calcium hydrogen carbonate

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

Magnesium carbonate + water + carbon dioxide ⟶ magnesium hydrogen carbonate

$$MgCO_3(s) + H_2O\ (l) + CO_2\ (g) \longrightarrow Mg(HCO_3)_2(aq)$$

ii. **Photosynthesis:** In this process, green plants use carbon dioxide to make sugar using sunlight as a source of energy.

Carbon dioxide + water ⟶ sugar + oxygen

$$6CO_2\ (g) + 6\ H_2O\ (l) \longrightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

iii. **Reaction with calcium hydroxide:** Natural calcium hydroxide slowly reacts with carbon dioxide from the air to form calcium carbonate.

Calcium hydroxide + carbon dioxide ⟶ calcium carbonate + water

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

Self-evaluation Test 1.9

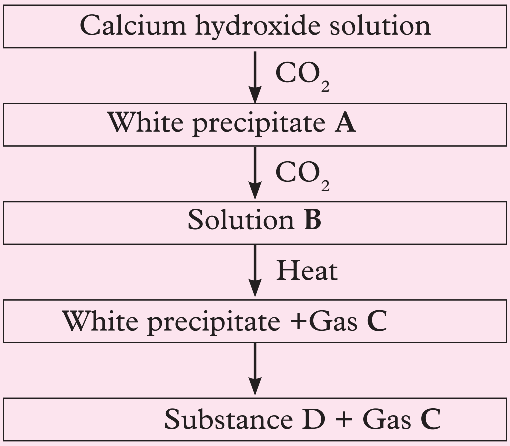
1. Carbon is a necessary evil. Explain.
2. If there is fire in the forest, there will be less___________ in the environment. (oxygen, carbon dioxide)
3. Why is it important to maintain a balance of carbon dioxide in the atmosphere?
4. Describe two processes that remove carbon dioxide from the atmosphere.
5. Carbon cycle is the movement of carbon to the Earth by the process of ___________ and ______________.

Unit summary 1

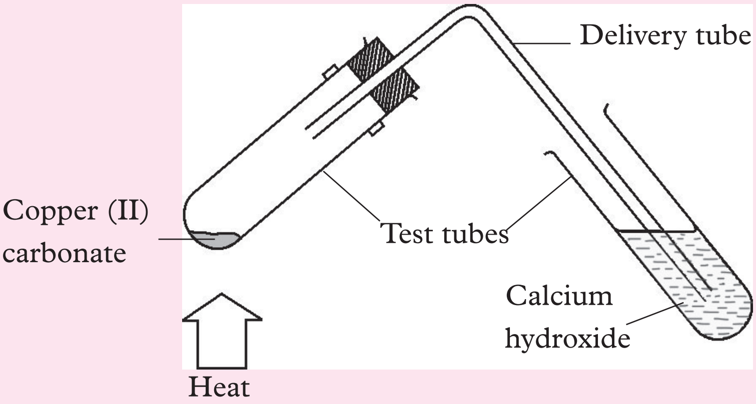
- Allotropy is the existence of an element in more than one form in the same physical state.
- Diamond and graphite are the two allotropes of carbon.
- The uses of diamond and graphite are linked to the type of bond joining atoms and layers together and the presence or absence of delocalized electrons.
- Carbon can exist as a free element or in compounds like oxides of carbon, carbonates and hydrogen carbonates.
- Carbon acts as a reducing agent by removing oxygen from compounds like iron oxide, concentrated nitric acid and concentrated sulphuric acid.
- In the laboratory, carbon dioxide is prepared by reacting a metal carbonate with any dilute mineral acid that forms a soluble salt.
- The reaction between dilute sulphuric acid and calcium carbonate cannot be used to prepare carbon dioxide as the reaction does not proceed to completion.
- Too much carbon dioxide in the atmosphere causes global warming.
- Carbon monoxide is a very poisonous and toxic gas.
- Water that contains dissolved calcium and magnesium ions is called hard water. Soft water lacks these ions.
- Soft water lathers easily with soap. Hard water does not.
- The whole process by which carbon dioxide circulates in the atmosphere is referred to as the carbon cycle.

Test your Competence 1

1. _______________ and ____________are two the crystalline allotropes of carbon while___________and _________are the non-crystalline forms of carbon.
2. The flow-chart below shows some processes involving some compounds of carbon. Use it to answer the questions that follow.



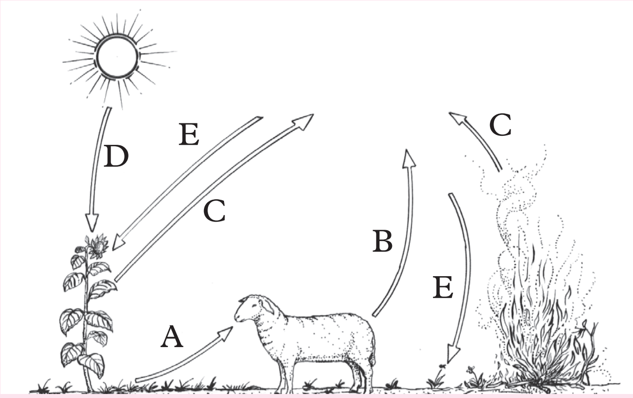
   (a) Name the white precipitate A, solution B and gas C.
   (b) Describe a test that can be used to identify carbon dioxide in the laboratory.
3. Akaliza set up an experiment as shown below.



   (a) What observation do you expect her to make in each test tube at the end of the experiment?
   (b) Write balanced equations that would lead to the observations expected in both test tubes.
   (c) Supposing she used sodium carbonate instead of copper carbonate, what observations would you expect her to make? Explain.

(d) Write a balanced chemical equation for the decomposition of ammonium carbonate by heat.

4. ___________ is a gas formed when carbon burns in enough oxygen supply while ___________ is formed when carbon burns in inadequate oxygen supply.
5. Zinc oxide is ___________ in colour when hot and ___________ when cold while lead oxide is ___________ in colour when hot and ___________ when cold.
6. Natural resources like fossil fuels are important and also harmful if not well utilised. Explain.
7. The following are forms of amorphous carbon. Which one is **not**?
   A. Lamp black
   B. Charcoal
   C. Soot
   D. Graphite
8. Hard water does not lather easily and hence wastes a lot of soap. Explain using chemical equations how you can convert hard water to soft water and use it in washing clothes.
9. The diagram below represents part of the carbon cycle.



Name processes A, B, C, D and E.

10. Explain how you can minimise emission of greenhouse gases in your locality.
11. Planting large numbers of trees results in ___________.
    A. A decrease in oxygen production.
    B. An increase in carbon dioxide production.
    C. A decrease in carbon dioxide built up in the atmosphere.
    D. An increase in oxygen use.

12. Carbon dioxide is a gas produced by ________________ and ________________ processes.

13. Identify and circle the raw materials required by plants during photosynthesis in the grid below.

S	U	N	L	I	G	H	T	B	D	V	T	T
K	D	F	C	E	S	B	D	C	A	E	F	C
C	A	R	B	O	N	D	I	O	X	I	D	E
L	D	C	S	Q	A	W	A	V	S	H	K	V
H	V	W	A	T	E	R	R	U	J	J	M	Q
G	N	Q	W	E	F	T	G	H	O	N	C	B
M	M	E	T	H	U	Y	F	E	R	M	Y	L

14. Identify and rewrite the correct word using the jumbled words given in the table below. Use the clues provided.

Scramble word	Correct word	Clues
cbnaor	________________	This element is found in carbohydrates, fats and proteins.
Etsssynitohhop	________________	Process where plants take in carbon dioxide from the air and convert it to glucose and oxygen.
Enroaitispr	________________	Process by which plants and animals convert glucose in the presence of oxygen into carbon dioxide and water.
Riaooxiddcenb	________________	Carbon containing gas taken in by plants during photosynthesis and released during respiration.

UNIT 2

Nitrogen and its inorganic compounds

Key unit competency

To be able to relate the properties of nitrogen and its compounds to their uses, describe how some compounds of nitrogen are prepared and discuss the related environmental issues.

Learning objectives

By the end of this unit, I should be able to:

- Describe the physical and chemical properties of nitrogen and its uses.
- Describe the impact of nitrogen compounds on the environment.
- Describe the industrial preparation of nitric acid and state its uses and those of some nitrates.
- Prepare and collect ammonia gas and state its uses.
- Interpret the effect of ammonia solution on different cations.
- Develop observation skills during the experiments on addition of ammonia solution on different cations.
- Develop self-confidence, a culture of working as a team and respect procedures in all experiments.
- Appreciate the dangers posed by nitrogen compounds to the environment.
- Carry out research on protection of the environment and hence protect natural resources.

Mind teaser

Like carbon, nitrogen is also important in our lives. Study these pictures carefully.

A **B** **C**

Fig 2.1

How is nitrogen used in each picture above? What is the significance of that? Which other uses of nitrogen do you know?

Introduction

In senior one, you studied nitrogen as one of the components of air. What is the percentage composition of nitrogen in air? What compounds do you know that contain nitrogen?

These compounds find application not only in Chemistry but also in other disciplines like Agriculture, Biology and Physics. For example, ammonia, nitric acid and some fertilisers are examples of compounds that contain nitrogen. It is therefore important to study how we can prepare nitrogen and use it for various purposes.

Fig 2.2: A bag of nitrogenous fertiliser

2.1 Industrial isolation of nitrogen

Activity 2.1

1. Carry out a research using textbooks or from the internet about industrial isolation of nitrogen from other gaseous components. Use the sites http://www.fuseschool.org and http://www.youtube.com to do the research.

Study questions

1. Is air a mixture or a compound?
2. Components of a mixture can be separated by physical means. Do you agree?
3. What is the percentage composition of nitrogen in air?

Discussion corner!

Discuss the above study questions in your groups. Conclude whether it is possible to separate the components of air by physical means.

The facts

Air is a mixture of gases. The components of air include nitrogen, oxygen, water vapour, carbon dioxide, noble gases and dust particles. These components can be separated by physical means. As earlier mentioned, nitrogen forms about 78% of air by volume.

Nitrogen can be isolated from air by fractional distillation of liquid air. This process involves eliminating other components of air hence remaining with nitrogen. It is a four step process as discussed below:

(a) Purification of air

This is the initial step. Air is purified to remove water vapour, dust and carbon dioxide. This is done by first passing air through filters to remove dust particles.

The dust-free air is then passed through driers to remove water vapour and finally it is passed through concentrated solution of sodium hydroxide. Sodium hydroxide absorbs carbon dioxide from the air. Note: It is important to remove water vapour and carbon dioxide before liquefaction to prevent the two from solidifying and blocking the pipes. Carbon dioxide and water vapour may also be removed by cooling the air to -80°C.

(b) Compression of air

This is the second stage whereby the remaining gas (mainly mixture of nitrogen, oxygen and noble gases) are compressed at very high pressures (200 atmospheres). This produces a lot of heat and so the compressed gases must immediately be cooled by passing them through cooling coils in a tank. The cooled compressed air is then allowed to expand by passing it through a narrow valve that then expands it rapidly. This makes its temperature to drop sharply. It is then returned to the compressor and the process repeated several times. After this repeated compression and expansion of air, cooling is done which liquefies air into a pale-blue liquid at -200°C. Neon and helium do not liquefy at this temperature and are hence removed at this stage.

(c) Fractional distillation

This is the third stage. The liquid air obtained is a mixture of colourless nitrogen (boiling point = -196°C), pale-blue oxygen (boiling point = -183°C) and argon (boiling point = -186°C). Air is thus passed through a fractionating column. The temperature of liquid air in the fractionating column is slowly raised. Nitrogen boils at -196°C and so distills first and is collected at the top of the fractionating column. Argon with a boiling point of -186°C and oxygen with a boiling point of -183 °C distills later and are collected from the lower part of the column in that order.

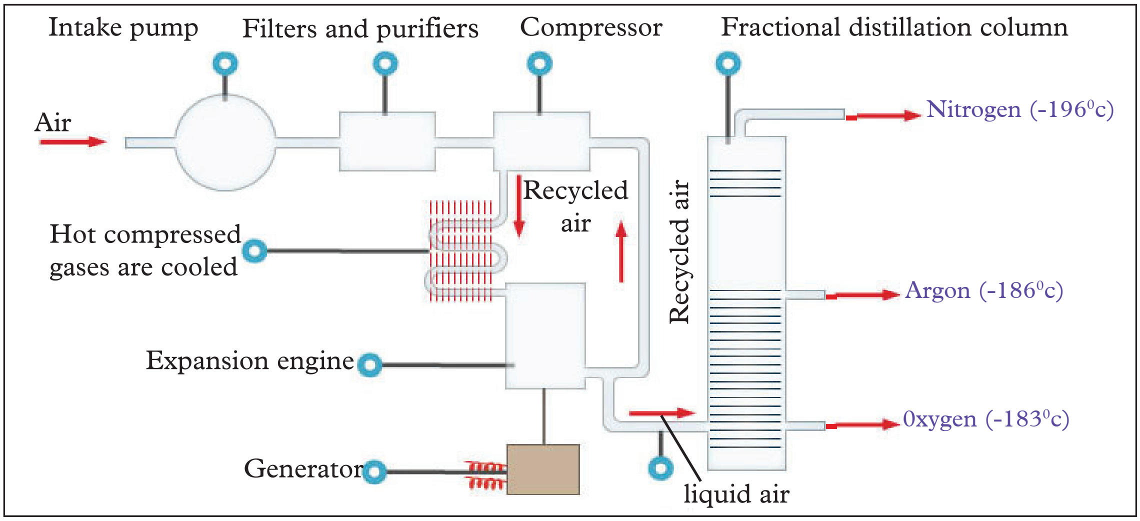


Fig 2.3: Fractional distillation of liquid air

(d) Storage of nitrogen

The nitrogen obtained is liquefied and stored under pressure.

Self-evaluation Test 2.1

1. Why is it possible to isolate nitrogen from air?
2. What is the importance of the following during isolation of nitrogen from air:
   (a) Passing air through filters and concentrated sodium hydroxide solution?
   (b) Compression and expansion of air?
   (c) Fractional distillation of liquid air?
3. In your study groups, discuss the industrial isolation of nitrogen gas from air and make a presentation to the rest of class.
4. Air contains mainly _____________ gas with about ___________%. The main components of air are mainly separated by _______________process.
5. Liquid air has three components X, Y and Z whose boiling points are -186°C, -183°C and -196°C respectively. When the liquid air is fed into a tall fractional distillation column and near its bottom warmed up slowly:
   (a) Which component will be collected from near the bottom of the fractional distillation column? Why?
   (b) Which component will be collected from the top of the fractional distillation column? Why?
   (c) Which component will be collected from the middle part of the fractional distillation column? Why?
   (d) What could the components X, Y and Z be in real situation?

2.2 Laboratory preparation of nitrogen

Activity 2.2

Apparatus and reagents

Round-bottomed flask or boiling tube, Bunsen burner/any source of heat, trough, gas jars, sodium nitrite, ammonium chloride.

Procedure

1. Make a solution of a mixture of sodium nitrite and ammonium chloride in water.
2. Put the solution in a round-bottomed flask or boiling tube.
3. Arrange the apparatus as shown in Fig. 2.4 below.

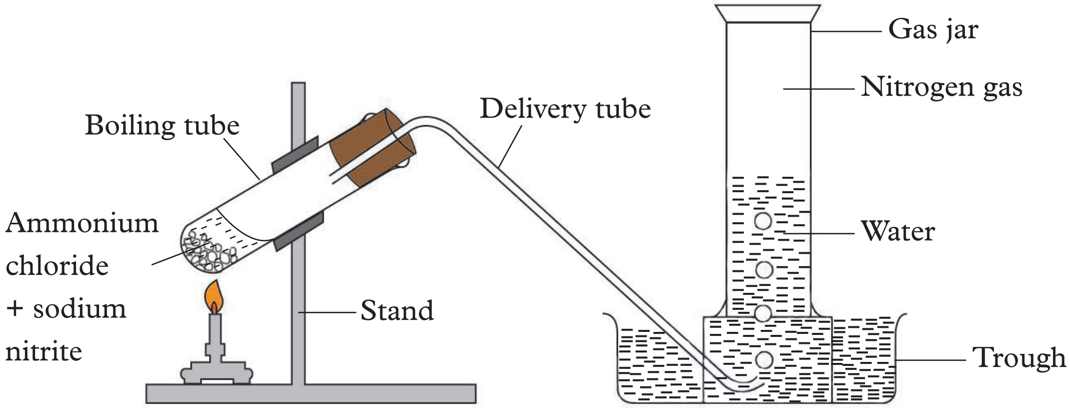


Fig. 2.4: Laboratory preparation of nitrogen

4. Warm the flask slightly and then remove the heat.
5. Record your observations in the flask and gas jar.

Study questions

1. Why was warming done only for a short time?
2. Write balanced chemical equations for the reactions that take place.

Discussion corner!

Discuss your observations in groups and present your findings is class.

I have discovered that…

When the flask is warmed, bubbles are produced and this continues even after warming has been stopped. A colourless gas collects in the upper part of the gas jar.

The facts

Sodium nitrite reacts with ammonium chloride to form ammonium nitrite and sodium chloride.

Sodium nitrite + ammonium chloride ⟶ sodium chloride + ammonium nitrite

$$NaNO_2\,(aq) + NH_4Cl\,(aq) \longrightarrow NaCl\,(aq) + NH_4NO_2\,(aq)$$

The ammonium nitrite is then decomposed to produce nitrogen gas.

Ammonium nitrite ⟶ nitrogen + water

$$NH_4NO_2\,(aq) \longrightarrow N_2(g) + 2H_2O(l)$$

The overall equation for the reaction:

Sodium nitrite + ammonium chloride ⟶ sodium chloride + water + nitrogen

$$NaNO_2\,(aq) + NH_4Cl\,(aq) \longrightarrow NaCl(aq) + 2H_2O(l) + N_2(g)$$

Note: If required dry, nitrogen gas can be passed through concentrated sulphuric acid. The gas is then collected in a syringe or by upward delivery.

Nitrogen can also be prepared by direct heating of ammonium nitrite. However, when ammonium nitrite is in small quantities, the reaction can be explosive.

Properties and uses of nitrogen

Activity 2.3

Apparatus and reagents

Wooden splint, blue and red litmus papers, aqueous calcium hydroxide, magnesium ribbon, sulphur powder, six gas jars full of dry nitrogen, a pair of tongs, deflagrating spoon.

Procedure

1. Copy Table 2.1 in your notebook.
2. Carry out the procedures 1 -7 and use your observations and conclusions to fill the table.

Table 2.1: Investigating the properties of nitrogen

Procedure	Observations	Conclusion
1. Observe the colour and smell of nitrogen in the first gas jar.		
2. Put a glowing splint into the first gas jar.		
3. Insert a burning splint into the second gas jar.		
4. Insert wet, blue and red litmus papers into the second gas jar.		
5. Add a few drops of aqueous calcium hydroxide into the third gas jar. Shake gently.		
6. Lower a burning magnesium ribbon into the fourth jar. Add water to the products. Smell the gas produced.		
7. Lower burning sulphur on a deflagrating spoon into the fifth gas jar.		

Study questions

Write chemical equations for the reactions that take place in procedure number 6.

Discussion corner!

- In your study groups, discuss the observations made in each of the investigations.
- Present your results to the rest of the class.

I have discovered that…

Nitrogen shows no observable reactions for most of the tests carried out.

The facts

The following are some properties of nitrogen gas.

Physical properties of nitrogen

- Nitrogen is colourless and does not smell (odourless).
- It is almost insoluble in water and hence can be collected over water.
- It is lighter than air and therefore is collected by upward delivery method.

Chemical properties of nitrogen

(i). Combustion

Nitrogen neither burns nor supports combustion. That is why a glowing or burning splint goes off when introduced in nitrogen.

(ii). Effect on litmus papers

Nitrogen is neutral. Litmus papers therefore show no colour change in nitrogen gas.

(iii). Effect on aqueous calcium hydroxide

Nitrogen has no effect on calcium hydroxide.

(iv). Reaction with magnesium and sulphur

Nitrogen reacts with burning magnesium to form magnesium nitride.

$$\text{Magnesium} + \text{nitrogen} \longrightarrow \text{magnesium nitride}$$

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

Magnesium nitride reacts with water to form magnesium hydroxide solution and ammonia gas.

$$\text{Magnesium nitride} + \text{water} \longrightarrow \text{magnesium hydroxide} + \text{ammonia}$$

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

However, when burning sulphur is introduced into a gas jar of nitrogen, no reaction occurs. This is because the heat produced from burning sulphur is not sufficient to break the triple covalent bonds ($N\equiv N$) of the nitrogen molecules for the reaction to occur.

Nitrogen is inert at room temperature. It does not react with other elements at room temperature. However, the high temperatures produced by burning metals like magnesium break down the triple covalent bonds of the nitrogen molecules to form single atoms of nitrogen which is now able to react with magnesium.

$$N\equiv N \longrightarrow 2N \text{ (very high temperature needed).}$$

Uses of nitrogen

1. It is used in the Haber process to manufacture ammonia gas.
2. It is used to freeze substances (coolant)due to its low boiling point of -196°C. It can be used to store body tissues that are required to last for a long period of time. For instance:
    - Tissues used in hospitals.
    - Bulls' semen for artificial insemination.
    - It can also be used to mend leaking pipes. When liquid nitrogen is poured onto the pipe, it freezes the liquid inside while repair is done.
3. It is used in food processing. When food is being packed, nitrogen is used to keep off oxygen and the food stays fresh for a long time because no oxidation can take place. Most bacteria cannot survive in nitrogen.
4. Because of its inert nature at low temperature, it is pumped into the ships' tanks that transport crude oil to remove traces of oxygen. This prevents dangerous explosions that could occur from crude oil vapours.

Self-evaluation Test 2.2

1. Nitrogen gas does not support combustion but a magnesium ribbon continues to burn in the presence of nitrogen. Discuss.
2. Liquid nitrogen boils at -196°C. What is the importance of this?
3. Give three physical properties of nitrogen gas that differentiate it from other gases.
4. In order to attain electronic configuration of noble gas, nitrogen needs to ______.
    - A. Lose one electron
    - B. Gain two electrons
    - C. Loose three electrons
    - D. Gain three electrons

2.3 Preparation, properties and uses of nitrogen compounds

Compounds to be studied under this section include:

- Nitrogen dioxide
- Ammonia
- Nitric acid

a) Nitrogen dioxide

Activity 2.4

Apparatus and reagents

Thistle funnel or dropping funnel, flat-bottomed flask , delivery tube, 3 gas jars , cardboard cover , copper turnings, concentrated nitric acid.

Caution: Nitrogen dioxide is poisonous. This experiment should therefore be performed in a fume cupboard or in an open space.

Procedure

1. Put copper turnings into a flat-bottomed flask.
2. Arrange the apparatus as shown in Fig. 2.5.

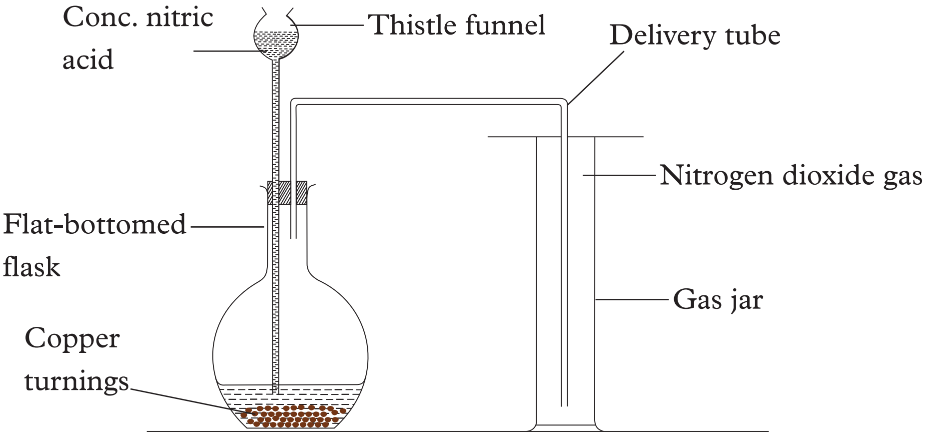


Fig 2.5: Laboratory preparation of nitrogen dioxide

3. Add concentrated nitric acid to the copper turnings in the flask.
4. Make your observations in the flask and the gas jar and record in your notebook.

Study questions

1. What observations did you make in the flask and the gas jar?
2. Write a balanced chemical equation for the reaction that takes place.

Discussion corner!

- In your study groups, discuss the results of the experiment and the study questions above.
- Share your findings with the rest of the class.

I have discovered that…

When concentrated nitric acid is added to copper turnings, effervescence occurs immediately producing red-brown fumes. A green solution is also formed in the flask.

The facts

Copper reacts with concentrated nitric acid to form a green solution of copper (II) nitrate, water and a red-brown gas (nitrogen dioxide).

Copper + nitric acid ⟶ copper (II) nitrate + water + nitrogen dioxide

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

More pure nitrogen dioxide can be prepared by heating a nitrate of a moderately reactive heavy metal. Lead (II) nitrate is preferred because unlike most nitrates, they do not contain water of crystallisation which would otherwise interfere with the preparation process since nitrogen dioxide is highly soluble in water.

$$2Pb(NO_3)_2(s) \xrightarrow{\text{Heat}} 2PbO(s) + 4NO_2\ (g) + O_2\ (g)$$

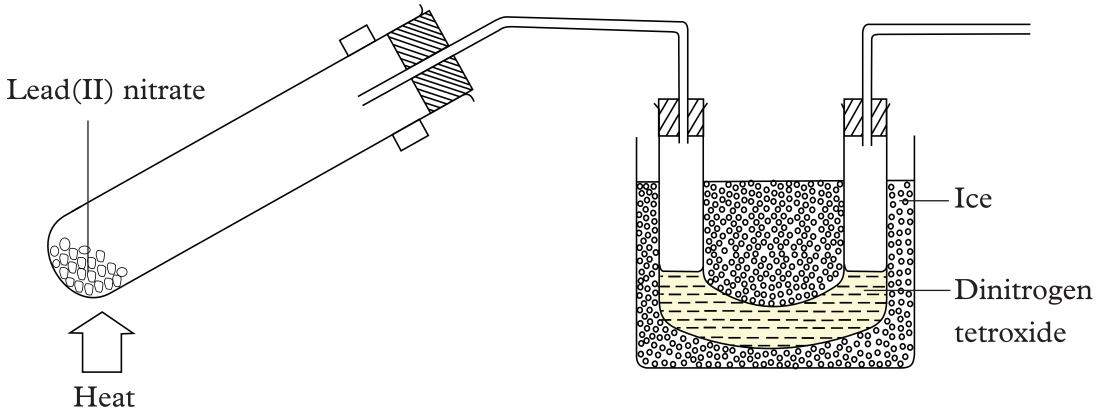


Fig 2.6: Preparation of nitrogen dioxide by heating lead (II) nitrate

In this method, the lead (II) nitrate decomposes on heating to give a mixture of nitrogen dioxide and oxygen. The nitrogen dioxide is cooled by freezing in a mixture ice and salt forming a pale-yellow liquid called **dinitrogen tetroxide** (N_2O_4). Oxygen passes through as a gas and is collected over water. This way, the two gases are separated. Dinitrogen tetraoxide easily dissociates into nitrogen dioxide when the temperature is raised as per the equation below.

$$\underset{\text{(Pale - yellow)}}{N_2O_4(l)} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \underset{\text{(Red - brown)}}{2NO_2(g)}$$

Properties of nitrogen dioxide

a) *Physical properties of nitrogen dioxide*

- It is a red-brown gas
- It has an irritating pungent smell.
- It is soluble in water forming an acidic solution
  $2NO_2\ (g) + H_2O(l) \longrightarrow HNO_3(aq) + HNO_2(aq)$
- It is one and half times denser than air.
- It is easily liquefied on cooling (boiling point is 21°C).

b) *Chemical properties of nitrogen dioxide*

Activity 2.5

Apparatus and reagents

Litmus papers (blue and red), magnesium ribbon, phosphorus, sulphur, copper turnings, a pair of tongs, wooden splints.

Procedure

1. Copy Table 2.2 in your notebook.
2. Prepare nitrogen gas as shown in Activity 2.4 then carry out the following investigations 1-7. Record your observations and conclusions in the blank spaces provided.

Table 2.2: Investigating the chemical properties of nitrogen dioxide

Procedure	Observation	Conclusion
1. Put a glowing splint into the first gas jar.		
2. Put a burning splint into the second gas jar.		
3. Put wet litmus papers (blue and red) into the third gas jar.		
4. Put a burning magnesium ribbon into the fourth gas jar.		
5. Put vigorously burning phosphorus and sulphur into the fifth and sixth jars respectively.		
6. Put red-hot copper into the seventh gas jar.		

7. Heat the gas in the corked test-tube and then place a glowing splint into the tube.		

Study questions

1. Account for the observations in procedures 4 and 5.
2. Explain the observations made in procedure 6.

Discussion corner!

- In your study groups, discuss the observations made in each procedure and the study questions above.
- Draw appropriate conclusions and present your findings to the rest of the class.

I have discovered that…

Both glowing and burning splints goes off on exposure to nitrogen dioxide gas. Burning magnesium, phosphorus and sulphur continue to burn in the presence of nitrogen dioxide.

The facts

The following are the chemical properties of nitrogen dioxide.

i. Combustion

Nitrogen dioxide neither burns nor supports combustion. That is why it puts off both the glowing and the burning splints.

ii. Reaction with water

Nitrogen dioxide reacts with water to form nitric acid and nitrous acid.

Nitrogen dioxide + water ⟶ nitric acid + nitrous acid

$$2NO_2\,(g) + H_2O(l) \longrightarrow HNO_3(aq) + HNO_2(aq)$$

The gas is therefore acidic and so the solution formed turns wet blue litmus paper red.

iii. Effects on burning elements

Burning magnesium continues to burn in nitrogen dioxide forming magnesium oxide and nitrogen gas.

Magnesium + nitrogen dioxide ⟶ magnesium oxide + nitrogen gas

$$4Mg(s) + 2NO_2(g) \longrightarrow 4MgO(s) + N_2(g)$$

The heat from burning magnesium decomposes this gas into its constituent elements nitrogen and oxygen.

The oxygen produced supports the continued burning of magnesium.

The same type of reaction takes place if vigorously burning phosphorus and sulphur are introduced into gas jar containing nitrogen dioxide.

Phosphorus + nitrogen dioxide ⟶ phosphorus pentoxide + nitrogen gas
$8P(s) + 10NO_2(g) \longrightarrow 4P_2O_5(s) + 5N_2(g)$

Sulphur + nitrogen dioxide ⟶ sulphur dioxide + nitrogen gas
$2S(s) + 2NO_2(g) \longrightarrow 2SO_2(g) + N_2(g)$

Nitrogen dioxide oxidises red-hot copper to copper (II) oxide while itself is reduced to nitrogen gas.

Copper + nitrogen dioxide ⟶ copper (II) oxide + nitrogen gas
$4Cu(s) + 2NO_2(g) \longrightarrow 4CuO(s) + N_2(g)$

Note: Heated nitrogen dioxide easily reacts with burning substances because the heat decomposes it to nitrogen monoxide and oxygen.

Nitrogen dioxide $\xrightarrow{\text{heat}}$ nitrogen monoxide + oxygen
$2NO_2(g) \xrightarrow{\text{heat}} 2NO(g) + O_2(g)$

Nitrogen monoxide is further decomposed to nitrogen and oxygen.

Nitrogen monoxide $\xrightarrow{\text{heat}}$ nitrogen + oxygen
$2NO(g) \xrightarrow{\text{heat}} N_2(g) + O_2(g)$

Finally the oxygen combines with element in given reaction to form their respective oxides.

iv. Reaction with sodium hydroxide solution

Nitrogen dioxide reacts with sodium hydroxide solution to form two salts that is sodium nitrate ($NaNO_3$) and sodium nitrite ($NaNO_2$). The two salts are formed because nitrogen dioxide when dissolved in water forms two acids; nitric acid (HNO_3) and nitrous acid (HNO_2).

Nitrogen dioxide + water ⟶ nitric acid + nitrous acid
$2NO_2(g) + H_2O(l) \longrightarrow HNO_3(aq) + HNO_2(aq)$

These acids though present in the same solution, separately react with sodium hydroxide solution to form the two salts.

Sodium hydroxide + nitric acid ⟶ sodium nitrate + water
$NaOH(aq) + HNO_3(aq) \longrightarrow NaNO_3(aq) + H_2O(l)$
Sodium hydroxide + nitrous acid ⟶ sodium nitrite + water
$NaOH(aq) + HNO_2(aq) \longrightarrow NaNO_2(aq) + H_2O(l)$

The overall equation for the reaction between nitrogen dioxide and sodium hydroxide is therefore as shown below.

Sodium hydroxide + nitrogen dioxide ⟶ sodium nitrate + sodium nitrite + water
$2NaOH(aq) + 2NO_2(g) \longrightarrow NaNO_3(aq) + NaNO_2(aq) + H_2O(l)$

Uses of nitrogen dioxide

1. It is used as an intermediary in the manufacture of nitric acid by Ostwald Process.
2. It is an intermediary in the formation of very dilute nitric acid during

lightning. This acid is washed into the soil, thereby increasing the amounts of nitrates in the soil.

Health check!

Nitrogen dioxide and the environment

The exposure to high levels of nitrogen dioxide (NO_2) may result in respiratory disorders. Nitrogen dioxide in flames the lining of the lungs which could lead to reduced lung immunity. This can cause further problems such as wheezing, coughing, colds, flu and bronchitis.

Increased levels of nitrogen dioxide can have significant health risk on persons with asthmatic conditions because it can lead to more frequent and more intense attacks. Children with asthma and older people with heart diseases are most at risk.

Visit: www.greenfacts.org for more information

Self-evaluation Test 2.3

1. Nitrogen dioxide dissolves in water to form two acids. Write the chemical equation for the reaction that takes place.
2. Nitrogen dioxide neither burns nor supports combustion. However, burning magnesium continues to burn in the presence of the gas. Explain.
3. Give one use of nitrogen dioxide.

b) Ammonia

Laboratory preparation of ammonia

Activity 2.6

Apparatus and reagents

Round-bottomed flask, glass tubing, spatula, drying tower, several gas jars, mortar and pestle, cardboard cover, calcium hydroxide, ammonium chloride, calcium oxide, litmus papers (blue and red).

Procedure

1. Place one spatula full of calcium hydroxide and one spatula full of ammonium chloride in a mortar.
2. Grind the mixture well using the pestle then put it in a round-bottomed flask.
3. Arrange the apparatus as shown in fig 2.7.

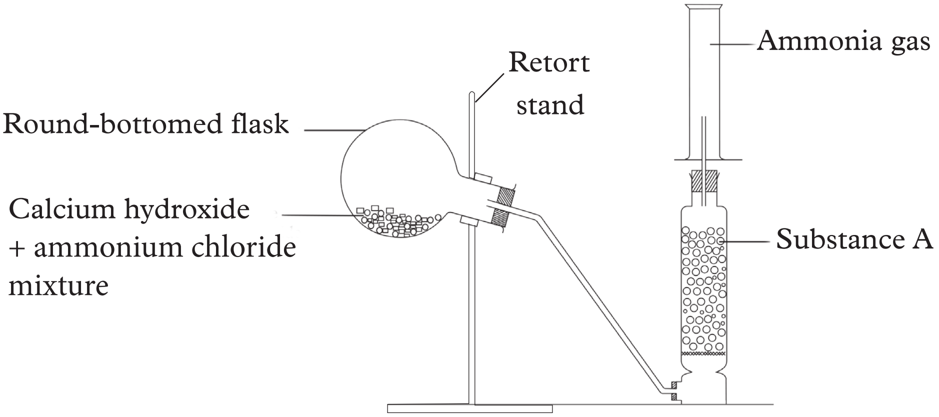


Fig 2.7: Laboratory set up for the preparation of ammonia

Caution: The flask should be set in a sloping position as shown in the diagram before the experiment starts.

4. Heat the mixture.
    - Note down what you observe on the cooler parts of the flask and the delivery tube.
5. Check whether the gas jar is full, by putting a moist red litmus paper at the mouth of the jar. It should turn blue if the gas jar is full.
6. Collect five jars of this gas for the next activity 2.7.

Study questions

1. Why should the flask be in a sloping position?
2. Identify the chemical labelled substance A in the Fig 2.7 and state its function.
3. Name the method used to collect ammonia gas and suggest why it is preferred.

Discussion corner!

- In groups, discuss the above method of preparing ammonia gas and the study questions above.
- Write a group report on the experiment and present it to the teacher for evaluation.

I have discovered that…

When the mixture is heated, water droplets are seen on the cooler parts of the flask and delivery tube. When a damp red litmus paper is placed at the mouth of the gas jar, it turns blue.

The facts

Calcium hydroxide reacts with ammonium chloride to form calcium chloride, water and ammonia gas. The equation for the reaction is as follows:

Calcium hydroxide + ammonium chloride ⟶ calcium chloride + water + ammonia

$Ca(OH)_2(s) + 2NH_4Cl(s) \longrightarrow CaCl_2(s) + 2H_2O(l) + 2NH_3(g)$

Substance A is calcium oxide. It is used to dry ammonia gas. Other drying agents like concentrated sulphuric acid and anhydrous calcium chloride would react with ammonia and hence cannot be used to dry the gas.

Ammonia + sulphuric acid ⟶ ammonium sulphate

$2NH_3\ (g) + H_2SO_4(l) \longrightarrow (NH_4)_2SO_4(aq)$

Ammonia + calcium chloride ⟶ calcium ammonium chloride

$4NH_3\ (g) + CaCl_2\ (aq) \longrightarrow CaCl_2.4NH_3(s)$

(Solid complex)

The flask should be in a slanting position to prevent drops of water vapour that condenses on the cooler parts of the flask from flowing back into the flask which would make the flask to crack.

Ammonia is less dense than air and is hence collected by upward delivery (downward displacement of air).

Properties of ammonia gas

Activity 2.7

Apparatus and reagents

Four gas jars of ammonia, wooden splints, trough, glass rod, concentrated hydrochloric acid.

Procedure

1. Copy table 2.3 in your notebook.
2. Carry out procedures 1-5 and use your results to fill the table.

Table 2.3: Investigating the properties of ammonia gas

Procedure	Observation	Conclusion
1. Waft the gas in the first jar towards your nose. Note the smell and colour of ammonia.		
2. Dip wet red and blue litmus papers into the second gas jar.		
3. Put a glowing splint and then a burning splint into the third gas jar.		
4. Invert the fourth gas jar in a trough of water. Note the level of the water inside the gas jar after 2 to 3 minutes.		

5. Dip a glass rod in concentrated hydrochloric acid and bring the dipped end to the mouth of the fifth gas jar.		

Study questions

1. Describe the smell of ammonia gas.
2. Is ammonia an acidic or alkaline gas?

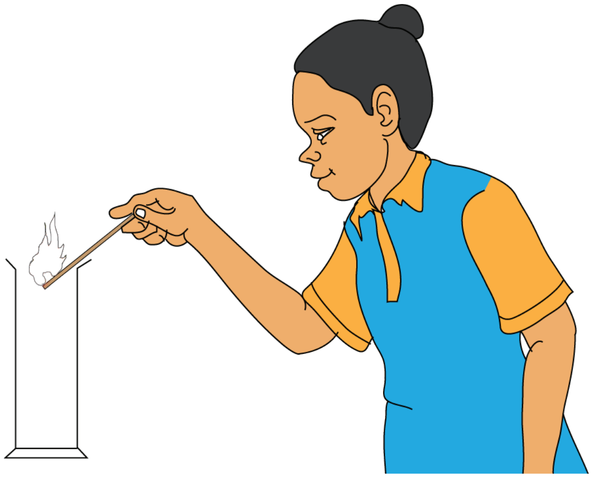


Fig 2.8: Testing ammonia gas using a burning splint

Discussion corner!

- In groups, discuss the study questions and properties of ammonia based on your observations in this activity.
- Write a report and present it to the teacher for evaluation.

I have discovered that…

Ammonia is a colourless gas with a pungent choking smell. It reacts with hydrogen chloride forming white fumes. It is very soluble in water ($1 cm^3$ of water dissolves about $800 cm^3$ of the gas).

The facts

Physical properties of ammonia

1. Ammonia is a colourless gas.
2. It has a pungent choking smell.
3. It is very soluble in water. Water rises rapidly in a round-bottomed flask full of ammonia gas and almost fills it immediately if inverted in a trough containing water. This can be demonstrated in a **fountain experiment** as shown in the set up that follows.

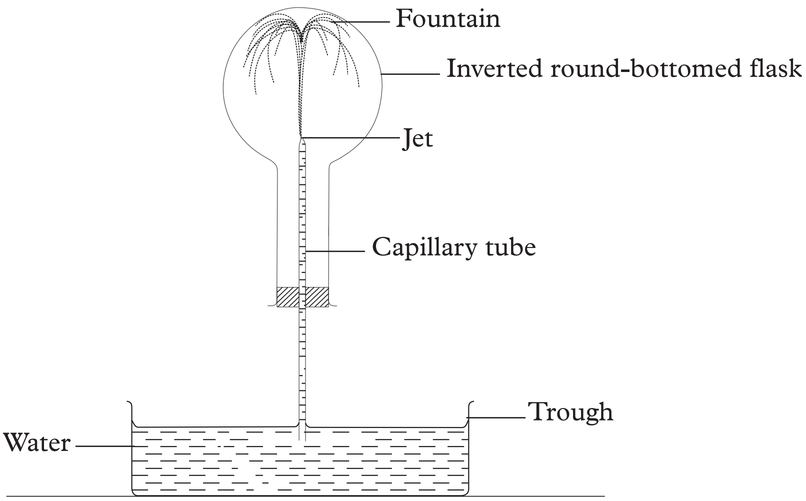


Fig 2.9: The fountain experiment

In the set up above, when water enters the flask, it dissolves so much ammonia that there is a partial vacuum in the flask. Water is then rapidly forced up the tube and enters the flask as a fountain.

Further activity

Carry out an experiment to demonstrate the fountain experiment. Arrange your apparatus as shown in the figure above.

Chemical properties of ammonia gas

1. Combustion

Ammonia does not burn in air neither does it support combustion.

2. Effect on litmus paper

Ammonia turns moist red litmus paper blue. It is the common known alkaline gas.

3. Reaction with hydrogen chloride

It reacts with hydrogen chloride gas to form dense white fumes of ammonium chloride.

$$NH_3\,(g) + HCl\,(g) \longrightarrow \underset{\text{(dense white fumes)}}{NH_4Cl(s)}$$

Note: This is the confirmatory test for ammonia.

4. Reaction of ammonia with oxygen

Ammonia reacts differently with air or oxygen. The reaction is even more different when a catalyst is introduced.

Activity 2.8

Apparatus and reagents

Open glass tube, delivery tube, glass beads or cotton wool, sources of oxygen and ammonia, wooden splint, cork/rubber bung.

Procedure

1. Arrange the apparatus as shown in the figure below.

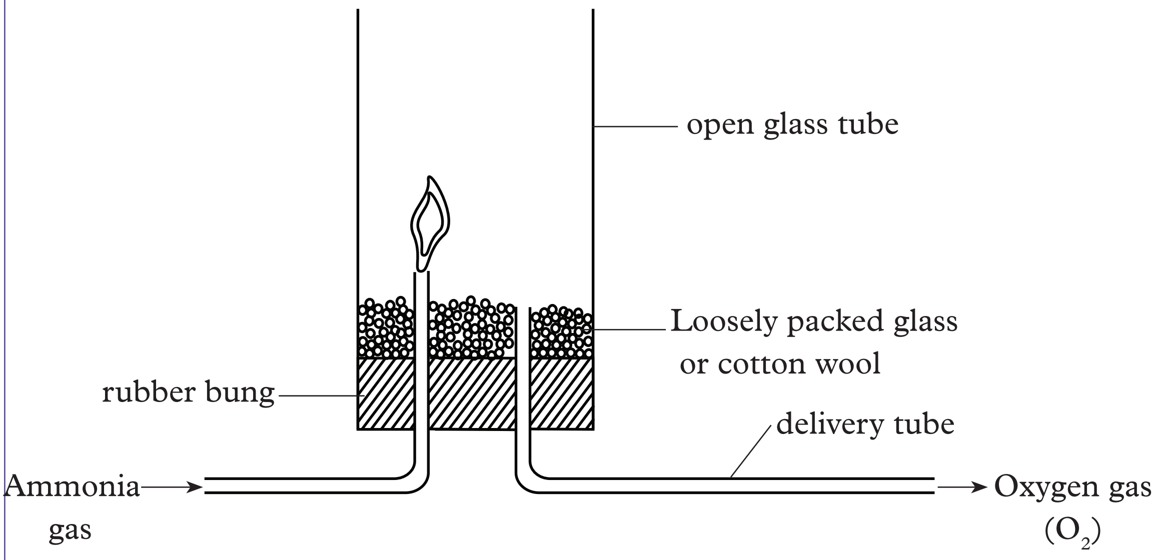


Fig 2.10: Burning ammonia in oxygen

2. Ignite ammonia gas at the jet delivering ammonia before allowing oxygen in as shown above.
3. Now adjust the rate of flow of both gases until a flame can be lit at the jet of ammonia tube.
4. Record your observations in your notebook.

Study questions

1. What is the use of the glass beads or cotton wool in this experiment?
2. What do you conclude from this experiment?

Discussion corner!

Discuss the experiment results in your groups and answer the study questions above. Present your findings to the rest of the class.

I have discovered that…

Ammonia burns in presence of oxygen with a yellow-brown flame.

The facts

Ammonia burns in oxygen to form nitrogen gas and water.

Ammonia + oxygen ⟶ nitrogen + water

$4NH_3 (g) + 3O_2 (g) \longrightarrow 2N_2 (g) + 6H_2O (l)$

The loosely packed glass beads or cotton wool helps to spread out the oxygen.

Note: Ammonia gas neither burns nor supports burning. It can only burn in oxygen or air enriched with oxygen.

Catalytic oxidation of ammonia

Activity 2.9

Apparatus and reagents

Platinum coil, delivery tube, beaker, concentrated ammonia solution, source of oxygen, litmus papers (blue and red).

Caution: The reaction can sometimes be explosive.

Procedure

1. Bubble oxygen through the concentrated ammonia solution for a few minutes.
2. Lower a red-hot platinum coil into the mixture of oxygen and ammonia solution.

Note: Concentrated ammonia solution produces ammonia gas.

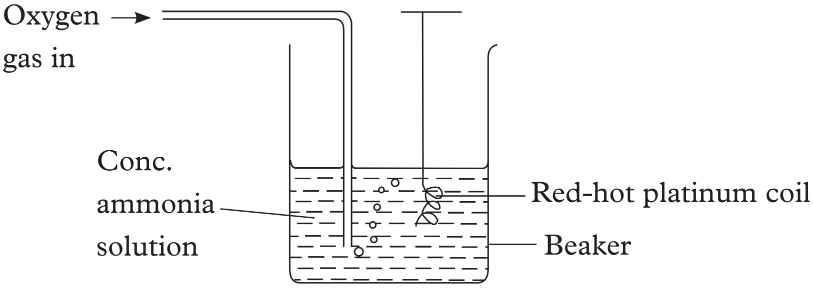


Fig 2.11: Catalytic oxidation of ammonia

3. Record your observations in your notebook.
4. Test the fumes produced using moist blue and red litmus papers.
5. Dip blue and red litmus papers in the solution after the reaction.

Study questions

1. What is the colour of the fumes that are seen coming out of the solution?
2. What is the effect of these fumes on moist blue litmus paper?
3. What is the effect of the resulting solution on blue litmus paper?
4. What conclusions can you make from your observations?

Discussion corner!

- Discuss the results of the experiment in your groups and answer the study questions above.
- Present your discussion report to the rest of the class.

I have discovered that…

Red-brown fumes are produced when the reaction takes place. The fumes turn blue litmus paper red but have no effect on red litmus paper.

The facts

Ammonia reacts with oxygen in the presence of platinum (Pt) - Rhodium (Rh) catalyst to form nitrogen monoxide and water as shown in the equation below.

$$\text{Ammonia} + \text{oxygen} \xrightarrow[900°\text{C}]{\text{Pt-Rh}} \text{nitrogen monoxide} + \text{water}$$

$$4NH_3(g) + 5O_2(g) \xrightarrow[900°\text{C}]{\text{Pt-Rh}} 4NO(g) + 6H_2O(g)$$

The nitrogen monoxide is further oxidised to form nitrogen dioxide gas characterised by red-brown fumes. This is an acidic gas.

$$\text{Nitrogen monoxide} + \text{oxygen} \longrightarrow \text{nitrogen dioxide}$$

$$2NO\ (g) + O_2\ (g) \longrightarrow \underset{\text{(Red-brown fumes)}}{2NO_2\ (g)}$$

When nitrogen dioxide is dissolved in water, nitrous acid and nitric acid are formed.

$$\text{Nitrogen (IV) oxide} + \text{water} \longrightarrow \text{nitric acid} + \text{nitrous acid}$$

$$2NO_2\ (g) + H_2O\ (l) \longrightarrow HNO_3\ (aq) + HNO_2\ (aq)$$

Nitrous acid is further oxidised to form nitric acid. Nitric acid is the one that turns blue litmus paper red.

$$\text{Nitrous acid} + \text{oxygen} \longrightarrow \text{nitric acid}$$

$$2HNO_2\ (aq) + O_2\ (g) \longrightarrow 2HNO_3\ (aq)$$

It is therefore possible to convert ammonia to nitric acid using platinum as a catalyst.

Manufacture of ammonia by Haber process

Research activity

1. In groups and using textbooks, research on the industrial manufacture of ammonia by the Haber process.
2. Compile a report then present it to the rest of the class.

The facts

The raw materials required for the manufacture of ammonia by the Haber process are:

i. **Nitrogen**-obtained from air by fractional distillation of liquid air.

ii. **Hydrogen**-obtained from natural gas (methane) or by electrolysis of acidified water or brine where electricity supply is cheap and readily available.

Optimum conditions needed during manufacture of ammonia are:

- A mixture of nitrogen and hydrogen in the ratio of 1:3 by volume.
- A pressure of between 200– 250 atmospheres.
- A temperature of between 400°C – 500°C.
- Iron catalyst in finely divided form and impregnated with aluminium oxide as a promoter of the catalyst.

Reactions that take place

When nitrogen and hydrogen react under the conditions mentioned above, only about 10% of ammonia is formed. The reaction is exothermic.

$$\text{Nitrogen} + \text{hydrogen} \xrightarrow[Al_2O_3]{Fe} \text{ammonia} + \text{heat}$$

$$N_2\,(g) + 3H_2(g) \xrightarrow[200 - 250\ \text{atm}]{400°C - 500°C\ Fe,\ Al_2O_3} 2NH_3(g) + \text{heat}$$

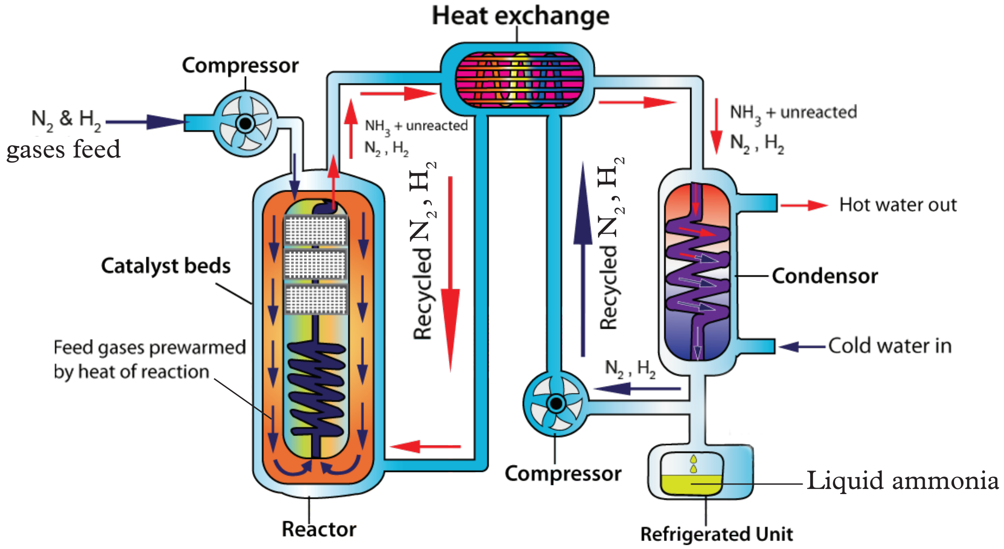


Fig 2.12: The Haber process

The heat produced during the reaction is taken back to the heat exchanger together with the unreacted nitrogen, hydrogen and the 10% ammonia formed. This helps to maintain the temperature needed for the reaction at 500°C. The ammonia produced and unreacted hydrogen and nitrogen are then taken to the liquifier or condenser where ammonia is liquified or dissolved in water. The unreacted nitrogen and hydrogen are re-circulated to the compressor.

Aqueous ammonia

Preparation of aqueous ammonia (ammonium hydroxide)

Activity 2.10

Apparatus and reagents

Beaker, delivery tube, filter funnel, source of ammonia.

Procedure

1. From the preparation flask, pass ammonia gas into water via an inverted filter funnel as shown in the figure 2.13 for some time. The filter funnel should just dip into the water surface.

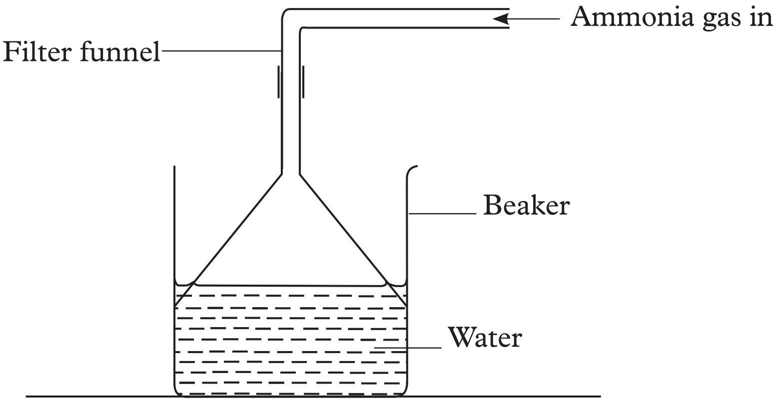


Fig. 2.13 Dissolving ammonia in water

2. Smell the aqueous solution in the beaker.
3. Record your observations in your notebook.

Study questions

1. Why is the funnel preferred to a delivery tube in dissolving ammonia in water?
2. Why is the ammonia gas used in this experiment not dried?

Discussion corner!

- Discuss the procedure and observations made in groups.
- Compile a report and present it to other class members.

I have discovered that…

After sometime, the water takes the smell of ammonia.

The facts

Ammonia gas dissolves in water to form aqueous ammonia or ammonia solution. Ammonia solution is also called **ammonium hydroxide.** The gas need not to be dried because water has been added to ammonia.

Ammonia + water $\rightleftharpoons$ ammonium hydroxide

NH_3 (g) + H_2O (l) $\rightleftharpoons$ NH_4OH (aq)

The funnel creates a large surface area for dissolution of ammonia as opposed to a delivery tube. In addition, if a delivery tube is used in place of the funnel, water would "suck back" into the hot preparation flask thereby cracking it.

Dissolution of ammonia in water is a reversible reaction. This is why aqueous ammonia produces ammonia gas at any temperature. It decomposes on warming to produce a strong smelling gas which is ammonia.

Properties of aqueous ammonia

Reaction of ammonia with dilute acids

Activity 2.11

Apparatus and reagents

Test tubes and test tube rack, 2M nitric acid, 1M sulphuric acid, 2M hydrochloric acid, methyl orange indicator, aqueous ammonia, labels.

Procedure

1. Place 2cm^3 of 1M sulphuric acid, 2M hydrochloric acid and 2M nitric acid in three separate test tubes and label them accordingly.
2. Add three drops of methyl orange indicator in each test tube.
3. Add aqueous ammonia drop wise into each test tube until a permanent pink orange colour is formed.

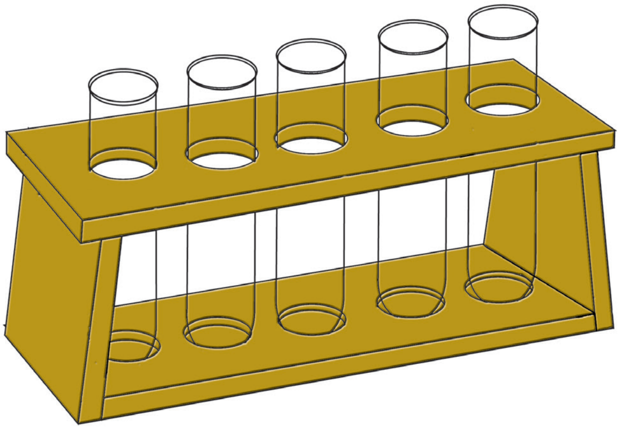


Fig 2.14: Experimental set-up for the reaction of ammonia solution with dilute acids

4. Record your observations.

Study questions

1. What name is given to the types of reactions that occur between aqueous ammonia and dilute acids?
2. Write the equations for the following reactions between aqueous ammonia and dilute:
   a. sulphuric acid
   b. hydrochloric acid
   c. nitric acid

Discussion corner!

- In your study groups discuss the observations made.
- Compare your findings with those of other groups in the class.

I have discovered that…

When ammonia solution is added to dilute acids, the solution turns yellow.

The facts

Aqueous ammonia neutralises dilute acids forming the corresponding ammonium salt and water only. The equations for the reactions that occur are as follows:

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

$$HCl\,(aq) + NH_4OH\,(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)$$

$$HNO_3\,(aq) + NH_4OH\,(aq) \longrightarrow NH_4NO_3(aq) + H_2O(l)$$

All ammonium salts are soluble in water and are normally prepared by neutralisation reaction. It is then followed by crystallisation of the salts.

Uses of ammonia

1. Large quantities of ammonia are used to make fertilisers.
2. Ammonia gas is used in the manufacture of nitric acid.
3. Liquid ammonia is used as a refrigerant in large-scale refrigerating plants and factories.
4. Ammonia solution is used as a solvent during cleaning (in laundries).
5. Ammonia is used in the manufacture of ammonium salts such as ammonium

Fig. 2.15: Ammonium nitrate fertiliser

chloride, which is used in dry cells and ammonium carbonate used in smelling salts. Smelling salts produce ammonia gas slowly at room temperature. This gas acts on the brain and prevents fainting or dizziness. In other words, it arouses consciousness.

6. Ammonia is used in the manufacture of dyes, wood pulp, plastics and fibres such as nylon.

Self-evaluation Test 2.4

1. What precaution would you take when preparing ammonia in the laboratory?
2. Calcium oxide is preferred in drying ammonia gas over other drying agents. Explain.
3. Uwimana, is a senior 3 student. She wanted to confirm whether a gas in a gas jar was ammonia. What chemical test would you advise her to carry out?
4. Draw the apparatus used to prepare aqueous ammonia in the laboratory.
5. Large-scale preparation of ammonia is carried out through the Haber process. Name two raw materials and two conditions required for this process.
6. Ammonia (NH_3) is a very useful chemical but for many years people could not make large amount of it. In 1908, a chemist called Fritz Haber invented the process for manufacturing ammonia on a large scale. That process is called the Haber process after its inventor.

   If nitrogen and hydrogen are mixed together with the right conditions, they will react to form ammonia. The chemical reaction has to be carried out at 450°C with a pressure 200 atmospheres air pressure.

   Finally, the reaction needs a catalyst (iron) to make it work. Most of the nitrogen and hydrogen gas are not used at first so they are channelled back into the reaction vessel until they are all used up. The chemical equation for the reaction is:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Use the write up above to answer these questions

(a) Ammonia is made out of many atoms.
   (i) How many hydrogen atoms are there in ammonia molecule?
   (ii) How many nitrogen atoms are there in an ammonia molecule?
(b) What is the most common use of ammonia?
(c) Who invented the Haber process?
(d) What does the Haber process produce?
(e) What type of catalyst is used in the Haber process?
(f) What happens to the unreacted hydrogen and nitrogen after the ammonia is removed.
(g) What is the ratio of nitrogen to hydrogen by volume entering the catalyst chamber?
   A. 3 volumes of H_2 : 3 volumes of N_2

B. 1 volumes of H_2 : 3 volumes of N_2
C. 3 volumes of H_2 : 1 volumes of N_2
D. 1 volumes of H_2 : 1 volumes of N_2

c) Nitric acid

Laboratory preparation of nitric acid

Activity 2.12

Apparatus and reagents

Retort glass bottle, round or flat-bottomed flask, concentrated sulphuric acid, potassium nitrate solution, glass cork, clamp stand, source of heat, running water from a tap.

Procedure

1. Arrange the apparatus as shown below.

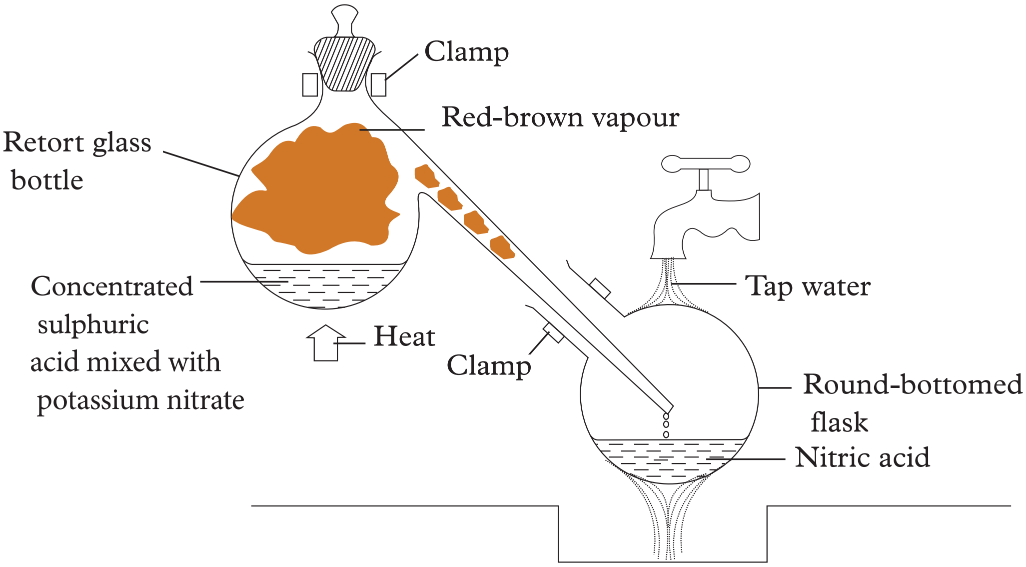


Fig 2.16: Laboratory set up for the preparation of nitric acid

2. Put some potassium nitrate in the retort glass bottle. Add concentrated sulphuric acid to cover it.
3. Open the tap water to cool the round or flat bottomed flask.
4. Heat the retort glass bottle gently.
5. Note down your observations in the retort glass bottle and flask.
6. Keep the acid formed for the next activity.

Study questions

1. Why are all the apparatus made of glass in this experiment?
2. What is the importance of tap water?

Discussion corner!

- Discuss the observations made in your study group.
- Compare your findings with those of other groups.

I have discovered that…

When the mixture is heated, effervescence is seen and red-brown fumes are seen in the retort flask.

The facts

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

Potassium nitrate + sulphuric acid ⟶ potassium hydrogen sulphate + nitric acid

$$KNO_3(s) + H_2SO_4(l) \longrightarrow KHSO_4(aq) + HNO_3(aq)$$

Nitric acid is more volatile than concentrated sulphuric acid and hence it vaporises. These vapours condense to form nitric acid (a yellow liquid).

The red-brown vapours seen in the retort flask are as a result of the decomposition of some of the nitric acid by heat, forming red-brown nitrogen dioxide.

Nitric acid ⟶ water + nitrogen dioxide + oxygen

$$4HNO_3(aq) \longrightarrow 2H_2O(l) + 4NO_2(g) + O_2(g)$$

The nitric acid collected in this experiment is red-brown because it contains some dissolved nitrogen dioxide, otherwise, pure nitric acid is colourless. The red-brown colour can be removed by bubbling air through the acid.

The apparatus used in this experiment is entirely made of glass because hot nitric acid vapours is highly corrosive and would attack cork or rubber stoppers and tubings.

Potassium nitrate is preferred in preparing nitric acid in the laboratory because it has no water of crystallisation which can dilute the acid produced. Sodium nitrate can also be used but it is **hygroscopic**.

Sodium nitrate + sulphuric acid ⟶ sodium hydrogen sulphate + nitric acid

$NaNO_3(s) + H_2SO_4(aq) \longrightarrow NaHSO_4(aq) + HNO_3(aq)$

Note: Nitric acid is kept in dark bottles. This is because light decomposes the acid into water, nitrogen dioxide and oxygen.

Further activity

In groups, discuss the industrial manufacture of nitric acid by Ostwald process. Prepare a report on your findings and present it to the class.

Manufacture of nitric acid by Ostwald Process

This process involves the catalytic oxidation of ammonia.

Raw materials

The raw materials required include:

- **Ammonia**-obtained through the Haber process.

- Air and water.

The optimum conditions needed are:

- A temperature of between 850°C - 900°C
- Platinum - rhodium catalyst

Note: The air and ammonia must be free of dust and other impurities to avoid poisoning of the catalyst.

Reactions that take place

In the first stage, reaction in this process produces nitrogen monoxide and steam. The heat produced in this step maintains the catalyst temperature and therefore no further heating is required.

Ammonia + oxygen ⟶ nitrogen monoxide + steam

$$4NH_3\,(g) + 5O_2\,(g) \longrightarrow 4NO\,(g) + 6H_2O\,(g)$$

The nitrogen monoxide and steam are cooled by air from the compressor. In the second stage, nitrogen monoxide is further oxidised by the air to produce nitrogen dioxide.

Nitrogen monoxide + oxygen ⟶ nitrogen dioxide

$$2NO\,(g) + O_2\,(g) \longrightarrow 2NO_2\,(g)$$

In the third stage, nitrogen dioxide is passed through water to form nitric acid (HNO_3) and nitrous acid (HNO_2).

Nitrogen dioxide + water ⟶ nitrous acid + nitric acid

$$2NO_2\,(g) + H_2O\,(l) \longrightarrow HNO_2\,(aq) + HNO_3\,(aq)$$

Excess air is passed to oxidise nitrous acid to nitric acid.

Nitrous acid + oxygen ⟶ nitric acid

$$2HNO_2\,(aq) + O_2\,(g) \longrightarrow 2HNO_3\,(aq)$$

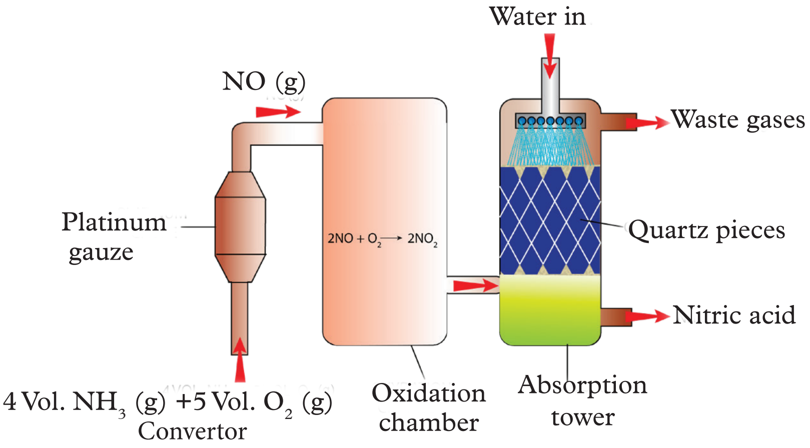


Fig 2.17: The Ostwald process

The nitric acid produced by this process is about 65 % pure. However, the purity of the acid can be increased by distillation. The unreacted air, ammonia and nitrogen dioxide are re-circulated back to the compressor.

Properties of dilute nitric acid

Activity 2.13

Apparatus and reagents

Test tubes, wooden splint, sodium carbonate, sodium hydroxide, copper(II) oxide, magnesium ribbon, dilute nitric acid, phenolphthalein indicator, zinc metal.

Procedure

1. Copy Table 2.4 in your notebook.
2. Carry out procedures "a-d" and use your results to fill the table.

Table 2.4: Investigating the properties of dilute nitric acid

Procedure	Observation	Conclusion
(a) Add 2 or 3 drops of phenolphthalein indicator to a solution of sodium hydroxide in a test tube. Then add dilute nitric acid drop by drop till a change is observed.		
(b) Add dilute nitric acid to copper(II) oxide in a test tube.		
(c) Put magnesium ribbon in a test tube. Add dilute warm nitric acid. Repeat the same experiment but now with very dilute cold nitric acid.		
(d) Put zinc in a test tube. Add to it dilute nitric acid.		

Study question

Write chemical equations for the reactions that take place in procedure a, b, c and d.

Discussion corner!

- In your study groups, discuss the results of the experiment and the study question above.
- Write a report and submit to the teacher for evaluation.

I have discovered that...

There is effervescence in reaction of dilute nitric acid with magnesium ribbon and zinc metal. The magnesium ribbon finally disapears and a colourless solution is formed. Sodium hydroxide solution is neutralised by dilute nitric acid in procedure (a).

The facts

Dilute nitric acid acts as any other mineral acid on carbonates, hydroxides and some oxides. The following are chemical properties of dilute nitric acid.

a) Nitric acid reacts with metal carbonates and hydrogen carbonates to form salt, water and carbon dioxide gas.

Metal carbonate + dilute nitric acid ⟶ metal nitrate + water + carbon dioxide

For example

Calcium carbonate + nitric acid ⟶ calcium nitrate + water + carbon dioxide

$CaCO_3$ (s) + $2HNO_3$ (aq) ⟶ $Ca(NO_3)_2$ (aq) + H_2O (l) + CO_2 (g)

b) It reacts with metal hydroxides to form salt and water. This is a neutralisation reaction.

Metal hydroxide + dilute nitric acid ⟶ metal nitrate + water. For example:

Sodium hydroxide + nitric acid ⟶ sodium nitrate + water

NaOH (aq) + HNO_3 (aq) ⟶ $NaNO_3$ (aq) + H_2O (l)

c) It reacts with metal oxides to form salt and water.

Metal oxide + dilute nitric acid ⟶ metal nitrate + water , for example:

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

CuO(s) + $2HNO_3$ (aq) ⟶ Cu $(NO_3)_2$(aq) + H_2O (l)

Nitric acid has oxidising properties that is, it donates its oxygen very easily). It therefore does not usually give off hydrogen with metals. This is because the hydrogen produced is immediately oxidised to form water.

For example, zinc metal reacts with dilute nitric acid to form zinc nitrate, nitrogen dioxide and water.

Zinc + nitric acid ⟶ zinc nitrate + nitrogen dioxide + water

Zn(s) + $4HNO_3$ (aq) ⟶ $Zn(NO_3)_2$ (aq) + $2NO_2$ (aq) + $2H_2O$(l)

An exception is the reaction of magnesium with very dilute and cold nitric acid.

Magnesium + nitric acid ⟶ magnesium nitrate + hydrogen

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

Properties of concentrated nitric acid

Concentrated nitric acid is a very powerful oxidising agent. Its oxidising power depends on the concentration, temperature and the type of reaction.

Activity 2.14

Apparatus and reagents

Iron (II) sulphate solution, dilute sulphuric acid, sulphur powder, 50% concentrated nitric acid, distilled water, barium chloride or barium nitrate solution, lead (II) nitrate solution, copper turnings.

Procedure

1. Copy Table 2.5 in your notebook.
2. Carry out the procedures outlined in the table below and record your observations and conclusions in the blank spaces.

Table 2.5: Properties of concentrated nitric acid

Procedure	Observation	Conclusion
(a) Add a few drops of dilute sulphuric acid to about 2 cm^3 of a solution of iron (II) sulphate in a boiling tube.		
(b) Add a few drops of concentrated nitric acid to the solution formed in (a) above.		
(c) Put a few copper turnings in a boiling tube and then add 50% concentrated nitric acid.		
(d) (i) Add about 2 cm^3 of concentrated nitric acid to sulphur powder in a boiling tube and warm it. (ii) Filter the mixture.		
(e) Dilute the filtrate obtained in (d) above with some water and divide it into two portions in two test tubes. (i) To the first portion in test-tube, add a few drops of barium chloride or barium nitrate solution. (ii) To the second portion in test-tube, add a few drops of lead (II) nitrate solution.		

Study questions

1. Which reactions show the oxidation property of concentrated nitric acid?

Discussion corner!

- In your study groups, discuss the observations made in the procedures above and the study questions.
- Draw conclusions and write a report.

I have discovered that…

Concentrated nitric acid is a strong oxidising agent. In all reactions, it acts as an oxidizing agent. Concentrated nitric acid is reduced to either nitrogen monoxide or nitrogen dioxide.

The facts

a) Reaction of concentrated nitric acid with iron (II) sulphate

Green iron (II) sulphate solution acidified with dilute sulphuric acid reacts with concentrated nitric acid to give a yellow-brown solution of iron (III) sulphate. Concentrated nitric acid therefore oxidises iron (II) ions to iron (III) ions. In the process, the acid is reduced to nitrogen monoxide gas.

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

The same type of reaction can take place with iron (II) chloride solution. However, in this case, addition of sulphuric acid is not necessary.

The nitrogen monoxide produced is readily oxidised to red-brown nitrogen dioxide gas by oxygen from the air.

Nitrogen monoxide + oxygen ⟶ nitrogen dioxide

$$\underset{\text{colourless}}{2NO\,(g)} + O_2\,(g) \longrightarrow \underset{\text{red-brown}}{2NO_2\,(g)}$$

b) Reaction of concentrated nitric acid with copper metal

Copper metal reacts with 50% concentrated nitric acid producing copper (II) nitrate, water and nitrogen monoxide gas.

Copper + 50% conc. nitric acid ⟶ copper (II) nitrate + water + nitrogen monoxide

$$3Cu(s) + 8HNO_3\,(aq) \longrightarrow 3Cu\,(NO_3)_2\,(aq) + 4H_2O\,(l) + 2NO\,(g)$$

The nitrogen monoxide is oxidised to nitrogen dioxide by oxygen from the air.

Concentrated nitric acid reacts with copper to form copper (II) nitrate, water and nitrogen dioxide.

Copper + nitric acid ⟶ copper (II) nitrate + water + nitrogen dioxide

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

The same type of reaction takes place when concentrated nitric acid reacts with zinc.

Zinc + nitric acid ⟶ zinc nitrate + water + nitrogen dioxide

$$Zn(s) + 4HNO_3\,(l) \longrightarrow Zn(NO_3)_2(aq) + 2H_2O\,(l) + 2NO_2(g)$$

c) Reaction of concentrated nitric acid with sulphur

Concentrated nitric acid reacts with sulphur to form sulphuric acid, water and nitrogen dioxide.

Sulphur + nitric acid ⟶ sulphuric acid + water + nitrogen dioxide

$$S(s) + 6HNO_3\,(aq) \longrightarrow H_2SO_4\,(aq) + 2H_2O\,(l) + 6NO_2\,(g)$$

When barium chloride or barium nitrate solution is added to the sulphuric acid produced in the reaction, a white precipitate of barium sulphate is formed. The sulphuric acid will also form a white precipitate of lead (II) sulphate when reacted with lead (II) nitrate solution. This is the test for sulphate ions (SO_4^{2-}). However, if sulphite ions (SO_3^{2-}) were present, the precipitate would have dissolved in dilute nitric acid.

The ionic equations for the reactions are:

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

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

Uses of nitric acid

1. Manufacture of nitrogenous fertilisers such as sodium nitrate, potassium nitrate, ammonium nitrate and calcium ammonium nitrate.
2. Manufacture of explosives like trinitrotoluene (TNT) and dynamite.
3. Manufacture of dyes.
4. Manufacture of plastics.
5. It is used as an oxidising agent in textile industries.
6. It is used in refining of gemstones.

Self-evaluation Test 2.5

1. Write two chemical equations for reactions in which nitric acid acts as an oxidising agent.
2. Nitric acid is not only an important reagent in the Chemistry laboratory but finds application in other fields. Elaborate.
3. A Senior 3 student added concentrated sulphuric acid to potassium nitrate crystals.
   (a) What observation did the student make?
   (b) Write a balanced chemical equation for the reaction that took place.
   (c) Why is potassium nitrate preferred in this reaction over other nitrates?
   (d) Sodium nitrate can also be used in this reaction. Why is it not preferred?
4. Give two raw materials and two conditions required in Ostwald process.
5. Which catalyst is used in large scale preparation of nitric acid?
   A. Vanadium (V) oxide
   B. Pure platinum
   C. Iron
   D. A platinum-Rhodium alloy (Pt - Rh)
6. Why is air filtered to remove dust and oil before being used in the manufacture of ammonia?
   A. They may poison the catalyst.
   B. They will colour the nitric acid.
   C. They will react with the nitric acid.

D. They will block the filters.

7. What word can be used to describe the reaction during absorption of ammonia into water?
8. Identify the most common fertiliser made from nitric acid.
   A. Nitrogen
   B. Ammonium sulphate
   C. Ammonium nitrate
   D. Manure
9. Write the equation for the action of heat on potassium nitrate.
10. Name the nitrate which on heating gives oxygen as the only gaseous product.
11. Describe what you will observe when concentrated nitric acid is added to copper.
12. Write the equation for the preparation of nitric acid from potassium nitrate.

Environmental effects during the manufacture of nitric acid

Health check

- *Concentrated nitric acid and its vapours are highly corrosive to the eyes, skin and mucous membranes. Dilute solutions of the acid on the other hand cause mild skin irritation and hardening of the epidermis.*
- *Contact with concentrated nitric acid burns the skin and produces deep painful wounds. Eye contact can cause severe burns and permanent eye damage. Inhalations of high concentrations can lead to severe respiratory irritations and delayed effects including pulmonary oedema which may be fatal.*
- *Ingestion of nitric acid may result in burning and corrosion of the mouth, throat and stomach. An oral dose of 10 ml can be fatal in human beings.*

Always be careful when handling nitric acid.

c) Nitrates

Nitrates are salts formed when the hydrogen ions in nitric acid are replaced by a metal or ammonium ion.

Laboratory preparation of nitrates

Activity 2.15

Apparatus and reagents

Spatula, Bunsen burner, filter funnel, filter papers, beakers, stirring rod, measuring cylinder, 1M sodium hydroxide solution, magnesium ribbon, calcium carbonate, 1M nitric acid.

Procedure

1. Measure 25 cm^3 of 1M nitric acid solution and place in a beaker. To this, add 25 cm^3 of 1M sodium hydroxide solution and stir. Heat the resulting solution to dryness.
   - Note down your observations.
2. Add small pieces of magnesium ribbon to 25 cm^3 of dilute nitric acid solution in a beaker. Filter the solution to remove excess magnesium and heat the resulting solution to dryness.
   - Note down your observations.
3. Add excess calcium carbonate to 25 cm^3 of nitric acid in a beaker. When the reaction is complete, filter the mixture and heat the filtrate to dryness.
   - Note down your observations.

Study questions

1. Why is it important to add excess magnesium ribbons and calcium carbonate in procedures 2 and 3 respectively?
2. What is the colour of the salt formed in each procedure above?

Discussion corner!

- Discuss the observations made in your study group.
- Present your finding to the class.

I have discovered that...

There are three methods used to prepare nitrates in the laboratory. In all these cases, white solid crystals are formed.

The facts

The three methods used to prepare nitrates are:

(i) **Neutralisation of dilute nitric acid by a base/alkali**.

For example:

$$HNO_3\,(aq) + NaOH\,(aq) \longrightarrow NaNO_3(aq) + H_2O(l)$$

When the resulting solution is heated to dryness, a white solid nitrate is obtained. This method is suitable for preparation of nitrates of potassium, sodium and ammonium.

(ii) **Reaction between dilute nitric acid and a metal carbonate**.

Examples of such reactions include:

$$2HNO_3\,(aq) + CaCO_3(s) \longrightarrow Ca(NO_3)_2(aq) + CO_2\,(g) + H_2O(l)$$

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

The solution formed is then evaporated to concentrate it and allow crystallisation to take place during cooling. The mother liquor is poured off and the crystals dried between filter papers.

(iii) **Action of nitric acid on a metal**.

This method is not suitable for use with very reactive metals like sodium and potassium. It is also not suitable for preparing salts that are readily oxidised. An example of a salt that can be prepared by this method is magnesium nitrate.

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

Properties of nitrates

Action of heat on nitrates

Activity 2.16

Apparatus reagents

5 test tubes, labels, blue litmus paper, wooden splint, sodium nitrate, potassium nitrate, copper (II) nitrate, lead(II) nitrate, silver nitrate (if available).

Procedure

1. Copy Table 2.6 in your notebook.

Table 2.6: Action of heat on metal carbonates

Metal nitrate being heated	Observation made
Sodium nitrate	
Potassium nitrate	
Copper(II) nitrate	
Lead nitrate	
Silver nitrate	

2. Put a spatulaful of each of the metal nitrates provided in different test tubes and label them accordingly.
3. Heat each nitrate.
4. Test any gas evolved using a glowing splint and moist blue litmus paper.
5. Observe the residue both when hot and cold.
6. Record all observations in a table.

Study questions

1. What is the colour of the gas or the gas mixtures produced in each case?
2. What is the colour of the hot and cold solid residues in each case?

Discussion corner!

- In your study groups, discuss the results obtained in the experiment and the study questions above.
- Prepare a report on how nitrates are affected by heat and present it to the teacher for evaluation.

I have discovered that…

The ease of decomposition of metal nitrates increases with decrease in reactivity of the metal constituting the nitrate. The nitrates decomposed form a red-brown acidic gas and other reaction produce a gas that relights a glowing splint.

The facts

All nitrates decompose on heating.

(i) Sodium and potassium nitrates decompose on heating to form metal nitrite and oxygen gas.

$$2KNO_3(s) \xrightarrow{heat} 2KNO_2(s) + O_2(g)$$

$$2NaNO_3(s) \xrightarrow{heat} 2NaNO_2(s) + O_2(g)$$

(ii) Nitrates of moderately reactive metals like copper, lead and zinc decompose to form metal oxides, nitrogen dioxide and oxygen gas.

$$2Pb(NO_3)_2(s) \xrightarrow{heat} 2PbO(s) + 4NO_2(g) + O_2(g)$$

$$2Zn(NO_3)_2(s) \xrightarrow{heat} 2ZnO(s) + 4NO_2(g) + O_2(g)$$

$$2Cu(NO_3)_2(s) \xrightarrow{heat} 2CuO(s) + 4NO_2(g) + O_2(g)$$

(iii) Nitrates of least reactive metals like silver and mercury decompose to form the respective metal, nitrogen dioxide and oxygen gas.

$$2AgNO_3(s) \xrightarrow{heat} 2Ag(s) + 2NO_2(g) + O_2(g)$$

$$Hg(NO_3)_2(s) \xrightarrow{heat} Hg(l) + 2NO_2(g) + O_2(g)$$

Fumes of mercury are poisonous and so the reaction should not be carried out under normal school laboratory conditions.

Lithium nitrate behaves differently from other alkali metals as follows.

$$4LiNO_3(s) \longrightarrow 2Li_2O(s) + 4NO_2(g) + O_2(g)$$

Test for nitrates

Activity 2.17

The Brown-ring test for nitrates

Apparatus and reagents

Test tubes, metal nitrate solution, freshly prepared iron (II) sulphate solution, concentrated sulphuric acid

Procedure

1. In a test-tube, mix equal volumes of a metal nitrate solution with freshly prepared iron (II) sulphate solution.
2. With the test-tube tilted at an angle, carefully pour concentrated sulphuric

acid down its wall so that two layers of solutions are formed.

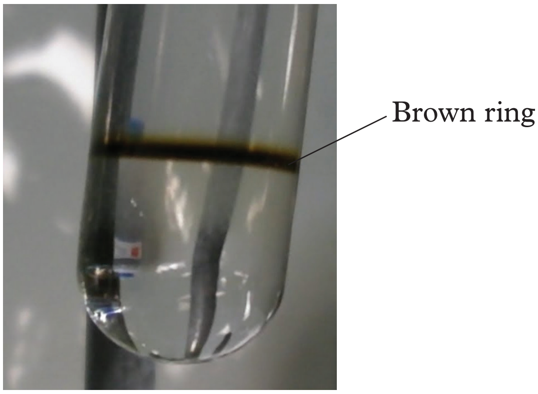


Fig 2.18: Brown-ring test for nitrates

3. Record your observation.
4. Shake the solution and record what happens.

Study questions

1. Why should we use freshly prepared iron (II) sulphate solution?
2. Explain what happened when you shook the test tube containing the mixture.

Discussion corner!

- Discuss in your study group the observations above.
- Compare your findings to those of other class member.

I have discovered that…

A brown ring forms between sulphuric acid and the rest of the solution. However, the ring disappears when the solution is shaken.

The facts

Concentrated sulphuric acid reacts with any nitrate salt to produce nitric acid. The nitric acid is reduced to nitrogen monoxide by iron (II) sulphate. The nitrogen monoxide formed reacts with unreacted iron (II) sulphate to form a brown nitroso-iron (II) sulphate ($FeSO_4.NO$) which appears as a **brown ring**. This is the test for nitrate ions presence in solution.

Iron(II) sulphate + nitrogen monoxide ⟶ nitroso-iron(II) sulphate

$$FeSO_4 (aq) + NO (g) \longrightarrow FeSO_4.NO (aq)$$

The ring disappears if the solution is shaken. This is because concentrated sulphuric acid mixes with the aqueous solution producing a lot of heat. This heat decomposes nitroso-iron(II) sulphate as shown in the equation below.

Nitroso-iron (II) sulphate ⟶ iron (II) sulphate + nitrogen (II) oxide

$FeSO_4.NO(aq) \longrightarrow FeSO_4(aq) + NO(g)$

Iron (II) sulphate solution is oxidised by oxygen from the air to iron(III) sulphate within a short time. This is the reason why it should always be freshly prepared.

Note: All nitrate salts are soluble in water. They all form nitrate ions when in solution which gives the "brown ring".

Uses of nitrates

1. They are used as fertilisers for example sodium nitrate, potassium nitrate, ammonium nitrate, calcium ammonium nitrate and NPK fertilizers.
2. Potassium nitrate is used in the manufacture of gun powder. (Gun powder is a mixture of potassium nitrate, charcoal and sulphur). When ignited, the mixture burns and explodes.
3. Ammonium nitrate is used as a fertiliser and also in the manufacture of explosives.
4. Nitrates in form of nitroglycerin are used in medicine for treatment of heart diseases.

Fig 2.19: NPK fertiliser

Further activity

Research and make a presentation on the uses of different compounds of nitrogen.

2.4 Environmental issues with nitrogen dioxide and nitrate fertilisers

Research activity

1. In groups, find out the environmental issues associated with nitrogen oxides and nitrate fertilisers.
2. Suggest ways that can be used to prevent pollution by nitrogenous compounds.
3. Write a report and hand it to the teacher for evaluation.

(a) Nitrogen dioxide

Nitrogen dioxide is found in the atmosphere as a result of industrial processes and emissions from automobile exhausts.

Nitrogen dioxide is toxic and can cause respiratory diseases.

Nitrogen dioxide also reacts with water vapour

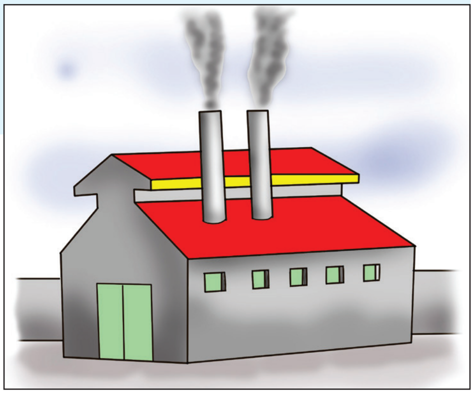


Fig. 2.20: Pollution by industries

in the atmosphere and/or rain water to form dilute nitric acid. Rain containing this acid or other acids is called "acid rain". Acid rain may interfere with animals and plants life on land and in water bodies. This is because the acid lowers the pH of soil and water making it difficult for some organisms to survive. Acid rain also attacks metallic objects such as iron roofs corroding them. This causes leaking in roofs.

Fig 2.21: Corroded iron sheets due to effects of acidic rain

(b) Nitrate fertilisers

The continued use of large quantities of nitrate fertilisers over long period of time pollutes the environment. This is because excess nitrates in the soil are washed down into rivers, lakes and other water bodies.

Once in water, nitrates cause aquatic plants (e.g. plankton and algae) to grow in large quantities (a process known **eutrophication)**. As a result dead plants accumulate faster than can be decomposed by bacteria and they settle to the bottom of the water bodies. Dead plant matter has a high **biological oxygen demand** (BOD) and some water bodies give up nearly all their dissolved oxygen in order to allow this natural decay process (oxidation) of the dead plant material to take place. This causes a decrease in the concentration of dissolved oxygen in water and this eventually leads to death of aquatic animals such as fish and plants.

Also due to large masses of dead plants and animals, the penetration of light in water decreases thereby affecting photosynthesis in aquatic plants.

If nitrates are consumed in drinking water, they are converted by bacteria in the intestines to more toxic nitrites. After absorption into the blood stream, the nitrites combine with haemoglobin forming **metahaemoglobin**. As a result, the oxygen carrying capacity of blood is reduced. Babies suffering from excess intake of nitrates have been known to change ther colour to "blue" due to lack of oxygen in their blood, a condition referred to as '**blue babies**'.

My environment, my life!

We should advise people in our community to avoid cultivating near river banks. This will prevent the washing away of fertilisers into water bodies hence prevent eutrophication.

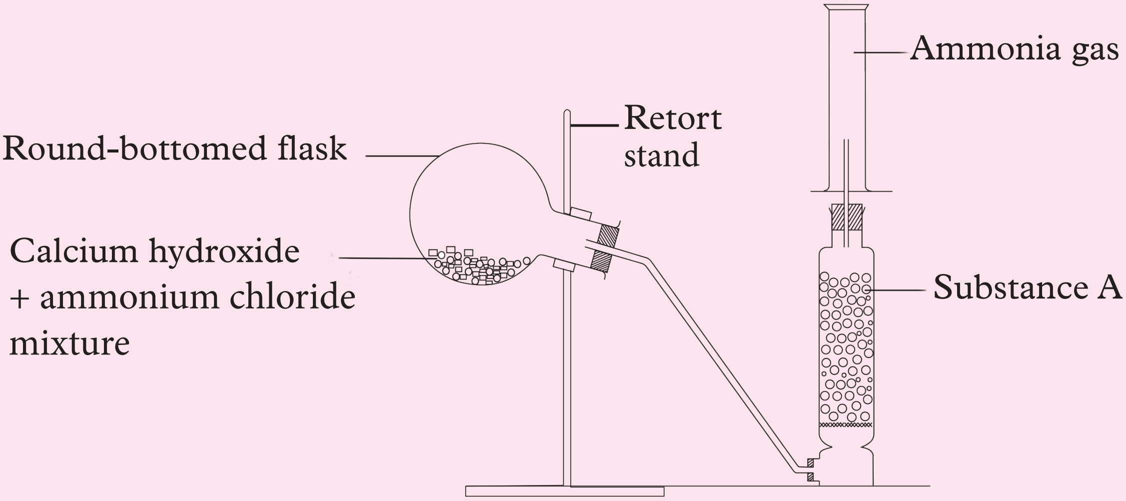
Unit summary 2

- Nitrogen can be isolated from air by fractional distillation of liquid air. The process involves removing other components of air like carbon dioxide, oxygen, noble gases and moisture (water vapour).
- In the laboratory, nitrogen is prepared by reacting ammonium chloride and sodium nitrite.
- Nitrogen is inert at room temperature because of the strong triple covalent bonds between the two nitrogen atoms which make the molecule.
- In the laboratory, nitrogen dioxide is prepared by reacting copper (II) nitrate with concentrated nitric acid.
- Nitrogen dioxide is a red-brown gas with an irritating pungent smell and dissolves in water to form nitrous and nitric acids. It is denser than air.
- Ammonia is an alkaline gas prepared in the laboratory by reacting calcium hydroxide and ammonium chloride. The flask containing reacting reagents should be slanting to prevent moisture that forms in the cooler parts of the apparatus from flowing back into the flask which can can cause it to crack.
- Large scale manufacture of ammonia is done through Haber process. The raw materials required are nitrogen and hydrogen in the ratio 1:3. Other conditions required are pressure of 200 - 250 atmospheres and a temperature of 400^0C - 500^0C.
- Ammonia is used in the manufacture of fertilisers, nitric acid, dyes, wood pulps, plastics and fibres.
- Laboratory preparation of nitric acid is done by reacting potassium nitrate and sulphuric acid.
- Nitric acid is prepared on large scale through Ostwald process. The raw materials needed are ammonia, air and water.
- Concentrated nitric acid is a strong oxidising agent. It is reduced to either nitrogen monoxide or nitrogen dioxide.
- Nitric acid is used in the manufacture of nitrogenous fertilisers, explosives, dyes and plastics. It is also used as an oxidizing agent in textile industries.
- Nitrates are salts formed when the hydrogen ion in nitric acid is replaced by a metal or an ammonium ion.
- The three methods used to prepare nitrates are neutralisation of dilute nitric acid by an alkali, reaction between dilute nitric acid and a metal carbonate and action of nitric acid on a metal.
- All nitrates decompose on heating. However, the degree of decomposition increases with decrease in reactivity of the metal forming the nitrate.

- Brown ring test is a test used to identify nitrates. It involves mixing equal volumes of a nitrate with freshly prepared iron (II) sulphate. Concentrated sulphuric acid is then carefully added.
- Nitrogen dioxide released into the atmosphere pollutes the environment. It dissolves in rain water forming acid rain. Acid rain destroys vegetation, animals and buildings. Excess use of nitrate fertilisers also leads to water pollution and eutrophication.

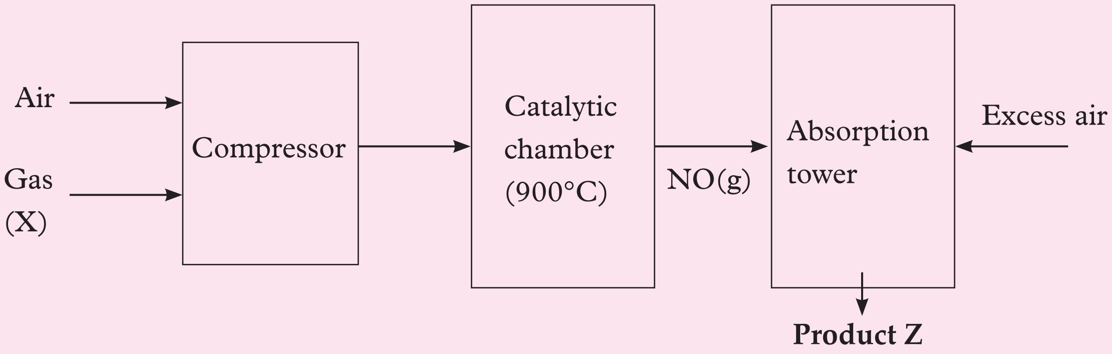
Test your Competence 2

1. Cultivating near water bodies leads to environmental pollution. Explain this with reference to effects of nitrogen compounds on the environment.
2. Nitrogen dioxide is a __________ gas in colour that has a __________ smell while ammonia is a colourless gas with a __________ smell.
3. Mutesi dissolved nitrogen dioxide gas in water. She later put blue and red litmus papers in the solution formed. The blue litmus paper turned red while there was no effect on the red litmus paper.
   (a) Explain the observations made above.
   (b) Write a chemical equation for the reaction that took place.
4. In the preparation of ammonia, a student set up an experiment as shown below.



(a) Identify and correct the mistake in the set up.
  (b) Name substance A and state its function.
  (c) Name the method used to collect ammonia gas in the set up and explain why it is suitable.
  (d) Write a chemical equation for the reaction that took place in the round-bottomed flask.

5. Large scale manufacture of ammonia is carried out through ______ process while manufacture of nitric acid in large scale is done through ______ process.
6. Name two raw materials and two conditions necessary for the large scale manufacture of ammonia.
7. Study the flowchart below then answer the questions that follow.



  (a) Name the process illustrated in the flowchart above ______________.
  (b) Identify (i) Gas X _________
      (ii) Product Z _________
  (c) Write a chemical equation for the reaction that takes place in the catalytic chamber.
  (d) Why is it recommended that all the apparatus in the laboratory preparation of this product be made of glass?

8. The following methods are suitable for preparation of nitrates except _______?
  A. Reacting dilute nitric acid with sodium hydroxide.
  B. Reacting sodium metal with nitric acid.
  C. Reacting nitric acid with lead carbonate.
  D. Reacting dilute nitric acid with magnesium.

9. Write **true** or **false** for each of the following:
  (a) All nitrates decompose on heating. _________
  (b) Ammonia is the only common basic gas. _________
  (c) Nitrogen dioxide dissolves in water to form nitric acid only. _________
  (d) Brown-ring test is used to identify nitrates.

UNIT 3

Sulphur and its inorganic compounds

Key unit competency

To be able to relate the properties of sulphur and its compounds to their uses, describe how some compounds of sulphur are prepared and discuss the related environmental issues.

Learning objectives

By the end of this unit, I should be able to:

- Recall the occurrence, extraction, properties and uses of sulphur.
- Prepare, test and collect sulphur dioxide gas and explain the impact of sulphur oxides on the environment.
- Describe the industrial preparation of sulphuric acid by the contact process.
- Develop skills in observation in preparing sulphur dioxide gas and testing for the presence of sulphates and sulphites in given solutions.
- Protect natural resources.
- Develop self confidence in discussions and presentation of research findings.
- Develop a culture of working in a team during research and discussions.
- Respect procedures during experiments.

Mind teaser

Study the chart below carefully. What is the chart about?

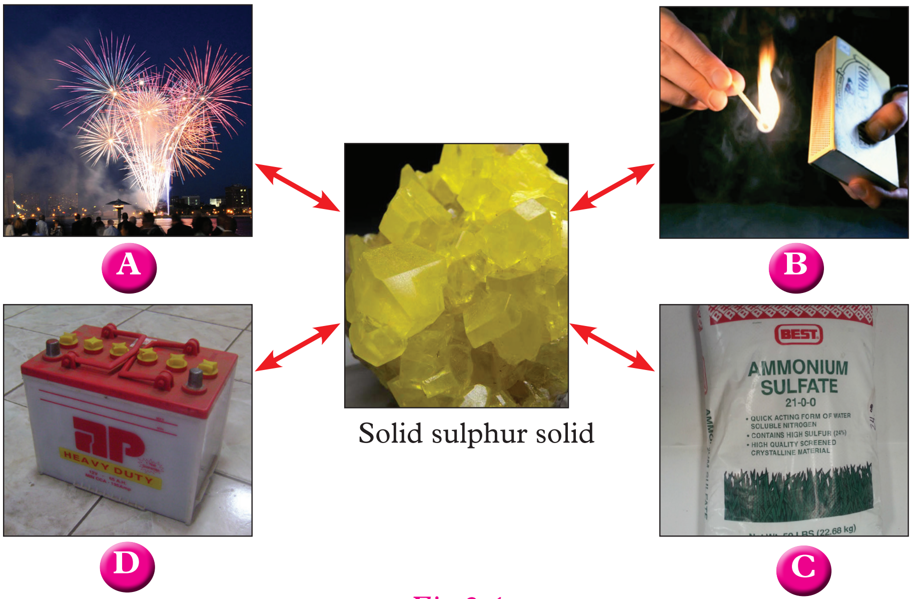


Fig 3.1

Try to identify the properties or uses of sulphur and its compounds associated with the pictures above. Which other ways do we use sulphur in our daily lives?

Introduction

In senior one, you learnt that sulphur is a non-metal located in group VI and period 3 of the Periodic Table.

It is a **bright yellow** non–metallic element which has been known since ancient times. One of the many compounds of sulphur is sulphuric acid. This acid is very important in Chemistry. For example, it is a major component of car batteries. It is also used to make fertilisers such as ammonium sulphate. Other sulphur compounds include sulphates, sulphites and gases like hydrogen sulphide, sulphur dioxide and sulphur trioxide.

3.1 Occurrence, extraction, properties and uses of sulphur

Research activity

In your study groups, research and discuss the following:

1. Occurrence of suphur in nature.
2. Extraction of sulphur (focus on sulphur obtained as a by product of crude oil refining and metal sulphide ore refining).
    (i) What process is used to extract sulphur from its underground deposits?
    (ii) Describe the process named in (i) above.
    (iii) Give some uses of sulphur.
3. Prepare a report and hand it to the teacher for evaluation.

The facts

Occurrence

Sulphur occurs naturally in volcanic regions or regions where volcanoes were formerly active. As a free element sulphur is found occurring about 150 m underground in Texas and Louisiana in USA and Sicily in Italy. Sulphur also occurs as underground deposits in Japan.

Sulphur can occur as a "free" element or combined in compounds such as:

- Hydrogen sulphide found in petroleum, coal gas and natural gas.
- Metal sulphides, for example, zinc blende (ZnS), iron pyrites (FeS_2), copper pyrites ($CuFeS_2$), galena (PbS) etc.
- Metal sulphates, for example gypsum ($CaSO_4 . 2H_2O$).

Extraction of sulphur

Sulphur is normally extracted from its underground deposits by **Frasch process**, developed by American Herman Frasch. The sulphur deposits usually occur at about 150 metres below the ground level in beds of about 30 metres diameter. These deposits lie beneath layers of clay, sand and limestone. It is therefore not easy to extract it using ordinary mining methods.

The basic principle behind Frasch process is melting the underground sulphur and then pumping it to the earth's surface. **Three concentric pipes** of different diameters are drilled through into sulphur beds. Super heated water at 170°C and a pressure of 10 atmospheres is pumped down the outermost pipe to melt the sulphur deposits. The melting point of sulphur is 113°C.

Hot compressed air at pressure of 15 atmospheres is pumped down the innermost pipe. This changes the molten sulphur into a froth which is then forced up through the middle pipe and collects in large containers where it solidifies to yellow solid sulphur. The sulphur obtained this way is usually over 99.8% pure.

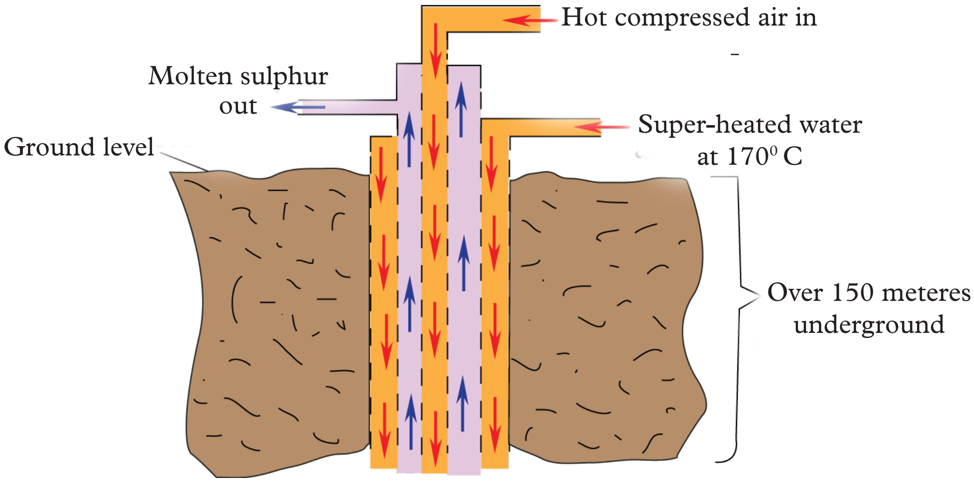


Fig 3.2: Extraction of sulphur by Frasch process

Other methods used to extract sulphur

Sulphur can also be extracted from crude oil gases. Crude oil gases include hydrogen sulphide (H_2S), carbon dioxide (CO_2) and methane (CH_4). This gaseous mixture is passed through an alkaline solution. Since hydrogen sulphide and carbon dioxide are acidic gases, they are absorbed by the alkaline solution and can be regenerated by heating the solution. The hydrogen sulphide produced is later oxidised to sulphur by air. The hydrogen sulphide is oxidised in the presence of aluminium oxide catalyst as shown in the equation below. This is called the **Claus process**.

$$2H_2S(g) \ + \ O_2(g) \xrightarrow{Al_2O_3} 2S(s) \ + \ 2H_2O(g)$$

Sulphur can also be produced during the extraction of metals from metal sulphides like zinc blende. When the ore is roasted in air it forms zinc and sulphur dioxide. The sulphur dioxide is reduced by reacting it with hydrogen sulphide to give sulphur.

$$2ZnS(s) + 3O_2(g) \longrightarrow 2ZnO(s) + 2SO_2(g)$$
$$SO_2(g) + 2H_2S(g) \longrightarrow 3S(s) + 2H_2O(g)$$

3.2 Properties and uses of sulphur

Activity 3.1

Apparatus and reagents

Test tubes, beakers, spatula, candle, sulphur powder, carbon dioxide, distilled water, boiling tube, Bunsen burner, deflagrating spoon, carbon disulphide, iron filings, concentrated sulphuric acid, concentrated nitric acid.

Procedure

1. Put about 3 cm^3 of distilled water in a test tube and add a half full spatula of sulphur powder. Shake well and record your observations.
2. Put about 3 cm^3 of carbon disulphide in a test tube and add a half full spatula of sulphur powder. Observe and record what you have seen.
3. Place three spatula endful of sulphur in a boiling tube. Heat gently using candle flame. Do not let the molten sulphur darken in colour. Tilt the boiling tube slightly so that you can see how easily the liquid flows.
   - Comment on its colour and viscosity.
4. Now continue heating the dark molten sulphur using a Bunsen burner flame. From time to time tilt the boiling tube to see how easily the molten sulphur flows. Do not let any liquid flow out of the boiling tube.
   - Comment on the colour and viscosity.
5. Place sulphur powder on a deflagrating spoon and heat in a Bunsen burner flame until it starts to burn. Test the gas evolved with wet blue litmus paper and record your observations.

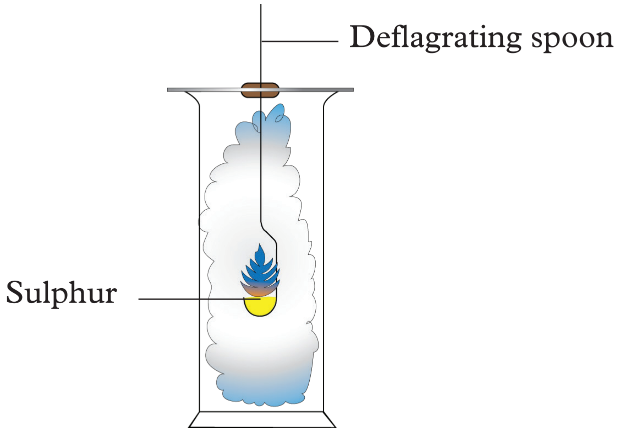


Fig 3.3: Burning sulphur in a deflagrating spoon

6. Mix equal masses of sulphur powder and iron fillings in a test tube and heat gently then strongly.
   - Record your observations.
7. To 3 cm^3 of concentrated sulphuric acid in a test tube, add a half spatula endful of sulphur powder and warm the mixture gently. Test any gases produced using wet blue litmus paper.

Caution! This procedure should be carried out in a fume cupboard since sulphur dioxide is a poisonous gas.

8. Repeat procedure (7) above but now using concentrated nitric acid and record your observations.

Caution! This procedure should also be carried out in a fume cupboard as nitrogen dioxide gas produced is poisonous.

Study questions

1. What is the colour of flame produced by burning sulphur?
2. Write chemical equations for the reactions that took place in procedures 5-8.

I have discovered that…

Sulphur does not dissolve in water but dissolves in carbon disulphide. When heated, sulphur undergoes a series of physical changes. Sulphur burns with a blue flame.

The facts

Physical properties of sulphur

1. It is a bright yellow solid substance at room temperatures.
2. Sulphur is soluble in organic solvents such as carbon disulphide and benzene but it is insoluble in polar solvents like water.
3. It melts at 113°C and boils at 444°C.

Chemical properties of sulphur

1. Sulphur burns easily with a blue flame in the presence of air to give fumes which turn wet blue litmus paper red. The gas formed is sulphur dioxide. Sulphur dioxide dissolves in water to form sulphurous acid (H_2SO_3).

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$
$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

2. Sulphur combines with metals directly and easily to form the corresponding metal sulphides. For example, when a mixture of iron and sulphur is heated, it produces a red hot glow and iron (II) sulphide salt is formed. This reaction is exothermic and once it starts, the glow continues even after heat is withdrawn.

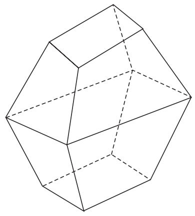
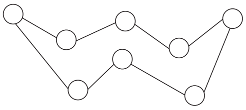
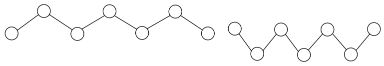
$$Fe(s) + S(s) \xrightarrow{\text{heat}} FeS(s)$$

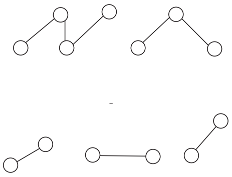
Iron (II) sulphide is a black solid. Similarly, hot powdered copper combines with sulphur on heating to form copper (II) sulphide.

$Cu(s) + S(s) \xrightarrow{heat} CuS(s)$

3. When heated with concentrated sulphuric acid, sulphur is oxidised to sulphur dioxide which turns wet blue litmus paper red.
   $S(s) + 2H_2SO_4(l) \longrightarrow 3SO_2(g) + 2H_2O(l)$
4. Hot concentrated nitric acid oxidises sulphur to sulphuric acid while the acid itself is reduced to brown nitrogen dioxide gas. This is a slow reaction.
   $S(s) + 6HNO_3(aq) \longrightarrow H_2SO_4(aq) + 6NO_2(g) + 2H_2O(l)$
   Note: These experiments should be done in the fume cupboard since both SO_2 and NO_2 are poisonous.
5. When heated, sulphur undergoes several changes as shown in the following table 3.1.

Table 3.1: Changes that take place in sulphur when heated at different temperatures

Temperature(°C)	Physical change	Explanation
113	Sulphur melts to give a mobile amber coloured liquid; further increase in temperatures makes the liquid to darken.	Intermolecular forces are broken by heat. 
180	The liquid becomes viscous. Colour changes gradually from red to black.	The S_8 rings that forms sulphur molecules open up forming long chains. These chains entangle making them viscous. 
250	Liquid becomes very thick and viscous and cannot flow.	The rings of 8 atoms open up as the molten sulphur is heated more strongly. The long chain entangle and make liquid sulphur viscous. 

Temperature(°C)	Physical change	Explanation
Above 250	Colour becomes brighter again and the liquid becomes mobile again.	Short units escape from liquid as gas (sulphur vapour). 

Uses of sulphur

1. It is used in the manufacture of sulphuric acid.
2. To vulcanise (harden) rubber; ordinary rubber is soft and elastic. However, when heated with sulphur, rubber becomes hard and tough making it suitable for use in making tyres.
3. Manufacture of matches, fireworks and gun powder. The explosive head of a match stick contains sulphur among other substances.
4. Used as a fungicide. Powdered sulphur is usually mixed with petroleum jelly and the mixture used as a skin ointment to treat fungal diseases in animals and plants.
5. Used in the production of important chemicals such as carbon sulphide (CS_2) and calcium hydrogen sulphite $Ca(HSO_3)_2$. Calcium hydrogen sulphite is used during the manufacture of paper to bleach wood pulp.
6. Making sulphur dyes which are used for dyeing cotton and artificial hair.
7. Making of sulphur drugs. Some antimalarial drugs contain sulphur.
8. Manufacture of sodium thiosulphate used in photography.

(a) (b) (c)

Fig 3.4: Products made of sulphur

Self-evaluation Test 3.1

1. In nature, sulphur occurs as a free element or in combined states. Name any three compounds of sulphur.

2. When sulphur is heated, it melts and flows easily. However, on further heating the liquid becomes viscous and does not flow. Explain.
3. State the observations made when:
   a) Sulphur is heated in air.
   b) A mixture of sulphur and concentrated nitric acid is heated.
   c) Write equations for the reactions in(a) and (b) above.
4. State four uses of sulphur.

3.3 Inorganic compounds of sulphur and their properties

Sulphur dioxide

Activity 3.2

Apparatus and reagents

Flat bottomed flask, dropping funnel, gas jar, wash bottle, sodium sulphite, dilute hydrochloric acid, Bunsen burner, flower petals, concentrated sulphuric acid, sodium hydroxide solution, blue and red litmus papers, beakers, distilled water.

Procedure

Caution: Sulphur dioxide is poisonous. This experiment should be carried out in a fume chamber.

1. Place sodium sulphite in a flat bottomed flask and set up the apparatus as shown below.

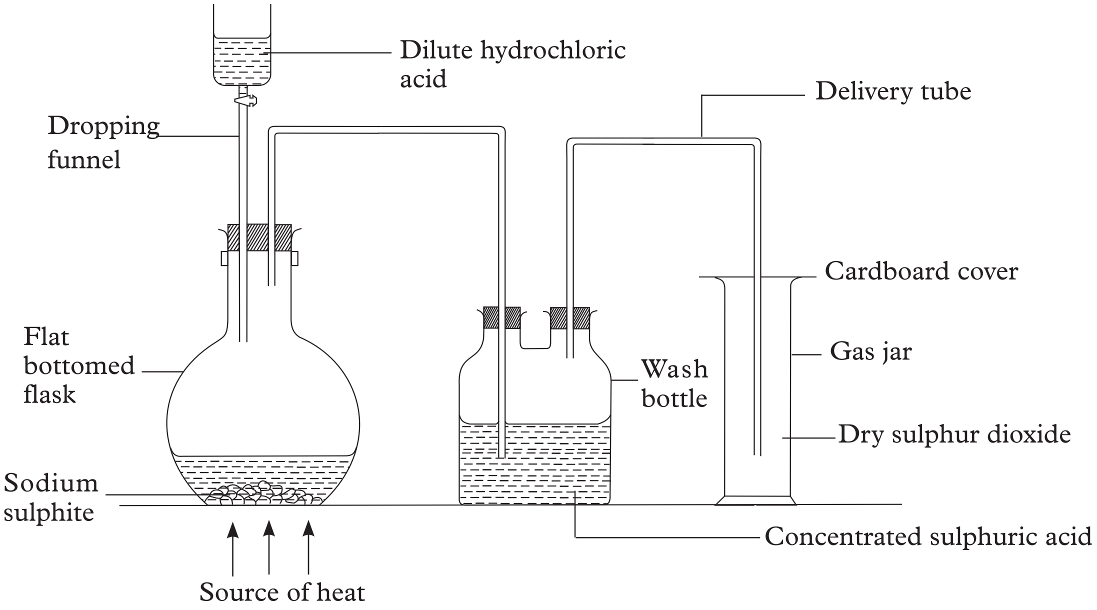


Fig 3.5:Preparation of sulphur dioxide using dilute hydrochloric acid and sodium sulphite

2. Run down the hydrochloric acid into flat bottomed flask.
3. Collect six gas jars of sulphur dioxide to be used in the next experiments.
    - Note down the colour of the gas collected.
4. Invert a gas jar full of the gas in water using a funnel as shown below. Drop blue and red litmus papers into the solution.

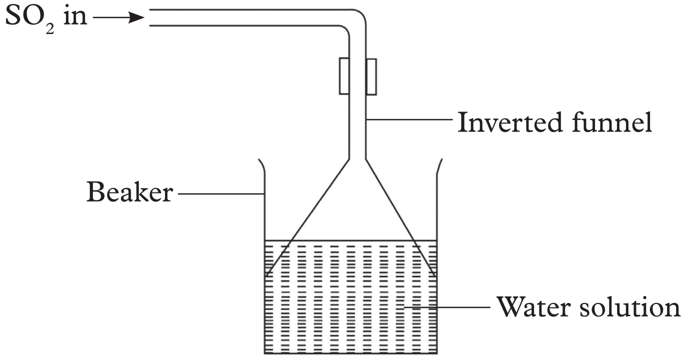


Fig 3.6: Reacting sulphur dioxide with water

    - State and explain your observations.

5. Drop pieces of wet blue and red litmus papers into a test tube containing sulphur dioxide.
    - Note down your observations.
6. Bubble sulphur dioxide through sodium hydroxide solution. Drop blue and red litmus papers into the resultant solution.
    - Note down your observations.
7. Drop red flower petals into a gas jar containing sulphur dioxide.
    - Record your observations after 5 minutes.

Study questions

1. In the preparation of sulphur dioxide gas, why was the gas passed through concentrated sulphuric acid?
2. Name the method used to collect sulphur dioxide gas and explain why it is suitable.

Discussion corner!

1. In your study groups, discuss the observations made in the experiment and the study questions above.
2. Prepare a report and present it in class.

I have discovered that…

Sulphur dioxide is collected by downward delivery. Wet litmus papers and coloured flower petals when placed in a jar containing sulphur dioxide, lose their colour. Blue litmus paper *dropped in a solution of sulphur dioxide turns red. A solution of sulphur dioxide in water turns blue litmus paper red.*

The facts

Sulphur dioxide is prepared in a fume cupboard or in the open air because it is poisonous. When a mixture of sodium sulphite and dilute hydrochloric acid is warmed, effervescence occurs producing a colourless gas. The gas is sulphur dioxide.

$$Na_2SO_3\,(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + SO_2(g) + H_2O(l)$$

The gas is passed through concentrated sulphuric acid to dry it. It is then collected by downward delivery since it is denser than air. Sulphur dioxide gas may also be prepared in the laboratory by heating a mixture of copper turnings with concentrated sulphuric acid using the following set up.

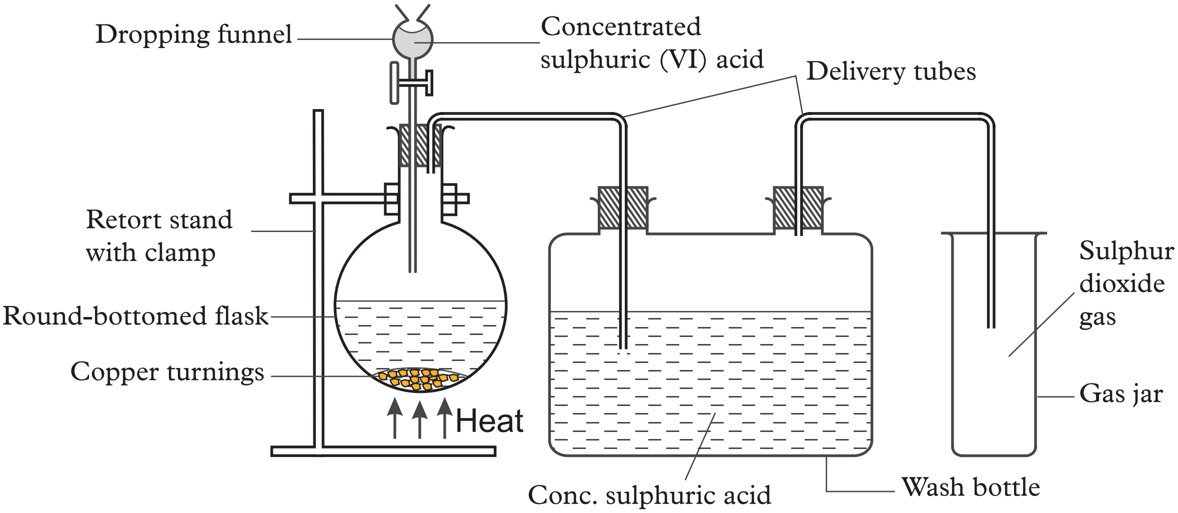


Fig 3.7: Preparation of sulphur dioxide gas using concentrated sulphuric acid and copper turnings

When the reaction mixture in the flask gets hot, effervescence occurs and a colourless gas is collected by downward delivery. The equation for the reaction is as follows.

$$Cu\,(s) + 2H_2SO_4(l) \longrightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$$

When the gas jar is full of the gas, the moist potassium dichromate paper placed at the mouth of the gas jar turns from orange to green because it is oxidised.

Properties of sulphur dioxide

Physical properties

- It is a colourless gas.
- It has a pungent choking smell.

- It is very soluble in water. This explains why water rushed into the gas jar in procedure 4 above.
- It is denser than air. This explains why it is collected by downward delivery.

Chemical properties

1. Reaction with water

Sulphur dioxide reacts with water to form sulphurous acid. This is a weak acid that partially dissociates to yield few hydrogen and sulphite ions. The presence of hydrogen ions explains why a solution of the gas turns blue litmus paper red.

$$H_2O\ (l) + SO_2\ (g) \rightleftharpoons H_2SO_3(aq) \rightleftharpoons 2H^+(aq) + SO_3^{2-}(aq)$$

2. Reaction with alkalis

When sulphur dioxide is passed through sodium hydroxide solution, neutralisation reaction occurs. With excess alkali, sodium sulphite is formed as shown in the following equation.

$$2NaOH\ (g) + SO_2\ (g) \longrightarrow Na_2SO_3\ (aq) + H_2O(l)$$

With excess sulphur dioxide, sodium hydrogen sulphite is formed as shown in the equations below.

$$NaOH\ (aq) + SO_2\ (g) \longrightarrow Na_2SO_3(aq) + H_2O\ (l)$$

$$Na_2SO_3(aq) + H_2O\ (l) + SO_2\ (g) \longrightarrow 2NaHSO_3(aq)$$

These solutions have bleaching properties.

3. Bleaching action

Flower petals and wet litmus papers when placed in a gas jar of sulphur dioxide, they become bleached. However, dry litmus papers are not bleached. This is because sulphur dioxide gas must first dissolves in water to form sulphurous acid. It is this acid that thus reduces the dye in petals and wet litmus papers by removing the oxygen atom from them. The dye is thus said to be bleached.

$$H_2O(l) + SO_2(g) \rightleftharpoons H_2SO_3(aq) \rightleftharpoons SO_3^{2-}(aq) + 2H^+(aq)$$

$$SO_3^{2-}(aq) + \underset{\text{Coloured}}{\text{dye}} \longrightarrow SO_4^{2-}(aq) + \underset{\text{Colourless}}{\text{(Dye-O)}}$$

Since this type of bleaching occurs by removal of an oxygen atom from a substance, it is called bleaching by reduction. Sulphur dioxide can only bleach in presence of water.

The colour of the dye can be restored if the dye is exposed to environmental oxygen in the presence of sunlight. That is why old newspapers that had been bleached by reduction turn yellow/brown on exposure to air and sunlight after a long time.

Note: Bleaching means removing coloured materials from substances or removing colour.

Laboratory test for sulphur dioxide gas

Activity 3.3

Apparatus and reagents

Source of sulphur dioxide gas (from a generator), acidified potassium manganate (VII) solution, acidified potassium dichromate (VI) solution, iron (III) chloride solution, test tubes, labels.

Procedure

1. Place about 5 cm^3 of acidified potassium manganate (VII) solution, acidified potassium dichromate (VI) solution and iron (III) chloride solution into three separate test tubes. Label them as A, B and C.
2. Pass sulphur dioxide gas from a generator through each of the solution in the test tubes.
3. Copy table 3.2 in your notebooks and record your observations and conclusions.

Table 3.2: Observations and conclusions on test for sulphur (IV) oxide.

Solution tested	Observation	Conclusion
Acidified potassium manganate (VII) solution		
Acidified potasium dichromate (VI) solution		
Iron (III) chloride solution		

Discussion corner!

1. In pairs, discuss the observations made in each case.
2. Present your results to the class.

I have discovered that…

Sulphur dioxide decolourises an acidified solution of potassium permanganate. The gas also changes the colour of potassium dichromate from orange to green. A solution of iron (III) chloride changes from brown to green.

The facts

The three reagents potassium manganate (VII), potassium dichromate (VI) and iron (IIII) chloride are used to test for sulphur dioxide gas. The above colour changes and the bleaching action of sulphur dioxide are used to identify the gas.

Uses of sulphur dioxide

1. It is an intermediate in the manufacture of sulphuric acid.
2. Used as a preservative in jam and fruit juices.
3. Used in fumigation to remove sources of infection from a house or room.
4. Used to make calcium hydrogen sulphite, $Ca(HSO_3)_2$ that is used to bleach wood pulp in the manufacture of paper.
5. Liquid sulphur dioxide is used in refrigerators as a refrigerant.
6. It is used for bleaching wool, straw, silk and sponges. These cannot be bleached using chlorine as they react with it.

Self-evaluation Test 3.2

1. Write balanced chemical equations that show the two methods of preparing sulphur dioxide in the laboratory.
2. Blue litmus paper placed in a solution of sulphur dioxide turns red and then colourless. Explain.
3. Describe an experiment you can carry out in the laboratory to establish that a gas is sulphur dioxide.

3.4 Industrial preparation of sulphuric acid

You already know sulphuric acid as an important reagent in the Chemistry laboratory. Do you know how large amounts of this important acid are prepared?

Research activity

In pairs do the following:

1. Research on the method used for industrial preparation of sulphuric acid.
2. Research on the processes involved in the preparation of the acid by this method under the following subheadings:
    (a) The raw materials required.
    (b) Production of sulphur dioxide.
    (c) Conversion of sulphur dioxide to sulphur trioxide.
    (d) Production of oleum.
    (e) Dilution of oleum to form sulphuric acid.
3. Compile a report and present it in class.

The facts

The amount of sulphuric acid consumed by a country in tonnes can be used as a measure of the level of industrialisation of that country. This is so because sulphuric acid is a very important industrial chemical. The industrial process of making sulphuric acid in large scale is known as the **contact process**. The major raw materials for the contact process are sulphur or sulphide ore, air and water.

The process takes place in the following stages:

Stage I: Production of sulphur dioxide

Sulphur dioxide is produced by either burning sulphur or by burning sulphur containing ores such as iron pyrites, (FeS_2), zinc blende, (ZnS) or galena, (PbS) Sulphur dioxide may also be produced by burning hydrogen sulphide from crude oil. The following equations show methods of obtaining sulphur dioxide.

a) Burning sulphur from Frasch process in air.

$$S(s) + O_2\ (g) \longrightarrow SO_2\ (g)$$

b) Roasting sulphide ores in air during the extraction of metals.

$$2ZnS(s) + 3O_2\ (g) \longrightarrow 2ZnO(s) + 2SO_2\ (g)$$
$$4FeS_2(s) + 11O_2\ (g) \longrightarrow 2Fe_2O_3\ (s) + 8SO_2\ (g)$$
$$2PbS(s) + 3O_2\ (g) \longrightarrow 2PbO(s) + 2SO_2\ (g)$$

c) Burning hydrogen sulphide from crude oil in air.

$$2H_2S\ (g) + 3O_2\ (g) \longrightarrow 2SO_2\ (g) + 2H_2O(g)$$

Stage II: Oxidation of sulphur dioxide to sulphur trioxide

The sulphur dioxide produced in stage 1 is mixed with air and passed through a dust precipitator (purification chamber). Here dust particles and other impurities are removed then the gases dried. This is achieved through electrostatic precipitation. Purification is important to remove impurities that could otherwise render the catalyst useless (poison the catalyst). The purified gases are then heated to a temperature of (450 – 500)°C and compressed over heated Vanadium (V) oxide catalyst. Under these conditions, sulphur dioxide is oxidised to sulphur trioxide. The equation for the reaction is:

$$2SO_2(g) + O_2(g) \xrightleftharpoons[450^\circ C - 500^\circ C]{V_2O_5} 2SO_3(g) + \text{heat}$$

This reaction is exothermic and the temperature rises to about 600°C. Under these optimum conditions, there is 98% yield of sulphur trioxide. The Vanadium (V) oxide catalyst is spread on silica gel to increase the surface area for the reaction. Platinum catalyst can be used in place of Vanadium (V) oxide but it has a disadvantage of being easily susceptible to poisoning and is also expensive.

The gases coming from the catalyst chamber are at high temperatures due to the exothermic nature of the reaction. They are therefore taken back to the heat exchanger for cooling as they pre-heat the incoming mixture of sulphur dioxide and air. Once the reaction has begun, the heat released is used to maintain the operating temperatures.

Any sulphur dioxide that is not converted to sulphur dioxide should not be allowed to escape into the atmosphere as this would cause pollution. It is therefore reacted with calcium hydroxide in the chimneys forming salt and water.

Calcium hydroxide + sulphur dioxide ⟶ calcium sulphite + water

$$Ca\,(OH)_2(s) + SO_2(g) \longrightarrow CaSO_3(s) + H_2O(l)$$

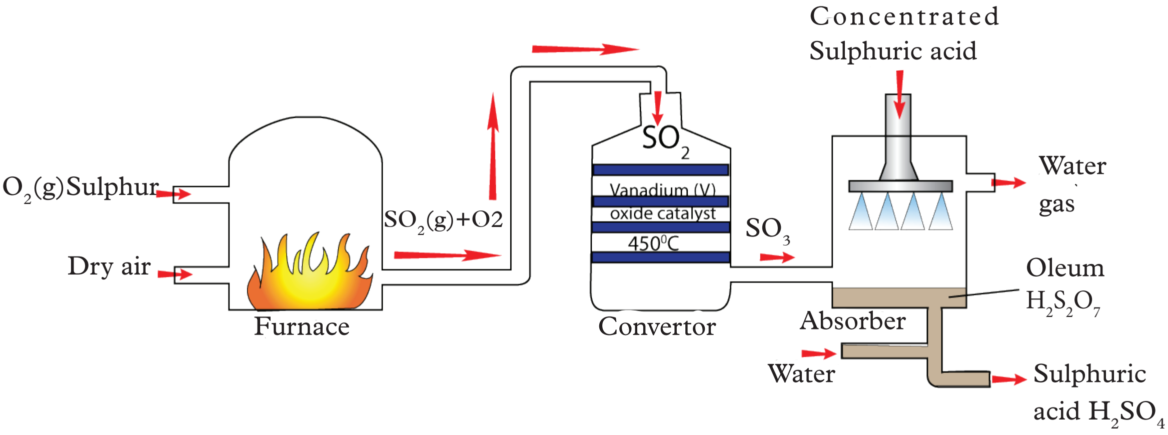


Fig 3.8: The Contact process

Stage III: Production of oleum

The cooled gases are passed to the absorption tower, packed with ceramic rings through which concentrated sulphuric acid flows. The sulphur trioxide reacts with concentrated sulphuric acid forming oleum (fuming sulphuric acid). Oleum is an oily liquid.

$$H_2SO_4\,(l) + SO_3(g) \longrightarrow \underset{\text{Oleum}}{H_2S_2O_7(l)}$$

Sulphur trioxide is not dissolved directly in water to form sulphuric acid. This is because if dissolved in water directly, the reaction will produce a lot of heat as it is highly exothermic. The heat produced boils the acid forming a mist of tiny droplets acid sprays which are dangerous and condense slowly.

Stage IV: Dilution of Oleum to form sulphuric acid

Oleum produced in stage III is mixed with a carefully calculated quantity of water to give concentrated sulphuric acid. The acid obtained from this process is about 98% pure.

$$H_2S_2O_7(l) + H_2O\ (l) \longrightarrow 2H_2SO_4(l)$$

Self-evaluation Test 3.3

1. (a) Name the process that is used to prepare sulphuric acid on large scale.
   (b) List the raw materials used in the process named 1(a) above.
   (c) Using equations show how sulphur dioxide is obtained for use in the process in 1(a) above.
2. (a) Outline the steps taken during the industrial preparation of sulphuric acid.
   (b) What is the estimate percentage of sulphuric acid formed by Contact process?
3. Why is sulphur trioxide not passed directly through through water to sulphuric acid but passed through concentrated sulphuric acid during its preparation?

3.5 Properties of concentrated and dilute sulphuric acid

(a) *Properties of concentrated sulphuric acid*

Activity 3.4

Apparatus and reagents

Concentrated sulphuric acid, beakers, 10 cm^3 measuring cylinder, sugar crystals, hydrated copper (II) sulphate crystals, distilled water, boiling tubes, pieces of cloth, piece of paper, piece of wood, petri dishes, stirrer, test tube holder, red and blue litmus papers, felt pen, copper turnings, magnesium ribbon.

Procedure

1. Observe the colour and smell of concentrated sulphuric acid. Drop blue and red litmus paper in it.
   - Note down your observations.
2. Pour some water in a boiling tube. Holding the tube with a test tube holder in a slanting position, slowly add a little concentrated sulphuric acid by sliding it along the side of the tube and shake gently. Feel the boiling tube with your palm.
   - Note down your observations.

3. Place concentrated sulphuric acid into a beaker and mark its level with a felt pen. Place the beaker in an open but secure place and observe after 48 hours.
   - Note down your observations.
4. Put 3 spatula endful of sugar crystals in 100 ml glass beaker. Add about 5 cm^3 of concentrated sulphuric acid into the beaker. Observe the changes in the beaker for about 10 minutes.
   - Note down your observations.
5. Put 3 spatula endful of hydrated copper (II) sulphate crystals in a 50 ml glass beaker. To it add 5 cm^3 of concentrated sulphuric acid.
   - Note down your observations.
6. a) Place half spatula endful of copper turnings in a test tube and add 3 cm^3 of concentrated sulphuric acid. Heat the test tube gently and test for any gas evolved by placing filter paper soaked in acidified potassium dichromate (VI) at the mouth of the test tube.
   - Note down your observations.

   b) Repeat procedure (6a) using magnesium metal in place of copper turnings. Note down your observations.

   c) Pour concentrated sulphuric acid onto a piece of cloth, piece of paper and piece of wood. Record your observations.

Study question

Write balanced chemical equations for the reactions that took place in procedures 4, 5 and 6 (a) and (b).

Discussion corner!

1. In groups, discuss the observations made in the experiment and the study questions above.
2. Write a report on your findings and present it in class.

I have discovered that…

Concentrated sulphuric acid converts sugar crystals to yellow, brown and finally to a black mass of substance. Blue copper (II) sulphate crystals in the acid turns into a white powder. On the other hand, brown copper turnings disappear in the acid forming a blue solution. When a piece of paper and a piece of cloth are placed in the acid, they get charred and torn into small pieces.

When concentrated sulphuric acid is placed in an open place for sometime its level in the container rises up.

The facts

Physical properties of sulphuric acid

1. It is a colourless, odourless oily liquid.
2. It has a density of 1.84 g/cm^3.
3. It dissolves in water with evolution of large amounts of heat, that is, why when added to water in the boiling tube, the boiling tube becomes hot.
4. It freezes at 10°C and boils at 338°C.

Note: When diluting the acid, always add the acid into water and never water into the acid and stir continuously. If water is added to the acid, a lot of heat is produced causing the solution to boil very violently, splashing the acid out of the container.

Chemical properties of concentrated sulphuric acid

i) It is hygroscopic

Concentrated sulphuric acid absorbs water from the atmosphere. That is why the level of water in the beaker in procedure (3) of activity 3.4 rose after 48 hours. This property makes concentrated sulphuric acid a suitable drying agent for gases which do not react with it.

ii) It is a dehydrating agent

It is capable of removing combined water from substances. This explains the changes that occurred to sugar crystals in procedure (4) of activity 3.4. The black mass formed is carbon. The equation for the reaction that took place is:

$$C_{12}H_{22}O_{11}(s) \longrightarrow 12C(s) + 11\,H_2O(g)$$

Concentrated sulphuric acid as well removes water from hydrated copper (II) sulphate converting it to a white powder of anhydrous copper (II) sulphate.

$$\underset{\text{Blue crystals}}{CuSO_4\,.5H_2O(s)} \xrightarrow{\text{conc.}H_2SO_4} \underset{\text{White powder}}{CuSO_4(s)} + 5H_2O\ (l)$$

Similarly concentrated sulphuric acid removes water from wood converting it into a black mass of carbon.

Note: Dehydration is the removal of chemically combined water or elements of water from a compound. A substance capable of removing chemically combined water or elements of water is called a **dehydrating agent**.

iii) It is corrosive

Concentrated sulphuric acid should not come into contact with skin or cloth as it will corrode them. The same happens to paper when put into the acid.

iv) It is an oxidising agent

Hot concentrated sulphuric acid acts as an oxidising agent. It oxidises almost all metals with the exception of those that are very low in the reactivity series such as gold and platinum. The following equations show the effect of concentrated sulphuric acid with metals.

i $Cu(s) + 2H_2SO_4(l) \longrightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$
(Brown) (blue)
$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

ii. $Mg(s) + 2H_2SO_4(l) \longrightarrow MgSO_4(aq) + SO_2(g) + 2H_2O(l)$
$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$

iii. $Zn(s) + 2H_2SO_4(l) \longrightarrow ZnSO_4(aq) + SO_2(g) + 2H_2O(l)$
$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

As seen from the above equations, the metals lose two electrons to form corresponding metal ions. This is oxidation. Concentrated sulphuric acid accepts the electrons lost and reduced into sulphur dioxide and water. This is reduction.

$$2H_2SO_4(l) + 2e^- \longrightarrow SO_4^{2-}(aq) + SO_2(g) + 2H_2O(l)$$

Note: Oxidation is the loss of electrons while reduction is the gain of electrons

(b)Properties of dilute sulphuric acid

Activity 3.5

Apparatus and reagents

Dilute sulphuric acid, test tubes, magnesium ribbon, litmus paper (blue and red), splints, phenolphthalein indicator, sodium hydroxide solution, copper (II) oxide, calcium hydroxide solution, sodium carbonate and sodium hydrogen carbonate.

Procedure

1. Observe the colour of dilute sulphuric acid.
2. Drop blue and red litmus paper in dilute sulphuric acid.
    - Record your observation.
3. Place 3 cm³ of dilute sulphuric acid in a test tube, add magnesium ribbon and test any gas produced using a burning splint.
    - Record your observation.
4. Place 3 cm³ of dilute sulphuric acid in a test tube. Add two drops of phenolphthalein indicator. Add sodium hydroxide solution drop wise until the mixture just changes colour.
    - Record your observations.

5. To 3 cm^3 of dilute sulphuric acid in a test tube, add some copper (II) oxide powder little at a time as you stir. Continue until no more copper (II) oxide dissolves.
   - Record your observations.
6. Put 3 cm^3 of dilute sulphuric acid into a test tube. Add sodium carbonate and connect it to another test tube containing calcium hydroxide solution as shown below.

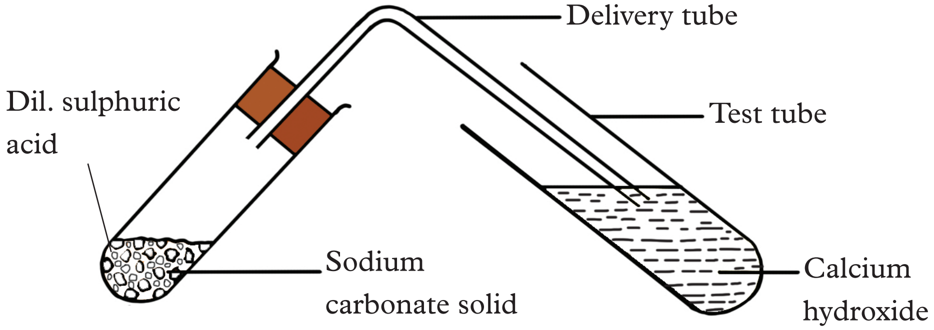


Fig 3.9: Reacting dilute sulphuric acid with sodium carbonate

7. Repeat the procedure using sodium hydrogen carbonate.
   - Note down your observations.

Study questions

1. Write balanced chemical equations for the reactions taking place in procedures 4, 5 and 6.
2. What was the role of the following:
   i. Phenolphthalein indicator in procedure 4?
   ii. Burning splint in procedure 3?
   iii. Calcium hydroxide in procedure 6?

Discussion corner!

1. In your study groups, discuss the results of the experiment and the study question above.
2. Prepare a report and present it to the class.

I have discovered that…

Blue litmus paper in dilute sulphuric acid turns red. Phenolphthalein indicator in a dilute sulphuric acid turns pink. Copper (II) oxide reacts with the acid forming a blue solution. When either sodium carbonate or sodium hydrogen carbonate reacts with dilute sulphuric acid, effervescence is seen due to production of a gas. The gas forms a white precipitate with lime water.

The facts

Physical properties of dilute sulphuric acid

- It is colourless and odourless.

Chemical properties of dilute sulphuric acid

i. Reaction with metals

Dilute sulphuric acid reacts with metals to form salt and hydrogen gas. However, metals like copper, silver and gold do not react with dilute sulphuric acid. They cannot displace hydrogen from the acid. Reaction of dilute sulphuric acid with some metals are shown below.

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

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

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

ii. Reaction with oxides and hydroxides

Dilute sulphuric acid reacts with hydroxides to form salt and water. This is a neutralisation reaction. Phenolphthalein indicator turns pink in a neutral solution. This is shown in the following reactions.

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

$H_2SO_4(aq) + 2KOH(aq) \longrightarrow K_2SO_4(aq) + H_2O(l)$

Dilute sulphuric acid also neutralises metal oxides to form salt and water. It reacts with copper (II) oxide forming a blue solution of copper (II) sulphate as shown in the following equation.

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

iii. Reaction with carbonates and hydrogen carbonates

Dilute sulphuric acid reacts with metal carbonates and hydrogen carbonates to form salt, water and carbon dioxide. The carbon dioxide produced forms a white precipitate with lime water. The following equations show the reactions that take place when the acid reacts with sodium carbonate and sodium hydrogen carbonate respectively.

$Na_2CO_3(s) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$

$2NaHCO_3(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$

Note: When the carbonates of lead, calcium and barium are added to dilute sulphuric acid, the reaction starts soon stops. This is because the sulphates formed are insoluble; they coat the rest of the metal carbonate preventing further contact between it and the acid hence the reaction stops.

Uses of sulphuric acid

Research activity

1. Carry out a research on the uses of sulphuric acid based on the following:
    - Manufacture of fertilisers.
    - Manufacture of detergents.
    - Manufacture of paints and dyes.
    - Manufacture of plastics, explosives and drugs.
    - As an electrolyte.
    - Processing of metal ores during extraction of metals.
    - Pickling iron and steel.
    - Production of other chemicals such as hydrochloric acid, hydrofluoric acid and metal sulphates.
    - Processing of natural fibres and manufacture of synthetic fibres.
2. Compile a report and present it to your teacher for evaluation.

Self-evaluation Test 3.4

1. Concentrated sulphuric acid is a dehydrating agent. Explain.
2. A student reacted dilute sulphuric acid with calcium carbonate. Explain the observation made. Use chemical equation in your explanation.
3. In the preparation of sulphuric acid by Contact process, sulphur trioxide is not directly dissolved in sulphuric acid. Explain why this is so and use equations to show how the dissolution occurs.
4. Give one industrial application of sulphuric acid.

Laboratory preparation of sulphates

Activity 3.6

Apparatus and reagents

Measuring cylinder, 100ml glass beaker, litmus papers, 2M sodium hydroxide, Bunsen burner, concentrated sulphuric acid, dilute sulphuric acid, copper turnings, sodium sulphate, lead (II) nitrate and calcium nitrate.

Procedure

1. Put 25 cm^3 of 2M sodium hydroxide into a beaker and drop in a piece of litmus paper. Add 25 cm^3 of 1M dilute sulphuric acid until the litmus paper just changes colour. Remove the litmus paper and evaporate to concentrate the

solution. Allow to cool and crystallise the salt in the usual way.

2. Put copper turnings into a beaker and add hot concentrated sulphuric acid.

   Caution: Do this experiment in a fume chamber.

   When the reaction is over (no more effervescence) filter the solution and heat the filtrate to concentrate it. Allow the solution to cool slowly and crystallise.

3. i) Mix a solution of sodium sulphate with a solution of lead (II) nitrate.
      - Record your observation.

   ii) Filter the resulting mixture and collect the residue on the filter paper. Wash the residue a few times with distilled water and dry the residue between filter papers.

   iii) Repeat procedure 3 (i) using calcium nitrate in place of lead (II) nitrate.

Study questions

1. Why should procedure 2 be carried out in a fume chamber?
2. Write chemical equations for the reactions that take place in the three procedures.

Discussion corner!

1. Discuss in groups the preparation of sulphates using the methods illustrated in this experiment and the study questions above.
2. Present your findings to the teacher for evaluation.

The facts

There are three ways of preparing sulphates in the laboratory.

1. By neutralising dilute sulphuric acid with an alkali.

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

   The solution is evaporated to concentrate it and when it slowly cools, sodium sulphate crystals form. Ammonium sulphate and aluminium sulphate can also be prepared using a similar reaction method.

2. By the action of concentrated or dilute sulphuric acid on an appropriate metal as shown by the equations below.

   $Cu(s) + 2H_2SO_4(l) \longrightarrow CuSO_4(aq) + 2SO_2(g) + 2H_2O(l)$

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

   Filtration is done and the filtrate evaporated then cooled slowly to form crystals of the salt.

3. By double decomposition

   This method is suitable for preparing insoluble sulphates. The reaction equation below show how this happens.

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

$$Na_2SO_4(aq) + Ca(NO_3)_2(aq) \longrightarrow 2NaNO_3(aq) + CaSO_4(s)$$

After filtration, the residue is washed with distilled water and dried.

3.6 Test for sulphates and sulphites

Activity 3.7

Apparatus and reagents

Test tubes, sodium sulphite and sodium sulphate solution, barium nitrate, dilute nitric acid.

Procedure

1. Place 2 cm^3 each of sodium sulphate and sodium sulphite in separate test tubes.
2. To each of the solutions in the test tubes, add four drops of barium nitrate solution and record your observations.
3. To each of the test tubes in procedure 2, add 2 cm^3 of dilute nitric acid and shake.
    - Record your observations in the following table.

Table 3.3 Results for test for sulphite and sulphate ions

Ion in solution	**Addition of four drops of barium nitrate**	**Addition of 2 cm^3 of dilute nitric acid**
Sulphite ion (SO_3^{2-})		
Sulphate ion (SO_4^{2-})		

Study question

Write ionic equations for the reactions that took place.

I have discovered that…

When a few drops of barium nitrate is added to a solution containing sulphite or sulphate ions, a white precipitate forms. On addition of dilute nitric acid, the precipitate formed with sulphite ions dissolves while that formed with sulphate ions persist.

The facts

Solutions containing sulphite and sulphate ions form white precipitate when barium nitrate or barium chloride solutions are added.

$$Ba^{2+}(aq) + SO_3^{2-}(aq) \longrightarrow \underset{\text{White precipitate}}{BaSO_3(s)}$$

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow \underset{\text{White precipitate}}{BaSO_4(s)}$$

To differentiate between sulphite and sulphate ions, dilute nitric acid or dilute hydrochloric acid is added to the test tube containing the white precipitate. If the salt is a sulphate, the white precipitate persists. However, if the salt is a sulphite the precipitate dissolves giving off sulphur dioxide and a colourless solution is formed.

$$BaSO_3(s) + 2HNO_3(aq) \longrightarrow Ba(NO_3)_2(aq) + SO_2(g) + H_2O(l)$$

$$BaSO_3(s) + 2HCl(aq) \longrightarrow BaCl_2(aq) + SO_2(g) + H_2O(l)$$

The identity of the sulphur dioxide gas may be confirmed by testing with a paper dipped in acidified potassium dichromate (VI) solution which changes colour from orange to green.

3.7 Environmental issues related to sulphur oxides

Research activity

1. Research on the environmental impact of sulphur oxides. Base your research on the following:
    - Sources of the sulphur oxides.
    - How the oxides gets into the environment.
    - The effect of the oxides on:
        - i. Plants and buildings.
        - ii. Soil acidity and absorption of minerals by plants.
        - iii. Aquatic life.
2. Use your findings to compile a report and present it to the teacher for evaluation.

Steps to reduce pollution by sulphur oxides

1. Unreacted sulphur dioxide and sulphur trioxide are recycled into the catalyst chamber in the Contact process.
2. Unreacted sulphur dioxide during Contact process undergoes scrubbing in the chimney.
3. Always use sulphur free fossil fuels.
4. Abide by the laws set up by the government to reduce the amount of sulphur oxides released to the atmosphere.

My environment, my life!

Always remember to protect the environment while undertaking any activity whether commercial or not. Preserve the environment for future generation.

Self-evaluation Test 3.5

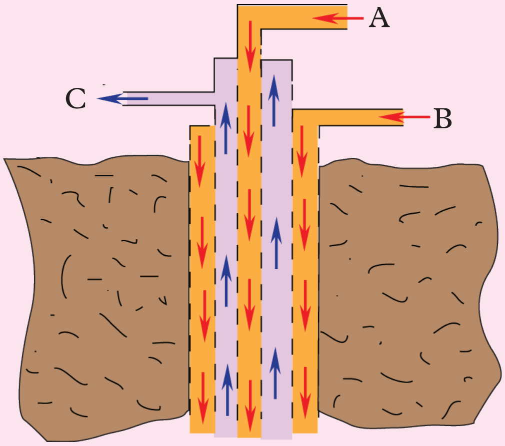
1. Explain ways in which sulphur compounds released to the atmosphere affect plants and aquatic animals.
2. Why is it important to protect our environment from pollution?

Unit Summary 3

- Sulphur occurs naturally as a free element or in combined form as hydrogen sulphide, metal sulphide or metal sulphates.
- Extraction of sulphur is done by the Frasch process. Claus process can also be used.
- Sulphur does not dissolve in water but dissolves in organic solvents like carbon disulphide.
- In the laboratory, sulphur dioxide is prepared by the action of dilute hydrochloric acid on sodium sulphite or heating a mixture of copper turnings and concentrated sulphuric acid.
- Sulphur dioxide is a colourless, poisonous, acidic gas with a choking smell. It is very soluble in water and denser than air.
- Sulphuric acid is industrially prepared by the Contact process. The main raw materials for this process are sulphur or an ore of sulphur, water and air.
- Concentrated sulphuric acid has oxidizing, dehydrating, corrosive and hygroscopic properties. Dilute sulphuric acid on the other hand has all the acidic properties of a strong mineral acid.
- Sulphites and sulphates are identified by reacting them with barium nitrate or barium chloride solution. A white precipitate of barium sulphite dissolves in dilute hydrochloric acid or dilute nitric acid but the barium sulphate precipitate persists in either of the dilute acids.
- Oxides of sulphur should not be allowed into the atmosphere as they are a major cause of environmental pollution.

Test your Competence 3

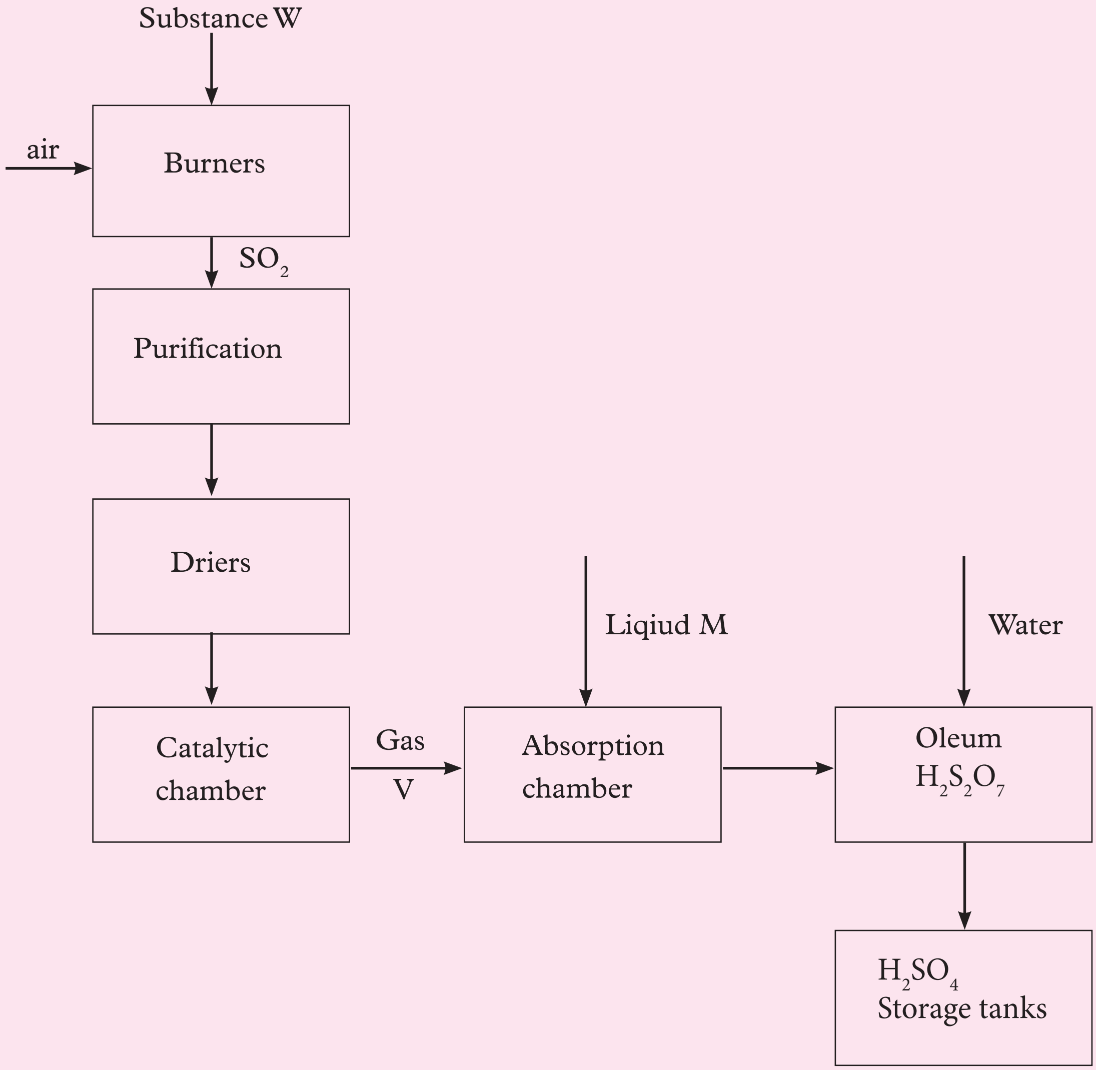
1. Extraction of sulphur is carried out by the __________ process while industrial preparation of sulphuric acid is done by the __________ process.
2. Kampire and Gasimba wanted to prepare dilute sulphuric acid using concentrated sulphuric acid and distilled water.
   (a) What procedure should they follow?
   (b) Why should they follow the procedure you have mentioned in 1(a) above?
3. The figure below shows the process used to extract sulphur.



   (a) Name the substance that passes through pipes A, B and C.
   (b) What is the role of the substances that pass though pipes A and B?
4. Sulphuric acid can be used for all of the following except ______________.
   A. Manufacture of fertiliser
   B. Making car tyres
   C. Dehydrating sugar
   D. Used in car batteries
5. Write **true** or **false** for each of the following statements.
   (a) Dilute sulphuric acid is a dehydrating agent. ______________
   (b) Metals react with concentrated sulphuric acid to form salts and hydrogen gas. ______________
   (c) Sulphur dioxide bleaches by reduction. ______________
   (d) Sulphur dioxide is an alkaline gas ______________
6. __________ is the catalyst used in the contact process.
7. Powdered sulphur was placed in deflagrating spoon and heated in a Bunsen burner flame until the sulphur started to burn.
   (a) What is observed as the sulphur burns?
   (b) Write an equation for the reaction in 7(a) above.

(c) What problem does this reaction pose to the environment?

8. Study the flowchart below showing manufacture process of sulphuric acid.



(a) Identify (i) Substance W

(ii) Gas V

(iii) Liquid M

(b) State the conditions necessary for the formation of Gas V.

(c) How would you control pollution that may come as a result of the above process?

9. You have been requested to give a speech on the environmental impact of sulphur compounds and ways of protecting the environment from pollution by sulphur compounds. Prepare the speech you would give.

10. In an experiment, 2.5g of sulphur was obtained by reacting hydrogen sulphide and chlorine as shown by the equation below.

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

   (a) Which of the reactants acts as a reducing agent in the above reaction.
   (b) State the effects of hydrogen sulphide on the environment.

11. Powdered sulphur was placed in deflagrating spoon and heated in a Bunsen flame until the sulphur started to burn.
   (a) What was seen when the sulphur burns.
   (b) Write an equation for the reaction in 11(a) above.
   (c) When the product formed in 11(b) is dissolved in water, name the solution that is formed and write the equation for the reaction.
   (d) If a few drops of litmus solution was added to the solution in 11(c) above, what would be the colour of the litmus solution?
   (e) Write an equation for the reaction that will occur between the solution formed in 11(c) above and sodium hydroxide solution.

12. (a) Suggest ways through which air pollution can be reduced.
   (b) To what extent is acid rain a problem in Rwanda?
   (c) Acid rain is an international problem. Justify the statement.
   (d) Should governments of different countries be doing something about the problem? Discuss.

UNIT 4

Chlorine and its inorganic compounds

Key unit competency

To be able to relate the properties of chlorine and its compounds to its use, describe how compounds of chlorine are prepared and discuss the related environmental issues.

Learning objectives

By the end of this unit, I should be able to:

- Describe the physical and chemical properties of chlorine and its compounds.
- Recall how chlorine is prepared in the laboratory, observe the colours formed during the preparation and be able to test for the presence of chlorine.
- Prepare hydrogen chloride in the laboratory.
- Test for the presence of chlorides and other halides in solution.
- Develop a culture of working in groups.
- Develop self-confidence in the presentation of research work and group discussions.
- Protect natural resources and appreciate the impacts of CFCs in the environment.
- Show respect for procedures in carrying out experiments.

Mind teaser

Chlorine like other elements have a variety of uses. Look at the photographs below.

A

B

Chlorine solution

D

C

Fig 4.1

Each of the items in Fig 4.1 represent a use of chlorine. Can you identify the use? How is it important in our lives? In which other ways is chlorine and its compounds important to us?

Introduction

You already know that chlorine is a gas found in period 3 and group VII of the Periodic Table. Can you recall the properties of chlorine and other halogens? What are the uses of chlorine? What compounds of chlorine do you know?

Chlorine is a green – yellow gas. The name chlorine was derived from the Greek word, 'chloros' which means green. The family of elements to which chlorine belongs to in the Periodic Table is called **halogens**.

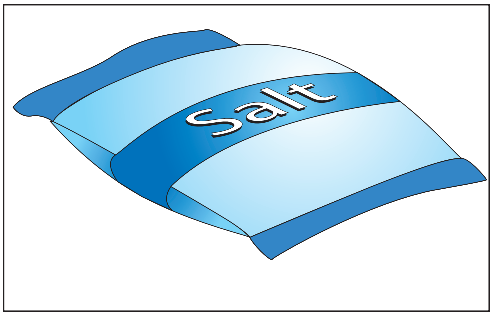


Fig 4.2: Packet of table salt

In nature, chlorine mostly occurs in a combined form with other elements. It is usually found combined with sodium forming sodium chloride (**rock salt**). Rock salt is the source of table salt that we use at home. Sea water also contains sodium chloride. Other naturally occurring compounds of chlorine are potassium chloride, magnesium chloride and calcium chloride. Chlorine is a highly reactive element hence it is not found in a free state. Compounds of chlorine together with chlorine itself have a variety of uses as you shall learn later.

4.1 Preparation and test for chlorine

Activity 4.1

Apparatus and reagents

Round bottomed flask, conical flask, gas jars, manganese dioxide, concentrated hydrochloric acid source of heat, two wash bottles, delivery tubes, dropping funnel, retort stand, blue and red litmus papers.

Caution: Chlorine gas is poisonous and should be prepared in a fume chamber or in open air.

Procedure

1. Put manganese dioxide in a round bottomed flask and arrange the apparatus as shown in the following set up.

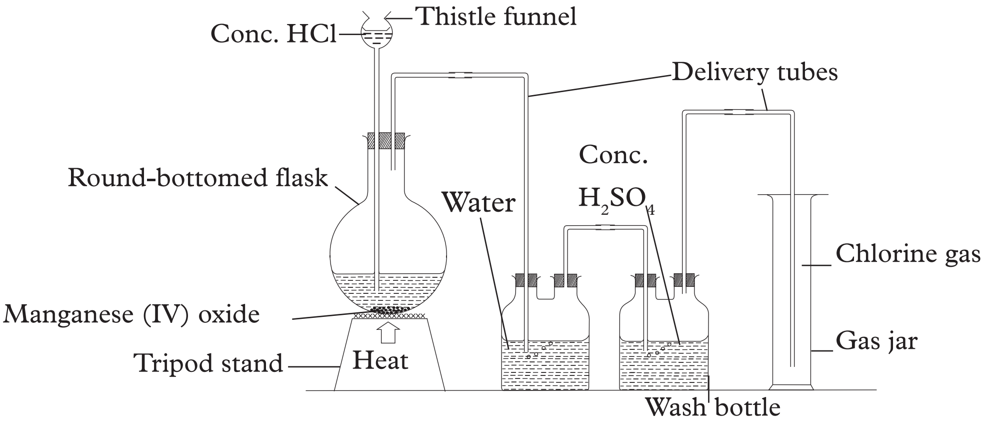


Fig 4.3: Preparation of chlorine gas in the laboratory

2. Add concentrated hydrochloric acid into the flask as shown and heat the mixture.
   - Note down your observations in the gas jar and in the flask.
3. Test the gas collected with moist blue and red litmus papers.
   - Note down your observations.
4. Smell the gas by wafting it towards your nose.
   - Describe the smell of the gas.

Study questions

1. What do you observe as soon as concentrated hydrochloric acid is added to manganese dioxide?
2. Write an equation for the reactions that take place.
3. What is the function of water and concentrated sulphuric acid in the two wash bottles?

Discussion corner!

1. Discuss the results of the experiment in pairs and the study questions above.
2. Write a report and present it to the teacher.

I have discovered that…

When dilute hydrochloric acid comes into contact with manganese dioxide, there is effervescence. A green-yellow gas collects in the gas jar. The gas turns wet blue litmus paper red and later bleaches it.

The facts

Hydrochloric acid reacts with manganese (IV) oxide to form manganese (II) chloride, water and chlorine gas.

$$MnO_2\,(s) + 4HCl\,(l) \longrightarrow MnCl_2\,(aq) + 2H_2O\,(l) + Cl_2\,(g)$$

Water in the wash bottle absorbs hydrogen chloride fumes. The gas evolved is dried by passing it through concentrated sulphuric acid. Chlorine gas can also be dried by passing it through anhydrous calcium chloride. It is collected by downward delivery since it is denser than air.

Chlorine gas can also be prepared using potassium permanganate instead of manganese dioxide as shown in the diagram below.

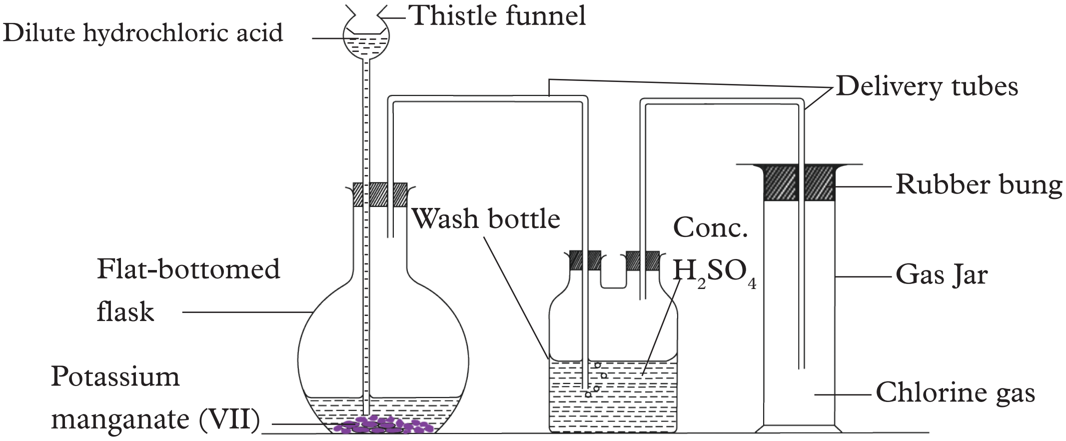


Fig 4.4: Preparation of chlorine using potassium permanganate and concentrated hydrochloric acid

In this method, no heating is required since potassium permanganate is a stronger oxidizing agent than manganese dioxide. The equation for the reaction that takes place is as follows.

$$2KMnO_4(s) + 16HCl\,(aq) \longrightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$$

Chlorine also can be prepared by heating a mixture of manganese (IV) oxide and sodium chloride with concentrated sulphuric acid. In this reaction, sulphuric acid first reacts with sodium chloride to form hydrochloric acid. The hydrochloric acid then reacts with manganese (IV) oxide to form chlorine gas.

$$H_2SO_4\,(l) + NaCl(s) \longrightarrow HCl(g) + NaHSO_4(aq)$$

$$MnO_2(s) + 4HCl\,(aq) \longrightarrow MnCl_2\,(aq) + 2H_2O\,(l) + Cl_2\,(g)$$

4.2 Physical properties of chlorine

1. It is a green – yellow gas with a pungent choking smell.
2. Boils at 34.05C and freezes at –101C.
3. It is 2.5 times denser than air hence collected by downward delivery.
4. It is fairly soluble in water (in aqueous solution it is called chlorine water).
5. It is easily liquefied under pressure.

Test for chlorine gas

Moist blue litmus paper placed in chlorine gas immediately turns red and then bleached.

Self-evaluation Test 4.1

1. Uwibambazi and members of her study group wanted to prepare chlorine gas in the laboratory.
    (a) Describe the method you would advise them to use.
    (b) What precautions would you advise them to take?
2. Gatete found a green-yellow gas in a gas jar. He wanted to confirm whether the gas was chlorine. Describe a method you would advise him to use.
3. What is the name of the group that contains the element chlorine in the periodic table?
4. Which of the following statement does not describe physical property of chlorine?
    A. The density of chlorine gas at standard temperature and pressure is 3.17g/l.
    B. It combines with sodium to form table salt.
    C. It is a yellow- green gas.
    D. It freezes at -101C.
5. How many electrons are found in the outermost energy level of chlorine atom.
6. Chlorine gas is prepared in the laboratory through the reaction between concentrated hydrochloric acid and manganese dioxide. Write a balanced chemical equation for the reaction that takes place.
7. Which of the following best describes chlorine?
    (a) A solid that readily combines with other elements.
    (b) A halogen.
    (c) A noble gas.
    (d) A mixture of carbon dioxide and ammonia.

4.3 Chemical properties of chlorine

a) Reaction of chlorine with metals

Activity 4.2

Apparatus and reagents

Wooden splints, glass rod, magnesium ribbon, iron wire, hydrogen, water, sodium hydroxide, gas jar, chlorine, source of heat, deflagrating spoon, conical flask, combustion tube.

Procedure

1. Wrap cleaned magnesium ribbon on a deflagrating spoon and heat it until it starts burning.

2. Lower the burning magnesium in a gas jar of chlorine.
   - Note down your observation.

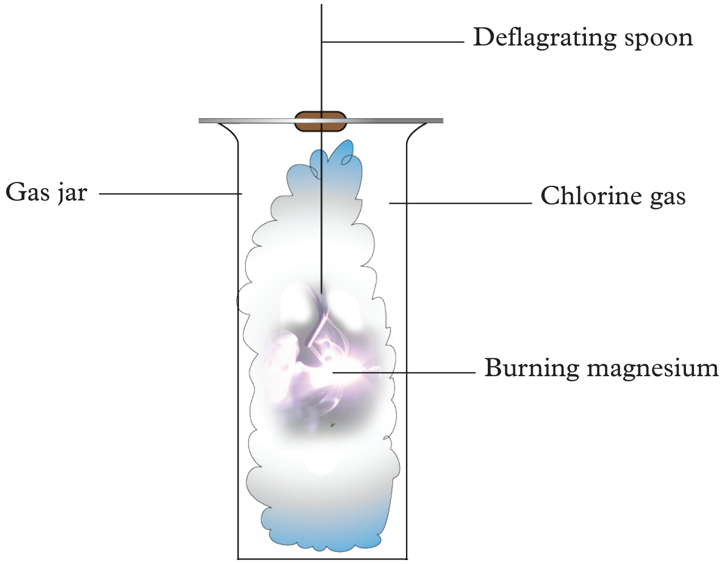


Fig 4.5: Reaction between magnesium and chlorine gas

3. Place a piece of iron wire in a combustion tube and arrange the apparatus as follows.

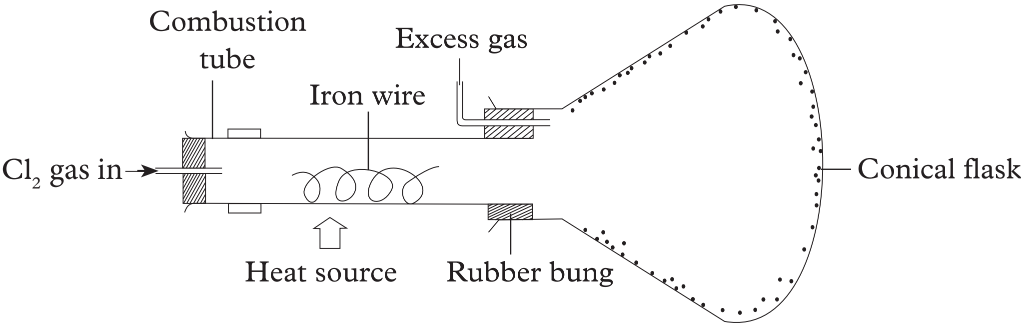


Fig 4.6: Reaction of chlorine with iron

4. Heat the iron wire until it glows red-hot.
5. Pass dry chlorine gas through the combustion tube over the red-hot iron wire.
6. Warm the end of the combustion tube near the conical flask.

Study questions

1. Write a chemical equation for the reaction that takes place when chlorine reacts with magnesium.
2. What observation did you make on the combustion tube when chlorine reacted with iron?

3. Write a chemical equation for the reaction between chlorine and iron.
4. Why is it possible to collect the product from the combustion tube into the conical flask as shown in fig 4.6?

I have discovered that…

Burning magnesium continues to burn in chlorine with a bright white dazzling flame. White fumes are formed that fill the gas jar. When chlorine gas is passed over heated red – hot iron wire, the wire starts to glow and the glowing continues for a long time. A black substance is then formed on the sides of the combustion tube. When the end of the combustion tube near the receiver is warmed, the black crystals sublime and collect in the receiver.

The facts

Chlorine directly combines with metals forming corresponding chloride salts. When reacted with magnesium, a white solid of magnesium chloride is formed according the following equation.

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

Chlorine reacts with iron to form black crystals of iron (III) chloride. When heated, this salt sublimes and condenses on the cooler sides of the combustion tube.

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

This reaction is exothermic and so the glow continues for a long time.

When exposed to moist air, iron (III) chloride slightly fumes. This is because iron (III) chloride reacts with water vapour in air to form white fumes of hydrogen chloride gas.

$2FeCl_3(s) + 6H_2O(g) \longrightarrow 2Fe(OH)_3(s) + 6HCl(g)$

Therefore, during the preparation of iron (III) chloride, the whole apparatus should be dry.

Note: Where a metal forms two chlorides, the chloride of the higher oxidation number is formed because it is more stable. This explains why iron (II) chloride is not formed in this case.

Similarly, sodium metal burns spontaneously in chlorine to form white solid of sodium chloride.

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

b) Reaction of chlorine with hydrogen, ammonia solution and water

Activity 4.3

Apparatus and reagents

Wooden splints, glass rod, concentrated ammonium solution, water, beaker, matchbox, source of hydrogen gas, chlorine gas, gas jar.

Procedure

1. Introduce a burning splint into a gas jar of chlorine.
   - Note down your observation
2. a) Introduce burning hydrogen from a generator into a gas jar of dry chlorine gas as shown in fig 4.7.
   - Note down your observation.

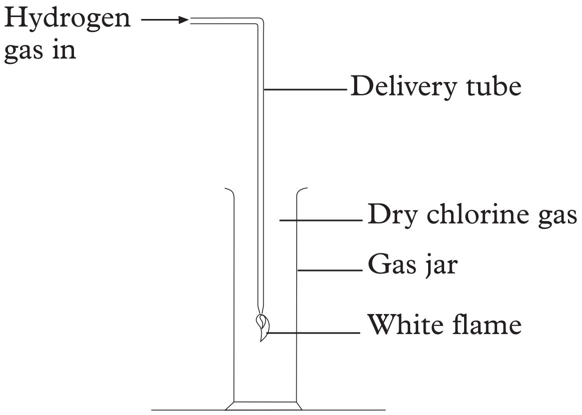


Fig. 4.7: Experimental set up for the reaction between chlorine and hydrogen

   b) Dip a glass rod into concentrated ammonia solution and then insert it into the gas jar of chlorine.
   - Record your observations.
3. a) Pass chlorine gas through distilled water in a gas jar as shown fig 4.8.
   - Record your observations.

   b) Fill a boiling tube with chlorine water. Invert the boiling tube in a beaker containing distilled water as shown in fig 4.8 and leave the set up in sunlight for 2 days.
   - Record your observation.

   c) Test any gas collected with a glowing splint.

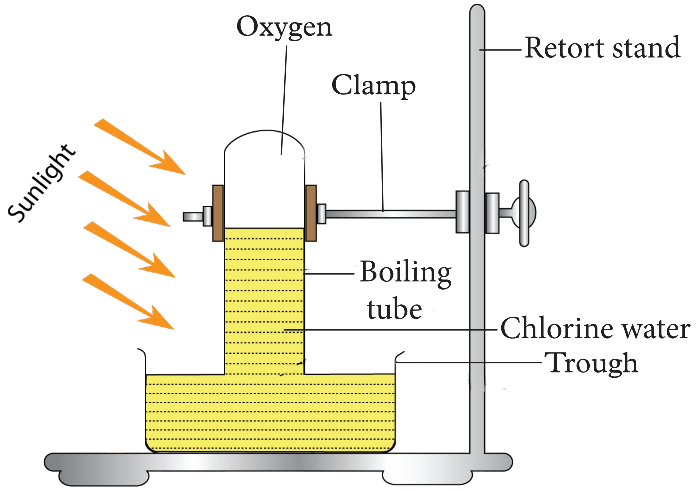


Fig 4.8 Chlorine water in sunlight

Study question

Write chemical equations for the reactions that take place in procedures 2 and 3.

Discussion corner!

1. In your study group, discuss the results of this experiment and the study questions above.
2. Present your results to the rest of the class.

I have discovered that…

Burning hydrogen continues to burn in a gas jar of chlorine. The green-yellow colour of chlorine disappears and a colourless gas is formed in the boiling tube. When chlorine water is exposed to sunlight, a colourless gas is formed in the boiling tube.

The facts

Chlorine reacts with hydrogen to form a colourless gas. The gas is hydrogen chloride.

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

Hydrogen chloride reacts with ammonia to form ammonium chloride as shown in the equation below. Ammonium chloride appears as dense white fumes.

$$NH_3(g) + HCl(g) \longrightarrow \underset{\text{White fumes}}{NH_4Cl(s)}$$

Note: The reaction between hydrogen and chlorine to form hydrogen chloride gas is used industrially to manufacture hydrochloric acid.

In bright sunlight, a mixture of hydrogen and chlorine reacts explosively forming hydrogen chloride gas.

When chlorine gas is bubbled in water, it reacts with water forming hydrochloric acid and hypochlorous acid. This solution of chlorine in water appears yellow and is referred to as chlorine water. Chorine water also contains dissolved chlorine gas.

$$Cl_2(g) + H_2O(l) \longrightarrow \underbrace{HCl(aq) + HOCl(aq)}_{\textit{Chlorine water}}$$

When chlorine water is exposed to sunlight for some time, oxygen gas is produced displacing the solution from the boiling tube as shownFig 4.8.

Hypochlorous acid is an unstable compound. In sunlight, it decomposes to form hydrochloric acid and oxygen.

$$2HOCl\ (aq) \longrightarrow 2HCl\ (aq) + O_2\ (g)$$

c) Bleaching action of chlorine

Activity 4.4

Apparatus, materials and reagents

3 gas jars of chlorine, blue and red litmus papers (dry and moist), coloured flower petals.

Procedure

1. Put moist blue and red litmus papers into a gas jar of chlorine.
   - Note down your observation.
2. Repeat procedure 1 using dry litmus papers.
   - Note down your observation.
3. Place wet coloured flower petals in a gas jar of chlorine.
   - Note down your observation.

Discussion corner!

1. Discuss in groups the observations made in this experiment.
2. Write a report and present it to the class.

I have discovered that…

Moist blue litmus paper placed in a gas jar of chlorine turns red and then white (bleached). The red litmus paper is also bleached. Bleaching also occurs on the coloured flower petals. However, dry blue and red litmus papers in dry chlorine gas are unaffected.

The facts

In the presence of water, chlorine is a bleaching agent. The bleaching action of chlorine is due to the formation of hypochlorous acid when chlorine dissolves in water.

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

As we have already learnt, hypochlorous acid is unstable and decomposes to hydrochloric acid and an oxygen atom.

$$HOCl(aq) \longrightarrow HCl(aq) + [O]$$

In the presence of a dye (coloured pigment) the free oxygen atom reacts with the dye to form an oxidised compound that is colourless (bleached).

$$\underset{\text{Coloured}}{\text{Dye}} + HOCl(aq) \longrightarrow \underset{\text{Colourless}}{\text{Dye + [O]}} + HCl(aq)$$

In the absence of water, bleaching does not occur as hypochlorous acid responsible for bleaching action is not formed.

Note: Whereas sulphur dioxide bleaches by reduction, chlorine bleaches by oxidation. Bleaching by chlorine is therefore permanent unlike bleaching by sulphur dioxide which is temporary (the bleached material may regain its colour by absorbing atmospheric oxygen).

d) Reaction of chlorine with alkalis

Activity 4.5

Apparatus and reagents

Source of chlorine gas, dilute and concentrated sodium hydroxide solutions, dilute and concentrated potassium hydroxide solutions.

Procedure

1. Bubble chlorine gas into cold dilute sodium hydroxide solution.
   - Note down your observations.
2. Repeat the procedure using concentrated sodium hydroxide solution.
   - Note down your observations.

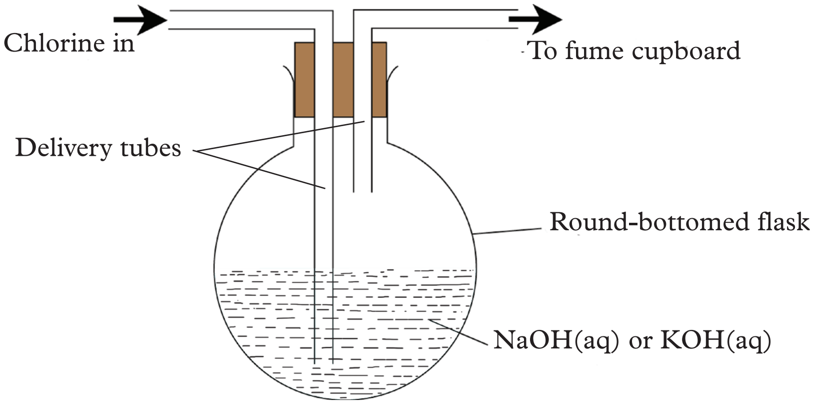


Fig. 4.9: Reaction of chlorine with concentrated and dilute sodium hydroxide

3. Repeat procedure 1 and 2 using potassium hydroxide in place of sodium hydroxide.
    - Note down your observations.

Discussion corner!

1. In pairs discuss the observations above.
2. Compare your findings with other groups.

I have discovered that…

Chlorine gas dissolves in both dilute and hot concentrated alkalis.

The facts

Chlorine reacts with cold dilute sodium hydroxide solution to form a yellow solution of sodium hypochlorite and sodium chloride as shown in the following equation.

$$Cl_2\,(g) + 2NaOH\,(g) \longrightarrow \underset{\text{(Sodium hypochlorite)}}{NaOCl\,(aq)} + NaCl\,(aq) + H_2O\,(l)$$

Just like chlorine sodium hypochlorite is able to bleach dyes.

$$\underset{\text{(Coloured)}}{Dye} + NaOCl\,(aq) \longrightarrow \underset{\text{(Colourless)}}{Dye} + [O] + NaCl\,(aq)$$

Common household bleaches contain sodium hypochlorite as the active ingredient. Chlorine also reacts with hot concentrated sodium hydroxide forming sodium chlorate and sodium chloride according to the following equation.

$$3Cl_2(g) + 6NaOH(aq) \longrightarrow NaClO_3(aq) + 5NaCl(aq) + 3H_2O(l)$$

Similar reactions take place when chlorine is bubbled through cold dilute potassium hydroxide and hot concentrated potassium hydroxide solution. The following are the equations for the reactions that take place.

Cold dilute potassium hydroxide solution

$$Cl_2(s) + 2KOH(aq) \longrightarrow \underset{\textit{(Bleaching agent)}}{\underset{\textit{hypochlorite}}{\underset{\textit{Potassium}}{KOCl(g)}}} + KCl(aq) + H_2O(l)$$

Hot concentrated potassium hydroxide

$$Cl_2(g) + 6KOH(aq) \longrightarrow \underset{\textit{Potassium chlorate}}{KClO_3(aq)} + 5KCl(aq) + 3H_2O(l)$$

Therefore, the product formed when chlorine reacts with alkalis depends on the temperature and concentration of the alkali.

Note: The chlorates (sodium chlorate and potassium chlorate) are used to make explosives, fireworks, matches, weed killer and medicines such as throat lozenges.

When chlorine is bubbled through calcium hydroxide solution, calcium chloride and calcium hypochlorite are formed.

$$2Ca(OH)_2(aq) + 2Cl_2(g) \longrightarrow CaCl_2(aq) + Ca(OCl)_2(aq) + 2H_2O(l)$$

When chlorine is passed over dry solid calcium hydroxide (slaked lime) a solid substance called bleaching powder ($CaOCl_2$) is formed.

$$Ca(OH)_2(s) + Cl_2(g) \longrightarrow \underset{\textit{Bleaching powder}}{CaOCl_2(aq)} + H_2O(l)$$

The green – yellow colour of chlorine disappears as it reacts with calcium hydroxide. When exposed to the atmosphere, bleaching powder smells strongly of chlorine and deteriorates fast. This is because it reacts with atmospheric carbon dioxide releasing chlorine.

$$CaOCl_2(aq) + CO_2\ (g) \longrightarrow CaCO_3(s) + Cl_2\ (g)$$

Self-evaluation Test 4.2

1. Write balanced chemical equations for the reactions that take place when chlorine reacts with the following metals:
   (a) Iron
   (b) Magnesium
2. Materials bleached with chlorine do not gain colour with time as opposed to those bleached using sulphur dioxide. Explain.

4.4 Uses of chlorine

Further activity

Research from the internet about the uses of chlorine and its compounds.

Discussion corner!

1. Discuss with a classmate the uses of chlorine you know.
2. Change roles and let your partner also mention some uses of chlorine.
3. Write down all your points and present them to the class.

The facts

Chlorine has a wide range of uses in our daily lives. Some of the uses of chlorine are outlined below.

1. It is used in the treatment of water and sewage systems. Low concentration of chlorine is used in killing micro-organisms in domestic water supplies. A higher concentration is used for sterilising swimming pool water. Chlorine kills germs.
2. Manufacture of domestic bleaches which consist mainly of sodium hypochlorite like 'Jik' and other bleaching agents used in cotton and paper industries.
3. Manufacture of hydrochloric acid.
4. Manufacture of weed killers such as sodium chlorate. Sodium chlorate is also used to make explosives, fireworks, matches and medicines such as throat lozenges. It is also used in the production of oxygen.
5. Manufacture of germicides, fungicides and pesticides such as DDT.
6. Manufacture of chlorofluorocarbons (CFC's) used to manufacture aerosol propellants, and used in refrigerators as 'freon' and air conditioning units due to their low boiling points.
7. Compounds of chlorine such as vinylchloride are used in the manufacture of plastics (PVC) used in making plastic water pipes, chairs, tables among other.
8. In the manufacture of trichloromethane (chlorofluoros) used as an anaesthetic.
9. Manufacture of cleaning solvents such as carbon tetrachloride.
10. Manufacture of fluid used in fire extinguishers such as pyrene.

4.5 Compounds of chlorine

In this section, you will learn about the following compounds of chlorine: hydrogen chloride gas, sodium chloride and hydrochloric acid.

(a) Hydrogen chloride gas

Activity 4.6

Apparatus, materials and reagents

Round bottomed flask, gas jars, splints, litmus paper, concentrated ammonia solution, Bunsen burner, wash bottle, funnel, beaker, water, glass rod, magnesium ribbon, zinc, iron fillings, aluminium turnings and copper turnings, concentrated sulphuric acid, sodium chloride, silver nitrate solution.

Procedure

1. Place sodium chloride in a round-bottomed flask and set up the apparatus as shown in figure 4.10.
2. Add concentrated sulphuric acid dropwise until the sodium chloride is immersed.
3. Heat the flask gently and collect several gas jars of the gas that you will use to test for the chemical properties of the gas.

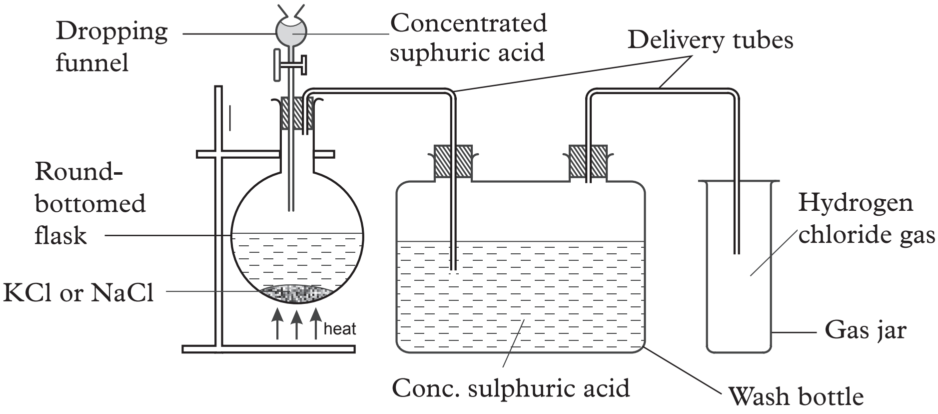


Fig. 4.10: Laboratory preparation of hydrogen chloride

4. a) Direct dry hydrogen chloride gas through the inverted funnel to a beaker of water as shown in fig 4.11.
   - Note down your observation.

   c) Insert moist blue and red litmus papers in a gas jar full of hydrogen chloride gas.
   - Note down your observation.

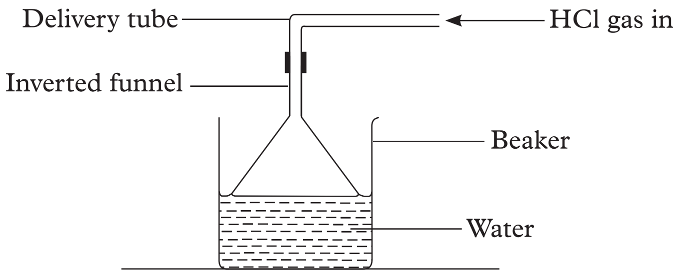


Fig 4.11: Dissolving hydrogen chloride gas in water

d) Dip a glass rod in concentrated ammonia solution and insert it into a gas jar of hydrogen chloride gas.

- Note down your observations.

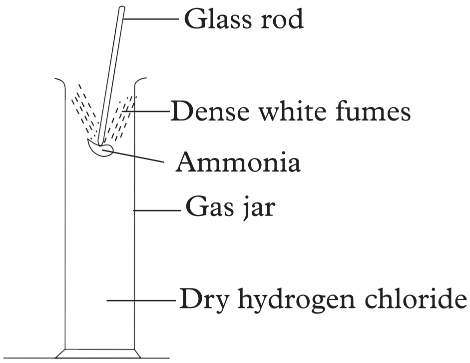


Fig 4.12: Testing for hydrogen chloride using ammonia

e) Lower a burning splint into a gas jar of hydrogen chloride as shown in Fig 4.13.

- Note down your observation.

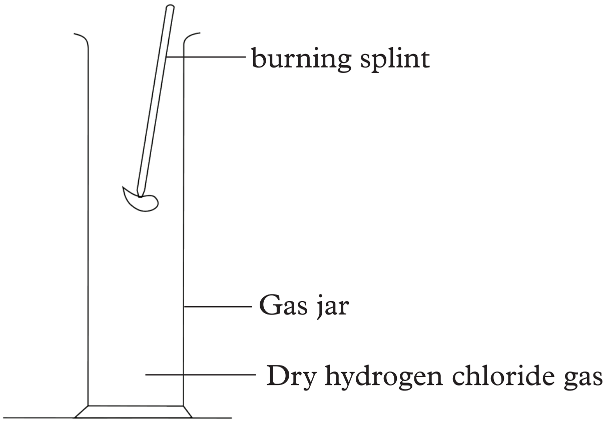


Fig 4.13: Testing for hydrogen chloride using silver nitrate solution.

f) Place a glass rod that has been dipped into silver nitrate solution in a gas jar of hydrogen chloride gas.

Study questions

1. Why is the gas passed through concentrated sulphuric acid?
2. Describe two tests for hydrogen chloride gas.
3. Write a balanced chemical equation for the reaction that took place in the flask.

Discussion corner!

1. In groups, discuss the results of the experiment and the study questions above.
2. Write a report and present it to the class

I have discovered that…

When a mixture of sodium chloride and concentrated sulphuric acid is heated, effervescence occurs and a colourless gas is evolved. On the other hand, when a gas jar full of hydrogen chloride gas is inverted over water, the water rises rapidly in the gas jar. A glass rod dipped in ammonia solution forms dense white fumes when inserted in a gas jar of hydrogen chloride gas. When another glass rod is dipped in silver nitrate solution and introduced in a gas jar of hydrogen chloride, a white precipitate is formed on the glass rod.

The facts

Hydrogen chloride gas is prepared in the laboratory by heating a mixture of sodium chloride and concentrated sulphuric acid. The equation for the reaction is as shown.

$$NaCl(s) + H_2SO_4(l) \longrightarrow NaHSO_4(aq) + HCl(g)$$

Hydrogen chloride is a colourless gas with a pungent irritating smell. The gas is passed through concentrated sulphuric acid to dry it. It is then collected by downward delivery because it is denser than air. The gas can also be dried using anhydrous calcium chloride in a packed U – tube.

Hydrogen chloride is highly soluble in water forming hydrochloric acid. Thus the 'fountain' experiment can be performed to confirm that the gas is very soluble in water.

Physical properties of hydrogen chloride gas

- It is a colourless gas.
- It has a pungent choking smell.
- Its boiling point is -85C and its melting point is -114C.
- It is very soluble in water.

Note: An inverted funnel is used to dissolve the gas in water to prevent 'sucking back'

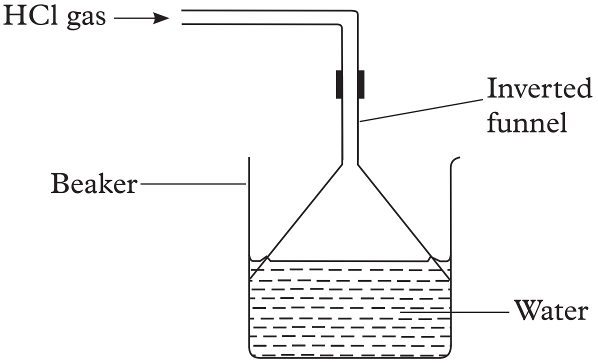


Fig 4.14: Apparatus used to dissolve hydrogen chloride gas in water

Chemical properties of hydrogen chloride gas

• Reaction with ammonia

Hydrogen chloride gas reacts with ammonia gas to form dense white fumes of ammonium chloride.

$$HCl(g) + NH_3(g) \longrightarrow \underset{\text{Dense white fumes}}{NH_4Cl(s)}$$

This reaction is used as a test for hydrogen chloride gas.

• Reaction with acidified silver nitrate solution

Another test for the gas is reacting it with silver nitrate demonstrated in activity 4.6. A white precipitate of silver chloride that forms on the glass rod confirms the presence of hydrogen chloride gas. In this reaction, hydrogen chloride gas first dissolves in water contained in silver nitrate solution to yield hydrogen ions and chloride ions, (hydrochloric acid). The hydrochloric acid then reacts with silver nitrate to form white insoluble silver chloride.

$$HCl(aq) + AgNO_3(aq) \longrightarrow \underset{\text{White ppt}}{AgCl(s)} + HNO_3(aq)$$

• Combustion

Hydrogen chloride gas extinguishes a burning splint. This means hydrogen chloride neither burns nor supports the burning of substances.

- Reaction with metals

Activity 4.7

Materials and reagents

Iron fillings, magnesium ribbon, aluminium powder, copper turnings, glass tube, hydrogen chloride gas from a generator, trough, gas jar, water, source of heat, wooden splint.

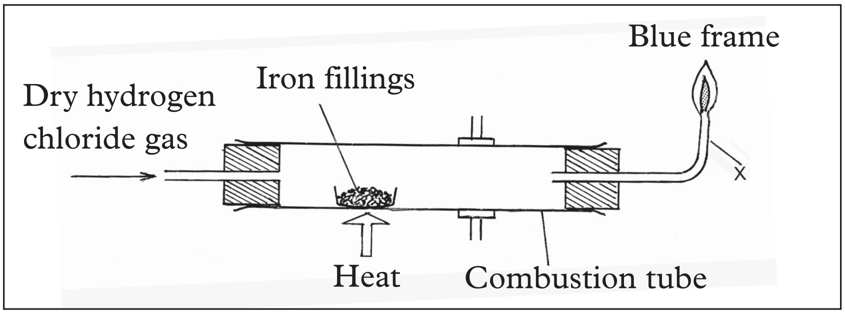


Fig 4.15: Reacting metals with hydrogen chloride

Procedure

1. Set up the apparatus as shown in fig 4.15.
2. Heat the iron fillings and record your observations.
3. Test the gas that is produced using a burning splint.
4. Repeat the experiment using cleaned magnesium ribbon, aluminium turnings and copper turnings instead of iron fillings.
   - Record your observations.

Study question

Write chemical equations for the reactions that take place.

I have discovered that…

When hydrogen chloride gas reacts with metals a gas that burns with a 'pop' sound is produced.

The facts

Hydrogen chloride gas reacts with heated metals above hydrogen in the reactivity series to form the corresponding metal chloride and hydrogen gas. The gas burns with a 'pop' sound. The equations below show the reactions that take place.

$$2HCl(g) + Fe(s) \longrightarrow FeCl_2(s) + H_2(g)$$

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

$$6HCl(g) + 2Al(s) \longrightarrow 2AlCl_3(s) + 3H_2(g)$$

Note: Where the metal exhibits more than one valence, the chloride of the lower valence is formed. Hydrogen chloride gas does not react with copper because copper is lower than hydrogen in the reactivity series.

Self-evaluation Test 4.3

1. Keza wanted to prepare dry hydrogen chloride gas in the laboratory. Draw a set up of apparatus that she would use.
2. Uwase found a certain gas in a gas jar. The gas had a pungent irritating smell. She however decided to carry out a chemical test to confirm if the gas was hydrogen chloride. Describe two ways she would do this.

(b) Preparation and properties of sodium chloride

Activity 4.8

Materials and apparatus

Sodium metal, deflagrating spoon, gas jars of chlorine gas or chlorine gas generator, Bunsen burner, knife, distilled water, concentrated ammonia solution, glass rod, electrical circuit components.

Procedure

1. Place a small piece of sodium metal on a deflagrating spoon and heat it until it starts burning.
2. Insert the burning sodium in a gas jar of chlorine.
    - Note down your observation.
3. Clean the solid formed and perform the following tests on it.
    i) Using a knife cut a small piece of the solid and put it in a test tube. Add 2 cm^3 of water and shake.
    - Note down your observations.

    ii) Test the electrical conductivity of the solution formed in 3(i) by connecting it to a electrical circuit.

    iii) To a small portion of solution in 3 (i) add silver nitrate solution.
    - Note down your observation.

    iv) To another portion of the solid in a test tube, add concentrated sulphuric acid. Dip a glass rod in a concentrated ammonia solution and introduce it at the mouth of the test tube.
    - Record your observation.

Study questions

1. Sodium chloride solution conducts electricity but the solid does not. Explain.
2. Write a chemical equation for the reaction that takes place in procedure 3(iv).

I have discovered that…

Sodium chloride is prepared by direct combination of sodium and chlorine. When burning sodium is introduced in a gas jar of chlorine, a vigorous reaction occurs and a white solid deposit is formed. The solid dissolves in water forming a colourless solution.

The facts

Sodium reacts with chlorine to form a white solids of sodium chloride.

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

Sodium chloride solids dissolve in water to form sodium chloride solution. This solution is a good conductor of electricity.

Physical properties of sodium chloride

- It is soluble in water.
- It is a white crystalline solid at room temperature.
- Has a high melting and boiling points.
- Does not conduct electricity in solid state. However it conducts electricity in molten or aqueous state.

Chemical properties of sodium chloride

- Sodium chloride reacts with concentrated sulphuric acid to produce hydrogen chloride gas which forms dense white fumes when brought near a bottle of ammonia.
- A solution of sodium chloride in water forms a white precipitate with silver nitrate showing that it is ionised.

$$Ag^+(aq) + Cl^-(aq) \longrightarrow \underset{\text{White precipitate}}{AgCl(s)}$$

Uses of sodium chloride

Activity 4.9

Identifying common uses of sodium chloride

Do the following in pairs.

1. Discuss with your classmate about the uses of sodium chloride you know.
2. Change roles and let your partner mention other uses of the salt.
3. Make a list of the uses and present it to the class.

The facts

The following are some of the uses of sodium chloride:

1. It is used in the manufacture of sodium carbonate by the Solvay process.
2. Used in the manufacture of sodium hydroxide, chlorine and hydrogen.

3. Used in the extraction of sodium metal.
4. Used in the manufacture of soap during 'salting out'
5. Used in food as condiment / seasoning food.
6. Medically, sodium chloride solutions are used in catheter flush injections or intravenous injections and cleaning of contact lenses in the form of saline.
7. Sodium chloride is used as a defrosting chemical in cold countries where icing occurs on roads. It is normally added to roads during winter to help melt the ice.
8. Sodium chloride is used as a food preservative where there are no refrigerators for example in preserving fish and other meat foods.

(c) Hydrochloric acid

The effect of hydrochloric acid on metals, bases, manganese (IV) oxide and carbonates

Activity 4.10

Apparatus and reagents

Hydrochloric acid, magnesium ribbon, zinc, iron, copper, sodium hydroxide, copper (II) oxide, calcium carbonate, zinc carbonate, copper (II) carbonate, sodium carbonate, manganese (IV) oxide, test tubes, sodium hydroxide, phenolphthalein indicator, wooden splint, calcium hydroxide solution, flat-bottomed flask, thistle funnel, gas jar

Procedure

1. i) Put some pieces of zinc metal in a test tube. Add 5 cm^3 of dilute hydrochloric acid and test for any gas evolved using a burning splint.

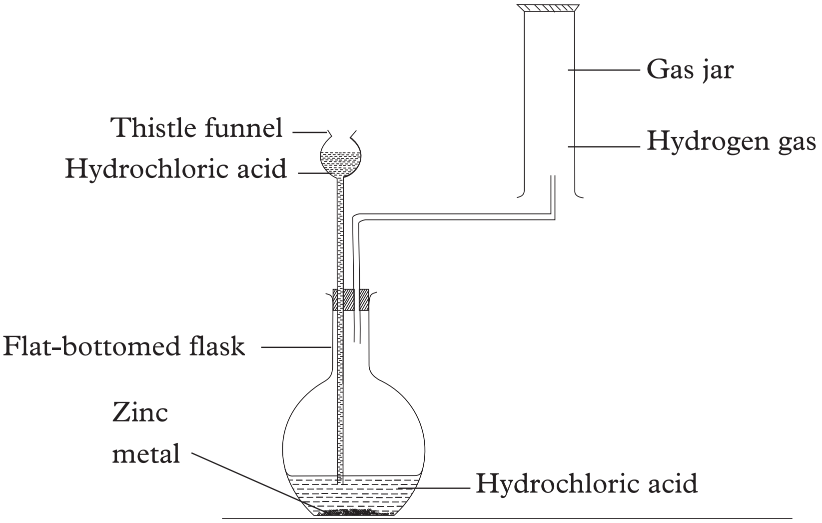


Fig 4.16: Reaction between zinc and hydrochloric acid

(ii) Repeat procedure (i) using magnesium iron and copper in place of zinc.

- Note down your observations.

2. i) Place 5 cm^3 of sodium hydroxide in a test tube and add three drops of phenolphthalein indicator. Add dilute hydrochloric acid to the test tube drop wise until the colour just turns colourless.
   ii) Place some copper (II) oxide in a test tube, add 5 cm^3 of dilute hydrochloric acid and warm.
   - Record your observations.
3 i) Place some calcium carbonate in a test tube and add 5 cm^3 of dilute hydrochloric acid. Test for any gas produced using calcium hydroxide solution using the set up below.

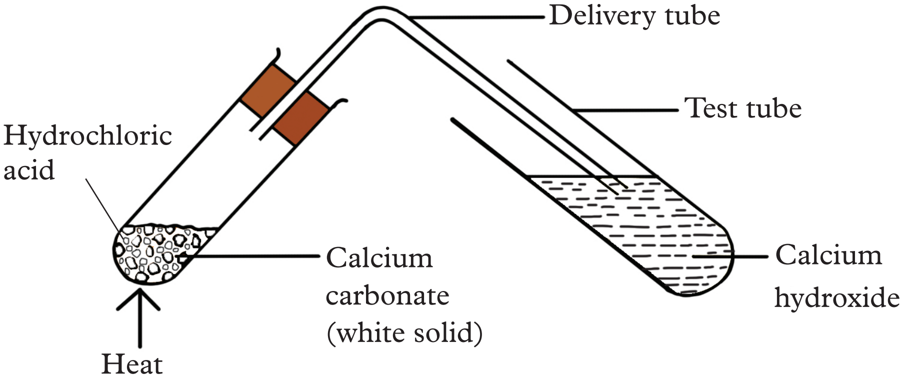


Fig 4.17: Reacting calcium carbonate with hydrochloric acid

   ii) Repeat procedure 3(i) using copper (II) carbonate, sodium carbonate and zinc carbonate.
   - Record your observations.
4. Place a small quantity of manganese (IV) oxide in a test tube and very carefully add some concentrated hydrochloric acid to it.
   - Record your observations.

Study questions

1. Explain the observation made when the gas produced in procedure (1) is tested using a burning splint.
2. Explain the observation made when the gas produced in procedure 3(i) is tested using calcium hydroxide solution.
3. Describe the smell and colour of the gas produced in procedure (4) and give its identity.

Discussion corner!

1. Discuss in groups the observations in the experiment and the study questions above.
2. Present your finding to the rest of the class.

I have discovered that…

Hydrochloric acid reacts with most metals to produce a gas that extinguishes a burning splint and gives a 'pop' sound. However no reaction occurs between copper and the acid. On the other hand, hydrochloric acid reacts with metal carbonates to produce a gas that forms a white precipitate with calcium hydroxide solution. A reaction between the acid and manganese (IV) oxide produces a green-yellow gas that has a pungent irritating smell.

The facts

The following are chemical properties of hydrochloric acid:

Reaction of hydrochloric acid with metals

Some reactive metals such as zinc, magnesium and iron react with hydrochloric acid liberating hydrogen gas and forming the corresponding metal chloride. This can be explained using the following equations.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

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

$$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$$

Very reactive metals such as sodium and potassium react violently with dilute hydrochloric acid forming a salt and hydrogen gas.

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

Copper metal is less reactive therefore, does not react with dilute hydrochloric acid.

Reaction with bases

Hydrochloric acid reacts with bases to form salt and water. This is a neutralisation reaction. The equations for the reactions that take place are as follows.

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

$$2HCl(aq) + CuO(s) \longrightarrow CuCl_2(aq) + H_2O(l)$$

The reaction of hydrochloric acid with oxides is used in the removal of oxides from metal surfaces (pickling) before coating them.

Reaction with carbonates

Hydrochloric acid reacts with metal carbonates to form the corresponding metal chloride, water and carbon dioxide gas as shown in the equations below.

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

$$ZnCO_3(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$$

$$CuCO_3(s) + 2HCl(aq) \longrightarrow CuCl_2(aq) + CO_2(g) + H_2O(l)$$

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

In all these reactions, hydrochloric acid is behaving as a typical acid.

Reaction with manganese (IV) oxide

Hydrochloric acid reacts with manganese (IV) oxide to form manganese chloride, water and chlorine gas.

$$MnO_2(aq) + 4HCl(aq) \longrightarrow MnCl_2(aq) + Cl_2(g) + 2H_2O(l)$$

In this reaction manganese (IV) oxide acts as an oxidising agent. It oxidises hydrochloric acid to chlorine while itself is reduced to manganese chloride and water.

Uses of hydrochloric acid

Activity 4.11

Do the following in pairs

1. Discuss with a classmate substances that you know that are manufactured using hydrochloric acid.
2. Now change roles and let your partner tell you of some uses of hydrochloric acid.
3. Write a joint report and present it to the class.

The facts

The following are some of the uses of hydrochloric acid:

1. Used in cleaning the surfaces of iron to remove the oxide coating before galvanizing and tin plating or chromium plating. The process of using hydrochloric acid to clean metal surfaces is called '**pickling**'.
2. Manufacture of aniline dyes and drugs.
3. Used in the manufacture of photographic materials.
4. It is used in the production of chlorides such as calcium chloride and ammonium chloride.
5. Used to standardise pH of beers and wines.
6. A mixture of concentrated hydrochloric acid and concentrated nitric acid at a ratio of 3:1 respectively is called **aqua regia**. This is the only liquid that can dissolve gold and platinum and is used in cleaning and purifying them.
7. It is used in school and research laboratories.
8. Hydrochloric acid is used in production of photoflash bulbs and fireworks.
9. Used in sugar processing and making gelatin.
10. Hydrochloric acid is found in the human stomach where it helps in the digestion process.

Self-evaluation Test 4.4

1. Sodium chloride is an important substance in industries. Explain
2. Munezero reacted dilute hydrochloric acid with the following substances: manganese dioxide, copper granules, zinc granules and sodium carbonate.
   (a) What observation do you think he made in each case?
   (b) Explain the observations using chemical equations.
3 Hydrochloric acid is an important reagent. Give four uses of the acid in industries.

4.6 Tests for chlorides and other halides in solutions

Activity 4.12

Materials and apparatus

Dilute nitric acid, silver nitrate, sodium chloride solution, sodium bromide solution, sodium iodide solution, ammonia solution, test tubes.

Procedure

1. To 2 cm^3 of sodium chloride solution in a test tube, add 1cm^3 of dilute nitric acid followed by a few drops of silver nitrate.
   - Record your observations.
2. (i) To 2 cm^3 of sodium bromide solution in a test tube, add 1cm^3 of dilute nitric acid followed by a few drops of silver nitrate.
   - Record your observation.

   (ii) To the resulting reaction mixture, add concentrated ammonia solution.
   - Record your observation.
3. (i) To 2 cm^3 of sodium iodide solution in a test tube, add 1cm^3 of dilute nitric acid followed by a few drops of silver nitrate.
   - Record your observation.

   (ii) To the reaction mixture, add concentrated ammonia solution.
   - Record your observation

Discussion corner!

1. Discuss in groups the observations made in the experiment.
2. Compare your findings with those of the groups.

The facts

When silver nitrate solution is added to an acidified solution of chloride ions, a white precipitate is formed. The precipitate is silver chloride.

$Ag^+(q) + Cl^-(g) \longrightarrow AgCl(s)$

(White precipitate)

The solution is acidified by adding dilute nitric acid. The nitric acid helps removes other ions that might give confusing precipitates for example, carbonate ions (CO_3^{2-}). The white precipitate is soluble in concentrated ammonia solution to form a colourless solution. When silver nitrate solution is added to an acidified solution containing bromide ions, a pale-yellow precipitate is formed.

$Ag^+(aq) + Br^-(g) \longrightarrow AgBr(s)$

(Pale-yellow precipitate)

The pale-yellow precipitate is soluble in concentrated ammonia solution to form a colourless solution.

When silver nitrate solution is added to an acidified solution containing iodide ions a yellow precipitate is formed.

$Ag^+(aq) + I^-(aq) \longrightarrow AgI(s)$

(Yellow precipitate)

The yellow precipitate is insoluble in ammonia solution of any concentration.

Note: Silver chloride precipitate is obviously white, but silver bromide and silver iodide precipitates are not very different, hence addition of concentrated ammonia solution is used to differentiate between them as shown in table 4.1 below.

Table 4.1: Differences between the three silver halides

Original precipitate	**Observation on addition of concentrated ammonia solution**
AgCl	Precipitate dissolves in concentrated ammonia to give a colourless solution.
AgBr	Precipitate dissolves in concentrated ammonia to give a pale-yellow solution.
AgI	Precipitate (yellow in colour) is insoluble in ammonia solution of any concentration.

Self-evaluation Test 4.5

1. Describe how you can confirm the presence of the following ions in a given solution

   (a) Cl^- (b) Br^- (c) I^-

4.7 Environmental issues of chlorine and its compounds

Activity 4.13

Research on the toxicity of chlorine and its compounds and the impact on the environment

Research should be guided by the following questions. Write a report for presentation to the class.

1. Name some of the compounds of chlorine that affect the environment.
2. How do the compounds identified in (1) above get into the environment.
3. How does each of the compounds named in (1) affect the environment.
4. How can environmental pollution by chlorine and its compounds be minimised?

Health check!

Chlorine is a poisonous gas. It was used as a chemical weapon during World War 1 where it was dropped from aeroplanes. Due to its high density it would flow down to the ground mixing with air. Mixing of chlorine bleach (sodium hypochlorite) with ammonia or acidic cleaning agents is a common source of household exposure.

When chlorine is inhaled, it is pulmonary irritating with intermediate water solubility that causes acute damage in the lungs. Respiratory system is also the target of chlorine toxicity in animals. Inhaling of chlorine gas can cause difficulty in breathing, chest pains, cough, eye, irritation. Increases heart beat, rapid breathing and can result in death.

The facts

(a) **Chlorine compounds** also affect the environment when they are used. Chlorofluorocarbons (CFC's) are non-toxic, nonflammable organic chemicals containing atoms of carbon, chlorine and fluorine. They are used in the manufacture of aerosol sprays, blowing agents for foams and packing materials as solvents and as refrigerants and air conditioners. They are used as propellants in aerosols. CFC's are non-biodegradable and hence have a long life span. When released into the environment, the CFC's find their way into the upper atmosphere. Here they break down into free chlorine and fluorine atoms acted on by ultraviolet light. These free atoms are very reactive and react with ozone molecules (O_3) breaking them down into oxygen molecules and oxygen atoms.

$$O_3(g) \xrightarrow{\text{Chlorine atom}} O_2(g) + [O]$$

The free oxygen atoms are also very reactive and react with more ozone molecules.

$$O_3(g) + O(g) \longrightarrow 2O_2(g)$$

Thus chlorine destroys the ozone layer. The ozone layer is important because it protects us from harmful ultra violet radiation. Ozone layer normally absorbs harmful ultra violet (UV) light rays from the sun. Once the ozone layer is depleted, the harmful ultra violet rays find their way to the earth's surface. These ultra violet rays have adverse effects on the flora and fauna life. The rays cause certain cancers of the skin, immunity suppression, eye cataract and respiratory problems in humans and other animals. They also damage plants.

CFC's also absorb the heat from the earth and prevents it from escaping into space. They therefore contribute to the 'greenhouse' effect or global warming.

(b) Another (DDT) compound of chlorine is called **dichloro – diphenyl – trichloroethane (DDT)**. It was used as an insecticide and was very effective in control of mosquitoes. DDT can get into human body through direct consumption or inhalation. DDT is harmful to human beings and other animals and plants. It accumulates in the fat tissue of the human body where it is not easily excreted. When DDT is washed into water bodies it is absorbed by tiny sea plants and animals called planktons, which are easily eaten by fish. When fish are eaten by human beings and other animals, the DDT gets into their bodies.

Accumulated DDT damages the nervous system. DDT has made egg shells of some birds to become so thin and fragile that they break during incubation.

Use of DDT has been banned worldwide including Rwanda and other East Africa countries.

(c) **Dioxins** are compounds which contain carbon, chlorine, hydrogen and oxygen. These compounds are formed as by-products during the manufacture of organo chlorine compounds. Dioxins are also formed when wood treated with organ chlorines; plastics such as PVC or paper bleached with chlorine are burnt or incinerated. Dioxins are very toxic chemicals and cause birth defects, damage to liver and kidney and also cause cancer.

(d) Polychlorinated biphenyl (PCB) is used in transformers in electricity substances, electrical capacitors and in washing machines. PCBs are used in inks, tapes, plastics, paints, glues, waxes and polishes. They are harmful to fish and other aquatic life because they interfere with reproduction. In humans they cause skin lesions, liver ailments and also damage the nervous system.

Ways of reducing effects of chlorine and its compounds on the environment

i) Using alternative chemicals for insecticides rather than DDT.

ii) Reduce our reliability on organ chlorides.

iii) Stop burning or incineration of waste products especially PVC and painted or treated wood.

iv) Use of protective clothing when spraying organo chloride pesticides.

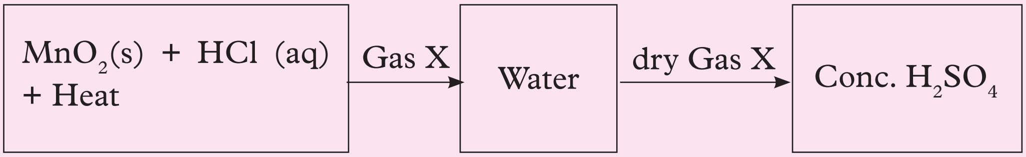
v) Reduce consumption of animal fat and avoid consuming fish from regions where organo chlorine pesticides are used.

Unit summary 4

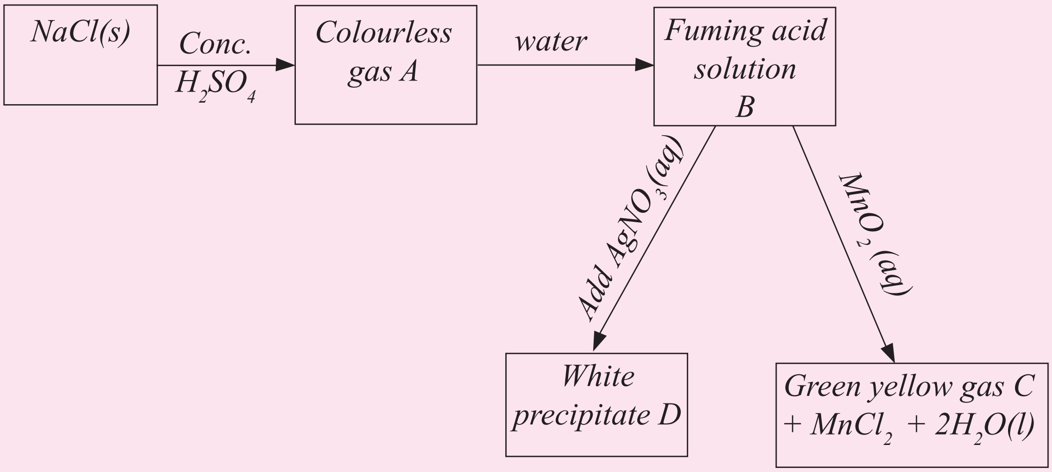
- Chlorine is a green yellow poisonous gas that dissolves in water to form hydrochloric and hypochlorous acids.
- In the laboratory chlorine is prepared by reacting either manganese dioxide or potassium manganate (VII) with concentrated hydrochloric acid. This experiment should be done in a fume chamber.
- Chlorine is a bleaching agent. Its bleaching action is due to the decomposition of hypochlorous acid to release an oxygen atom that combines with the dye in a substance or material.
- Some uses of chlorine include water treatment, manufacture of bleaching powders, and pesticides among others.
- Hydrogen chloride is a colourless gas with a pungent irritating smell and is very soluble in water to form hydrochloric acid.
- In the laboratory, hydrogen chloride is manufactured by heating a mixture of sodium chloride and concentrated sulphuric acid.
- Sodium chloride is a compound of chlorine that is prepared by a direct combination of sodium and chlorine.
- Sodium chloride is used in the manufacture of sodium carbonate by Solvay process, manufacture of soap, as a food preservative among other uses.
- A reaction between a solution containing chloride ions and acidified silver nitrate forms a white precipitate of silver chloride. This is the confirmatory test for chloride ions.

Test your Competence 4

1. How does chlorine occur in nature.
2. List the measures that can be taken to reduce the effects of chlorine and its compounds to the environment.
3. A small amount of chlorine added to water can ————————it.
4. What state is chlorine at room temperature?
5. Study the flow chart below then answer the questions that follow.



   (a) Name gas X.
   (b) What is the use of water?
   (c) What is the use of concentrated sulphuric acid?
6. What is observed when the following metals are separately reacted with dry chlorine?
   (a) Burning magnesium
   (b) Red-hot iron
7. (a) Write an equation for the reaction between chlorine and sodium hydroxide solution.
   (b) Explain how the resulting solution acts as a bleaching agent.
8. Use the flow chart below to answer the questions that follow.



(a) Name gas A and C.
(b) Name precipitate D.
(c) Name the solid which can be heated with solution B to produce gas C.

9. Explain three effects of chlorine to the environment.

UNIT
5

Rate of reactions

Key unit competency

To be able to describe and explain the effect of different conditions on the speed of reactions.

Learning objectives

By the end of this unit, I should be able to:

- Define the rate of a reaction and describe the effects of different conditions on the speed of reactions.
- Develop observation skills in experiments that involve changes of colour.
- Develop research and scientific reporting skills.
- Interpret mass against time graphs for different reactions.
- Develop a culture of working in a team during discussions and research.
- Appreciate that reactions in everyday life like burning and rusting happen at different rates.
- Respect the procedures in an experiment.

Mind teaser

Look at the photograph below. It shows milk processing plant. The raw materials involved are milk, water and empty packets.

Fig 5.1: Milk processing plant

In order to increase the amount of packets of milk processed per day, what do you think the management requires to do? In chemical reactions, what does your answer to the question above compare to?

Introduction

The rate of chemical reactions can be increased or decreased by altering certain factors that affect the reactions. For example, catalysts are used to speed up the rates of chemical reactions and remain unchanged at the end of the reaction. Other factors that affect the rate of chemical reactions are temperature, concentration of reactants and products and pressure. Others are particle size of reactants and light among others.

5.1 Definition of reaction rate (speed of reaction)

Activity 5.1

Apparatus and reagents

Nails, beaker, water in a tin, magnesium ribbon, a pair of tongs, Bunsen burner and charcoal.

Procedure

1. Dip nails into a beaker of water.
2. Remove the nail from the beaker and place them in an open place.
    - Note what happens to the nails after some days.
3. Using a pair of tongs place magnesium ribbon on a lit Bunsen burner. Note what happens to the magnesium ribbon.
4. Light the charcoal and record the time it takes to burn completely.

Study question

1. How long does it take the nails to rust?
2. How fast does the magnesium ribbon burn?
3. Give the time taken by the charcoal to burn completely.

I have discovered that…

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

The facts

Rate of reaction can therefore be defined as change of an amount or concentration of a particular reactant or product per unit time.

Different reactions have different reaction rates. Some are fast, and others slow or very slow. For example, burning of magnesium is very fast, burning of charcoal is slow and rusting is very slow.

5.2 Effect of different conditions on the speed of reaction

(a) Effects of temperature on the rate of reaction

Activity 5.2

Apparatus and reagents

250 cm^3 conical flask, thermometer, stop clock, piece of white paper, 50 cm^3 measuring cylinder, sodium thiosulphate solution(40 dm^3), hydrochloric acid(2mol/dm^3), test tubes, clamp and stand, tripod, gauze, graph paper.

It is not necessary for these exact temperatures to be used although the temperature used should **not** exceed 55°C.

Note: Dilute hydrochloric acid will be added to sodium thiosulphate solution at different temperatures in a series of experiments.

Procedure

1. Using a 50 cm^3 measuring cylinder, measure out 10 cm^3 of the 0.25 mol dm^{-3} sodium thiosulfate solution provided and 40 cm^3 water and pour this mixture into a conical flask.
2. Using a 10 cm^3 measuring cylinder, measure 5 cm^3 of the 2.0 mol dm^{-3} hydrochloric acid into a test tube.
3. Prepare a water bath by half-filling a beaker with water and place the beaker over a tripod and gauze.
4. Clamp the conical flask in the water bath and also place the test tube in the water bath.
5. Place the thermometer in the conical flask and very gently heat the water bath until the contents of the flask reach the required temperature for the experiment.
6. Remove the conical flask and the test tube from the water bath.
7. Place the conical flask on the paper marked with a cross and immediately add the acid from the test tube as you start the stopwatch. Gently stir the mixture with the thermometer. Record this as the initial temperature.

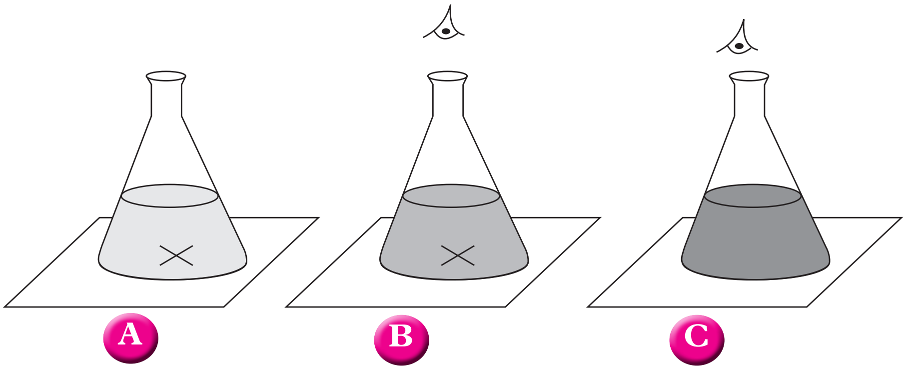


Fig 5.2: Experiment on the effects of temperature on rate of reaction

8. Observe the cross from above through the solution and stop the stopwatch as soon as the cross disappears from your view.
9. Record the time taken for the cross to be obscured and record the final temperature of the mixture.
10. Repeat the experiment for each of the five temperatures suggested in the Table 5.1 and record all of your results.

Table 5.1: Results of the conditions that affect reaction rate

Experiment	1	2	3	4	5
Temperature (°C)	20	25	30	35	40
Time taken for the cross to disappear(s)					
$\frac{1}{t}$					

As soon as possible, pour the solution down the sink (in a fume cupboard if possible) and wash away.

Study Questions

1. What is the name of yellow substance formed?
2. Write equation for the reaction that occurs.
3. What other observations are made during the reaction?
4. For each set of results, calculate the value of $\frac{1}{\text{time}}$. (This value can be taken as a measure of the rate of reaction for this experiment).
5. Plot a graph of $\frac{1}{\text{time}}$ on the vertical (y) axis and average temperature on the horizontal (x) axis.
6. Comment on the shape of the graph.
7. What is the effect of increasing temperature on the reaction time?
8. What is meant by saying that two quantities are directly proportional?

Discussion corner!

1. Discuss in your study groups the observations made and the study questions above.
2. Prepare a report of your findings and present it to the class.

I have discovered that…

When sodium thiosulphate reacts with an acid, a yellow precipitate of sulphur is formed. To follow this reaction we can measure how long it takes for a certain amount of sulphur to form.

This is done by observing how long it takes the black cross on white paper t to be obscurred.

We view it from above through the conical flask as indicated in Fig.5.3.

Sodium thiosulphate + hydrochloric acid ⟶ sodium chloride + sulfur dioxide + water + sulphur

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

The facts

The speed or rate of reaction can be expressed as 'x amount of sulphur' per unit time, so the rate is proportional to $\frac{1}{\text{time}}$ for a given set.

When gases or liquids are heated the particles gain kinetic energy and move faster. The rate of reaction increases because the velocity of all the reacting particles increase as the temperature rises.

Increase in velocity increases the frequency of collisions of reacting particles. This is because of the increase in kinetic energy which provides the particles with the necessary activation energy required for the reaction to occur. The reaction therefore proceeds at a faster rate.

However, this is not the main reason for the increased reaction speed.

Consider the following graph of volume of hydrogen produced against time when a metal reacts with dilute hydrochloric acid or sulphuric acid.

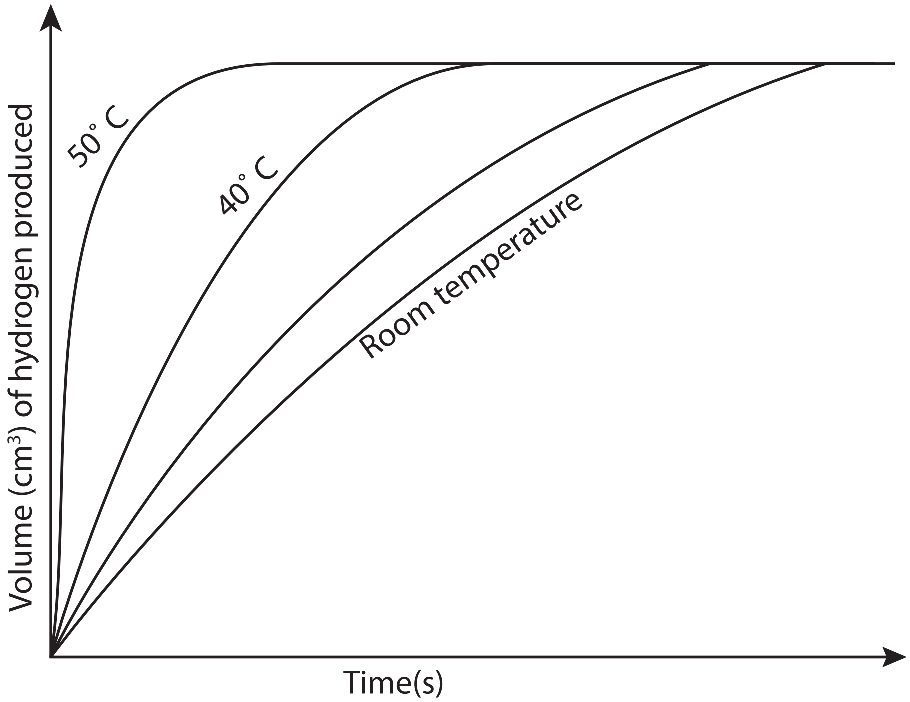


Fig 5.3: Graph of volume versus time

When the temperature of the reactants is increased a higher volume of hydrogen is produced per unit time. The volume of hydrogen produced per unit time is a measure of the rate of reaction because the velocity of all the reacting particles increase as the temperature increases.

b) Effects of concentration on the rate of reaction

Activity 5.3

Apparatus and reagents

- 250 cm^3 conical flask
- Piece of paper marked X
- 100 cm^3 measuring cylinder
- Sodium thiosulphate solution of different concentrations (0.32M, 0.25M, 0.18M, 0.13M and 0.06M).
- 2M hydrochloric acid
- Stop watch

Procedure

1. Put 50 cm^3 of 0.32M sodium thiosulphate solution into a conical flask.
2. Measure 5 cm^3 of 2M hydrochloric acid into the measuring cylinder.
3. Add the acid to the conical flask and immediately start the stopwatch. Swirl the flask to mix the solutions and place it on a piece of paper marked with a cross.

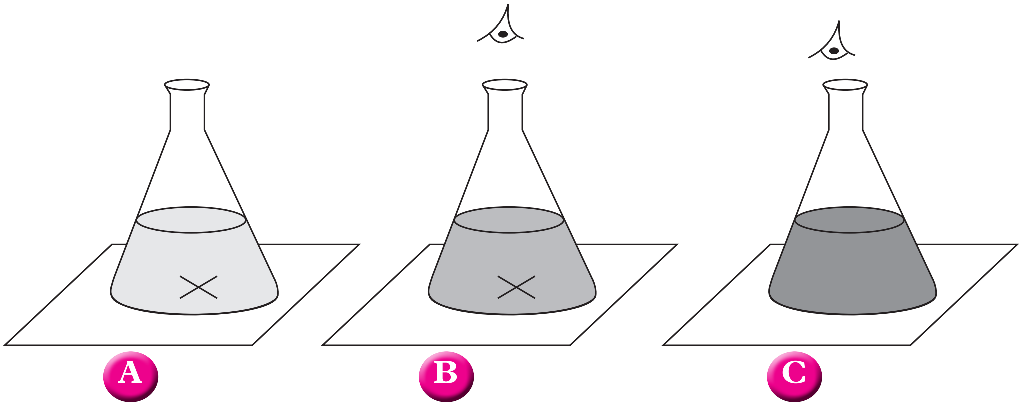


Fig 5.4: Experiment on effects of concentration on rate of reaction

4. Look down at the cross from above. When the cross disappears stop the watch and note the time taken. Record this in a table like the one below.

Table 5.2. Results of the experiment

Experiment number	**1**	**2**	**3**	**4**	**5**
Concentration of $Na_2S_2O_3$ in mol/dm^3	0.32	0.25	0.18	0.13	0.06
Time taken for cross to disappear(seconds)					
$\frac{1}{t}(s^{-1})$					

5. Repeat the experiment using the different concentrations of sodium thiosulphate solution. As soon as possible, pour the solution at the end of every reaction down the sink (in a fume cupboard if possible) and wash away.

Study questions

1. Write the equation for the reaction that occurs.
2. For each set of results, calculate the value of $\frac{1}{t}$ and fill the table. (This value can be taken as a measure of the rate of reaction).
3. Plot a graph of $\frac{1}{t}$ on the vertical (y) axis and concentration of the thiosulphate on the horizontal (x) axis.
4. Comment on the shape of the graph.
5. What is the effect of raising the concentration on the reaction?

I have discovered that…

When sodium thiosulphate and dilute hydrochloric acid are mixed, a yellow precipitate of sulphur is formed. The higher the concentration of sodium thiosulphate solution the less the time taken for the cross to disappear (get obscurred).

The facts

The higher the concentration of the reactants (see Fig. 5.5), the faster the rate of reaction. This is because increasing the concentration of reactants increases the frequency of fruitful collisions between reacting particles. The greater the number of collisions, the faster the rate of reaction.

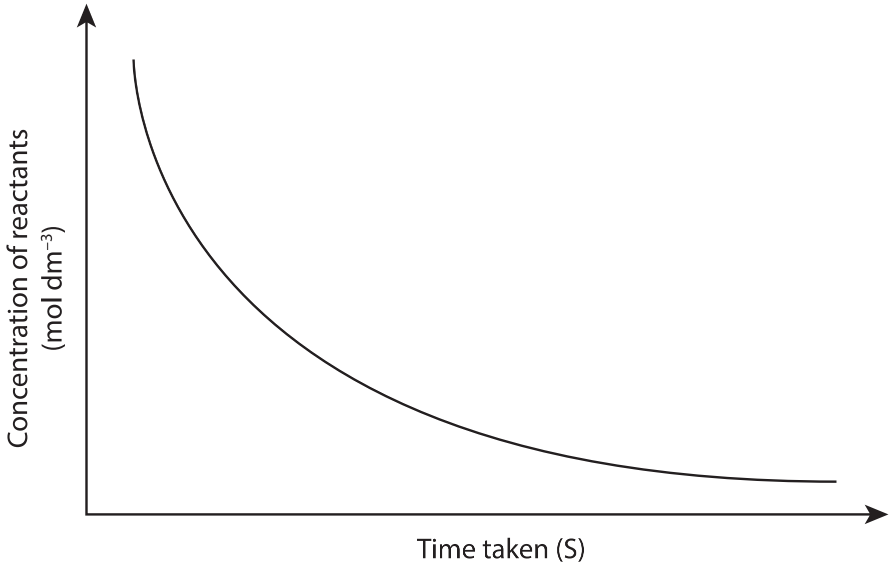


Fig 5.5: Graph of concentration of reactants versus time

The frequency of collisions between reacting particles is influenced by the number of particles closer to each other in given volume unit.

(c) Effects of particle size on the reaction rate

Activity 5.4

Apparatus and Reagents

Magnesium ribbon, 1 M hydrochloric acid, distilled water, 2 conical flasks (100 cm^3), measuring cylinder (100 cm^3), clamp stand, glass trough, stopwatch, sand paper and weighing balance.

Procedure

1. Clean the magnesium ribbon using a sand paper to remove any oxides that may be coating its surface. This will reduce reaction errors related to impurities.
2. Cut three equal sizes (10 cm each) of magnesium from the freshly cleaned magnesium ribbon, weigh each of them using a digital weighing balance and record their weights. Are they the same?
3. Wrap the magnesium pieces immediately in an aluminium foil to prevent them from being re-oxidised.
4. Fold the first magnesium ribbon and keep it for the next procedure.
5. Take the second magnesium ribbon cut it into smaller pieces, repeat this until you obtain very fine pieces of magnesium.
6. Measure 20 cm^3 of 1M hydrochloric acid using a clean dry measuring cylinder and pour into a clean 100 cm^3 conical flask. Repeat this one more time in the second conical flask. Label them A and B.

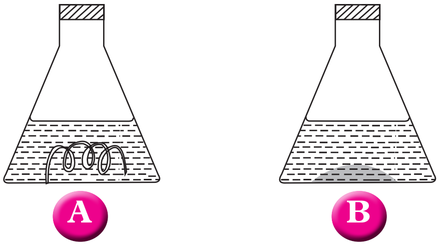


Figure 5.6: Experimental set-up for measuring reaction rates between HCl and magnesium of different surface areas

7. Simultaenously put the folded magnesium ribbon and the fine pieces magnesium powder into conical flasks A and B respectively as shown above. Immediately start the stopwatch.
8. Monitor the reaction progress closely and stop your stopwatch when the magnesium completely dissolves in the acid or reaction comes to stop. In each case, record the reaction duration in seconds in the table like the one shown below.
9. Repeat steps 7-8 using 0.5M, 1.0M, 1.5M, 2.0M, 2.5M and 3.0M HCl.

Table 5.3: Table of results showing HCl-Mg ribbon reaction

Concentration of HCl (M)	3.0	2.5	2.0	1.5	1.0	0.5
Reaction duration for ribbon (in seconds)						
Reaction duration for fine pieces (in seconds)						

Study questions

1. What observations did you make during the reaction?
2. Write a chemical equation for the reaction that occurs.
3. What conclusion can you make from this experiment? Why?

Discussion corner!

1. In your study groups, discuss the observations made and the study questions above.
2. Write a report and present it to the rest of the class.

I have discovered that…

The reaction involving magnesium pieces came to an end much faster than magnesium ribbon. Also effeverscence was more rapid in this experiment than in the reaction involving magnesium ribbon.

The facts

If solid reactants are broken down into smaller pieces, the rate of reaction increases. The increase in rate of reaction happens because many smaller pieces of the same mass of solid provide a greater surface area compared to larger pieces. This generally increase the contact surface of the reactants colliding with each other thereby increasing reaction rate.

At higher acid concentrations, the rate of reaction are even more rapid than in lower acid concentrations as shown in the graph below.

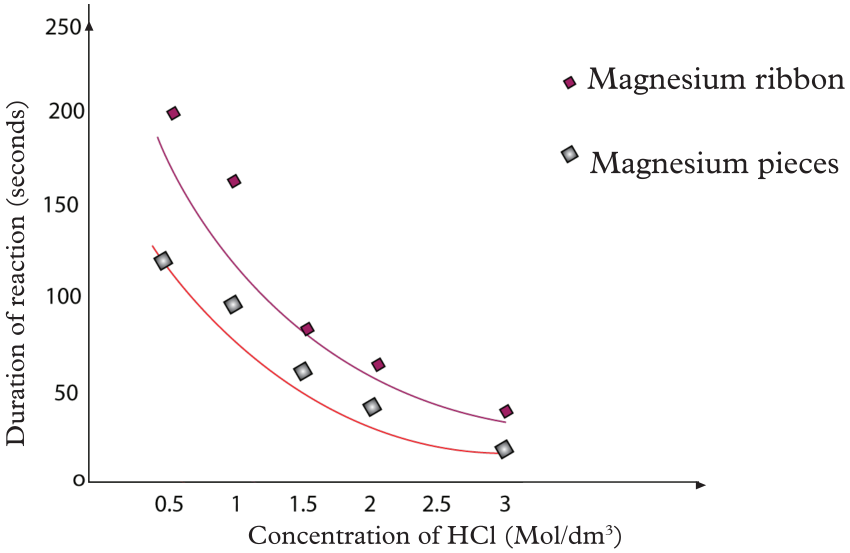


Fig 5.7 Graph of reaction rate of magnesium at different concentrations and particle sizes of hydrochloric acid

Self-evaluation Test 5.1

1. Which of the following gases is formed when an acid reacts with a carbonate?
   A. Carbon dioxide
   B. Chlorine
   C. Oxygen
   D. Hydrogen
2. In an experiment the following carbon dioxide gas volumes were recorded by a student from an acid-limestone reaction.

Time (min)	0.0	0.5	1.0	1.5	2.0	2.5	3.0	3.5	4.0	4.5	5.0	5.5	6.0
Vol (cm^3)	0	8	15	21	26	30	33	36	38	39	40	40	40

   At what time did the reaction stop?
   A. 6.0 min
   B. 5.5 min
   C. 5.0 min
   D. 4.5 min
3. A catalyst can increase the rate of a chemical reaction. Which of the following is also TRUE about catalysts?
   A. They have to be a solid.
   B. They are chemically the same at the end of the reaction.
   C. Their appearance is unchanged at the end of the reaction.
   D. They have to be a metal.
4. Calcium carbonate in the form of limestone dissolves in hydrochloric acid giving off carbon dioxide gas. Which of the following conditions would produce the fastest reaction?
   A. Temperature of 40°C and 10cm^3 of dilute acid
   B. Temperature of 20°C and 10cm^3 of dilute acid
   C. Temperature of 40°C and 5cm^3 of dilute acid and 5cm^3 of water
   D. Temperature of 20°C and 5cm^3 of dilute acid and 5cm^3 of water
5. Platinum catalyses the reaction between sulphur dioxide (SO_2 and oxygen (O_2) to form sulphur trioxide (SO_3) Which statement(s) is **not** correct?
   (i) Less sulphur trioxide is formed when more catalyst is used.
   (ii) The catalyst decreases the rate of reaction between sulphur dioxide and oxygen.
   (iii) When the reaction is complete the platinum catalyst has gained mass.
   A. i, ii and iii
   B. only i and iii
   C. only i
   D. only ii and iii
6. Potatoes cook more quickly in boiling vegetable oil than in boiling water. This is because oil ____________ ?

A. Boils at a higher temperature than water.
B. Contains water.
C. Replaces the water in the potatoes.
D. Catalyzes the cooking process.

7. Which chemical species below will react the fastest with hydrochloric acid?
   A. Powdered marble at 40°C
   B. Powdered marble at 25°C
   C. Marble chips at 25°C
   D. Marble chips at 40°C

(d) Effects of catalyst on the reaction rate

Activity 5.5

Apparatus and reagents

3 conical flasks, measuring cylinders (10 cm^3), dropping pipette, zinc granules, copper turnings or powder, dilute sulphuric acid, 0.5M copper (II) sulphate solution.

Procedure

1. Put a few pieces of zinc granules into each of the three conical flasks. Name them A, B and C. (Weigh approximately 5 grams of zinc into each conical flasks).
2. Add 5 cm^3 of dilute sulphuric acid to conical flask A. Measure the amount of gas produced as shown below.

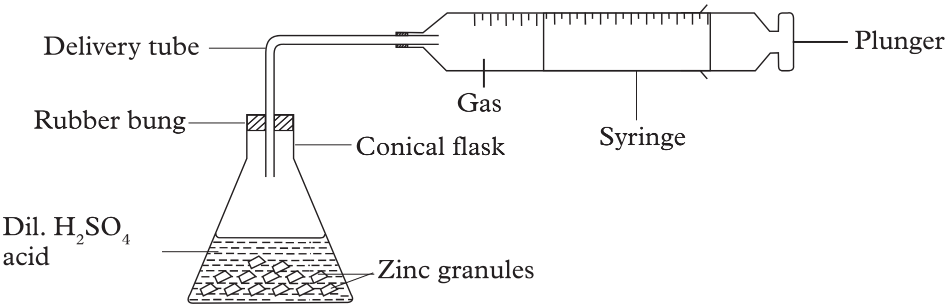


Figure 5.8: Experimental set-up for the reaction between HCl and magnesium

3. In conical flask B, add a few copper turnings. Make sure they are in contact with the zinc granules. Add 5 cm^3 of dilute sulphuric acid and insert the rubber bung quickly. Measure the amount of gas produced.
4. Add 5 cm^3 of dilute sulphuric acid to conical flask C. Add about 1 cm^3 of copper (II) sulphate solution using a dropping pipette. Insert the rubber bung quickly. Measure the amount of gas produced as in procedure 2 above.

5. Record the results in a table like the one shown below.

Table 5.4. Results of the experiment

Experiment	**A**	**B**	**C**
Amount of gas produced (in cm^3)			

Study questions

1. Write equation for the reaction between zinc and dilute sulphuric acid.
2. What was the purpose of adding copper (II) sulphate solution in experiment C?
3. Write the equation for the reaction between copper (II) sulphate and zinc metal.
4. Explain how you would test for the gas evolved.

Discussion corner!

Discuss in your study groups the observations made and the study questions above.

I have discovered that…

If the granulated zinc pieces are shiny then the reaction in conical flask is slow. The reaction may be faster if the zinc is not very pure. Bubbles of hydrogen form on the surface of the zinc. The reaction that occurs is:

$$\text{Zinc} + \text{Sulphuric acid} \longrightarrow \text{Zinc sulphate} + \text{Hydrogen}$$

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

In conical flask 2, copper turnings is the catalyst for the reaction, and the reaction is faster than in conical flask 1, but is not as fast as that in conical flask 3.

In conical flask 3, zinc displaces copper from the copper (II) sulphate solution and the surface of the zinc goes black. The displaced copper metal then acts as a catalyst for the reaction i.e.

$$\text{Zinc} + \text{copper (II) sulphate} \longrightarrow \text{zinc sulphate} + \text{copper}$$

$$Zn(s) + CuSO_4\,(aq) \longrightarrow ZnSO_4(\,aq) + Cu(s\,)$$

The facts

A catalyst does not cause the reaction but only speeds up the rate of a chemical reaction that would have otherwise taken place at a much slower rate in its absence. Catalyst remain unaffected at the end of the reaction.

e) Effect of pressure on reaction rate

Look at the syringes below. If the end is sealed, how can you increase the pressure of the gases inside the syringe?

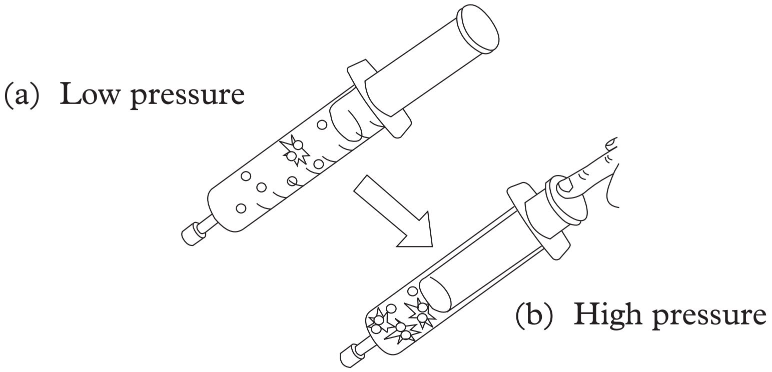


Fig 5.9 Syringes showing how pressure can be increased

To increase pressure, you push the plunger downwards (inward) as shown. By pressing the plunger in, you will have the same number of gas particles in a smaller volume. This reduces the space is which collisions take place thereby increasing the chances of fruitful collisions in a given time. As a result, the rate of reaction is increased.

f) Effect of light on reaction rate

Activity 5.6

Reagents and apparatus

Pair of scissors, cardboard, A4 paper, silver bromide solution, art brush, dark room

Procedure

1. Cut a piece of paper into a square shape using a pair of scissors then cut a piece of cardboard to the same size as paper.
2. In a dark room smear the paper with silver bromide solution.
3. Cover the lower part of the paper with card board as shown in the fig 5.10.
4. Cover the upper part with a thin piece of paper.

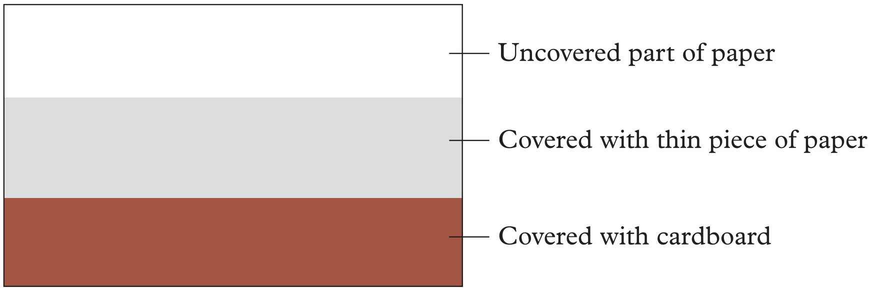


Fig 5.10: Experimental set up for determining the effect of light on silver bromide

5. Expose the improvised film to light, leave it for about 5 minutes.
6. Remove all the coverings then make your observations.

Study questions

1. Why was the paper smeared in dark place?
2. Why was the cardboard used in the experiment?
3. What observations did you make after exposure to light?

Discussion corner!

Discuss the observations made in this experiment and the study questions above.

I have discovered that…

Silver bromide decomposes to silver metal which will be seen as black spots. On the cardboard area no black spots seen because light cannot pass through it.

The facts

Silver chloride (AgCl), silver bromide (AgBr) and silver iodide (AgI) are all sensitive to light (i.e. are 'photosensitive') and all the three are used in the production of various types of photographic film to detect visible light, beta and gamma radiations. Each silver halide salt exhibit different sensitivity towards light. When the radiation hit the film the silver ions in the salt are reduced by electron gained to silver metal according to the following equation.

$Ag^+(aq) + e^- \longrightarrow Ag(s)$; the halide ion is oxidised to the halogen molecule i.e.

$2X^- \longrightarrow X_2 + 2e^-$

Note: The above reaction does not occur in absence of light.

AgI is used in X-ray radiography, AgCl is used in 'fast' film for cameras.

Photosynthesis in green plants is also sensitive to light. Light energy is used to combine water and carbon dioxide to form glucose during photosynthesis. The equation of the reaction is:

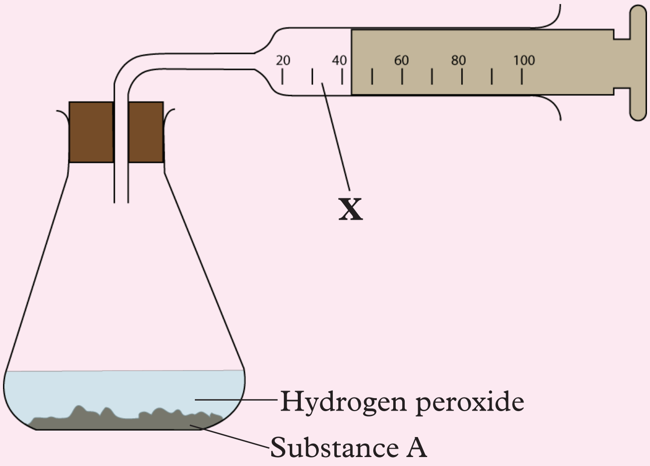
$$6H_2O\ (l) + 6CO_2\ (g) \xrightarrow[\text{chlorophyll}]{\text{light}} C_6H_{12}O_6\ (aq) + 6O_2(g)$$

Self-evaluation Test 5.2

1. The set-up below is used to study the rate of reaction during decomposition of hydrogen peroxide.

   (a) Name the gas labelled X in the syringe.

   (b) Identify substance A and its role in the experiment.



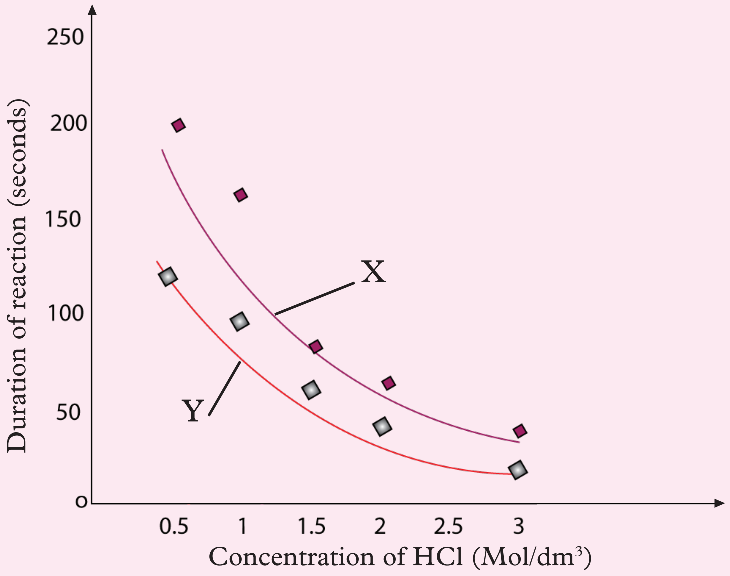
2. In the catalytic converter in a motor vehicle, which of the following sets of gases are produced?
   A. Hydrogen and oxygen.
   B. Nitrogen and carbon dioxide.
   C. Carbon monoxide and nitrogen oxide.
   D. Nitrogen dioxide and oxygen.

3. The rate of a chemical reaction may be defined as
   A. the change in concentration of any one of the products only.
   B. the change in concentration of any one of the reactants only.
   C. the time it takes for the reaction to be completed.
   D. the change in concentration of any one of the reactants or products.

4. Why does the rate of reaction increase when powdered calcium carbonate is used instead of marble chips?

5. Which one of the following reactions will start with the highest rate?
   A. 5 g of limestone powder in 100 ml of 0.1 M HCl at 45°C.
   B. 5 g of limestone chips in 100 ml of 0.1 M HCl at 25°C.
   C. 5 g of limestone powder in 100 ml of 0.1 M HCl at 25°C.
   D. 5 g of limestone chips in 100 ml of 0.1 M HCl at 45°C.

6. The following graphs were obtained in an experiment to determine the rate of the reaction between nickel carbonate and hydrochloric acid.



   a) Write a word equation for the reaction that is taking place.
   b) Which of the graphs shows the reaction that had the most concentrated hydrochloric acid? Explain your answer.

7. Hydrogen peroxide decomposes at room temperature to give oxygen and water according to the equation:

   Hydrogen peroxide ⟶ oxygen + water

   a) Which piece of apparatus could be used to measure the rate of this reaction?
   b) What effect would adding manganese dioxide have on the rate of this reaction?
   c) What name do we give to chemicals like manganese dioxide?

8. Most of the chemical reactions that take place in living organisms involve enzymes.
   a) Explain the term ‘enzyme’.
   b) Catalase is an enzyme that is found in liver. It can cause a change in the decomposition of hydrogen peroxide. What effect would you expect if a piece of liver was placed in hydrogen peroxide solution?
   c) If manganese dioxide is added to hydrogen peroxide, the rate of decomposition reaction increases. If the mixture is heated to 80°C the reaction is more vigorous. If the same is done with a piece of liver the reaction rate is less at 80°C. Explain these observations.

Unit summary 5

- Rate of chemical reaction refers to change in the concentration of one of the reactants or products per unit time.
- The rate of reaction is inversely proportional to time taken.
- The rates of chemical reactions are affected by concentration, pressure in the case gases), temperature, surface area for solid reactants and presence of a catalyst.
- Increase in temperature increases the velocity and frequency of collisions of reacting particles. This is because of increase in kinetic energy of the reacting particles.
- The higher the concentration the higher the rate of reaction. This is because more reacting particles come in contact very frequently as a result.
- Smaller particles provide a greater surface area increasing the chances of surface contacts for particles collision hence faster reaction rate compared to bigger particles.
- A catalyst speeds up the rate of reaction that would otherwise have taken place at a much slower rate.
- Increased pressure reduces the unit volume in which gas particles collide thereby increasing chances of fruitful collisions.

Test your Competence 5

1. (a) Explain why increasing the (i) concentration (ii) temperature of reacting particles increases the rate of reaction.
   (b) Explain why decreasing the particle size of the reactants increases the rate of reaction.
   (c) Explain why the usage of catalyst increases the rate of reaction.
   (d) State whether addition of catalyst increases the amount of energy released during a chemical reaction.
2. Describe briefly how the rates of reactions below could be monitored:
   (a) Calcium carbonate with hydrochloric acid.
   (b) Sodium thiosulphate with hydrochloric acid producing sodium chloride, water, sulphur dioxide and sulphur solid.
   (c) Decomposition of hydrogen peroxide producing water and oxygen gas.

Fill the blank spaces below with an appropriate word.

3. (a) Chemical reactions can only take place when particles ________ with each other.
   (b) There are several ways to change the ________ of a reaction.
   (c) Increasing the ________ increases the ________ energy of the particles and they ________ more ________ and with greater energy.
   (d) Increasing the ________ of a reactant in solution increases the ________ of a 'fruitful' collisions between particles.
   (e) A ________ reactant will react faster if it is broken into ________ pieces. This increases the ________ , ________ and allows more ________ between the reactants.
   (f) Adding a catalyst can also speed up the rate of a reaction because it helps in providing binding surfaces for reactants bringing them closer to each other for successful collisions as a result of lowered activation energy. However, although it helps the reaction to take place much faster, a ________ is not ________ at the end of the reaction and could be ________ again.
4. A student performed an experiment to establish how effective manganese dioxide was as catalyst in the decomposition of hydrogen peroxide. The results below were obtained using two different quantities of manganese dioxide. The volume of the gas produced was recorded against time.

Time(s)	**0**	**30**	**60**	**90**	**120**	**150**	**180**	**210**
Volume (cm$^{3)}$ for 0.3g	0	29	55	79	98	115	133	146
Volume (cm^{3}) for 0.5g	0	45	84	118	145	162	174	182

   (a) Draw a diagram of the apparatus you could use to carry out the experiments.

(b) Plot a graph of these results.

(c) Is the manganese dioxide acting as a catalyst in this reaction?

(d) (i) At what stage does the reaction proceed most quickly?

(ii) In terms of particles, explain why the reaction is quickest at the point you have chosen in d (i) above.

(e) Why does the shape of the graph become less steep as the reaction proceeds?

(f) What volume of gas was produced when using 0.3 g of manganese dioxide after 50s?

(g) How long did it take for 60 cm^3 of the gas to be produced when the experiment was carried out using 0.5g of manganese dioxide?

(h) Write a balanced chemical equation for the decomposition of hydrogen peroxide.

5. A flask containing dilute hydrochloric acid was placed on a digital balance. Excess limestone chips was added to this acid, a plug of cotton wool was placed in the neck of the flask and the initial mass of the apparatus was calculated. The table below shows the results obtained.

Time (min)	**0**	**2**	**4**	**6**	**8**	**10**	**12**	**14**	**16**
Loss in mass(g)	0	2.1	3.0	3.1	3.6	3.8	4.0	4.0	4.0

(a) Plot the results of this experiment in your notebook.

(b) Which of the results would appear to be incorrect? Explain your answer.

(c) Write a balanced equation to represent the reaction that takes place.

UNIT 6

Chemical properties of acids and bases

Key unit competency

To be able to prepare and carry out reactions of acids and bases with other substances.

Learning objectives

By the end of this unit, I should be able to:

- Prepare acids and describe the chemical properties of acids.
- Prepare bases and describe the chemical properties of bases.
- Predict the products of reactions of acids and bases with other substances.
- To work with others in team spirit and develop orderliness when performing experiments.
- To observe change of colour and gas evolution during experiments and write scientific reports.
- To appreciate the uses of acids in biological processes especially in digestion of proteins in human beings and other species and use of a base in balancing the pH of soil.
- Appreciate the potential dangers associated with strong acids and bases.

Mind teaser

Look at the pictures below. Do the things in the pictures look familiar? Name them.

Fig 6.1

What are the things made of ? Name other substances found at home made of acids and bases. Come up with a list of uses of acids and bases based on your findings above.

Introduction

The word 'acid' is commonly used in everyday life. It is regarded as a highly corrosive, substance with sour taste. On the other hand bases are bitter in taste and have soapy feel. Acids and bases react differently with different substances. The reaction between an acid and a base produces salt and water only and is known as **neutralisation** reaction.

6.1 Preparation of acids and bases

(a) Displacing a volatile acid by a less volatile one

Activity 6.1

Apparatus and reagents

Sodium chloride, concentrated sulphuric acid, round bottomed flask, thistle funnel, tripod stand, Bunsen burner, wire gauze, wash bottle, delivery tubes, gas jar, rubber bung.

Procedure

1. Place sodium chloride in a round bottomed flask.
2. Set up the apparatus as follows.

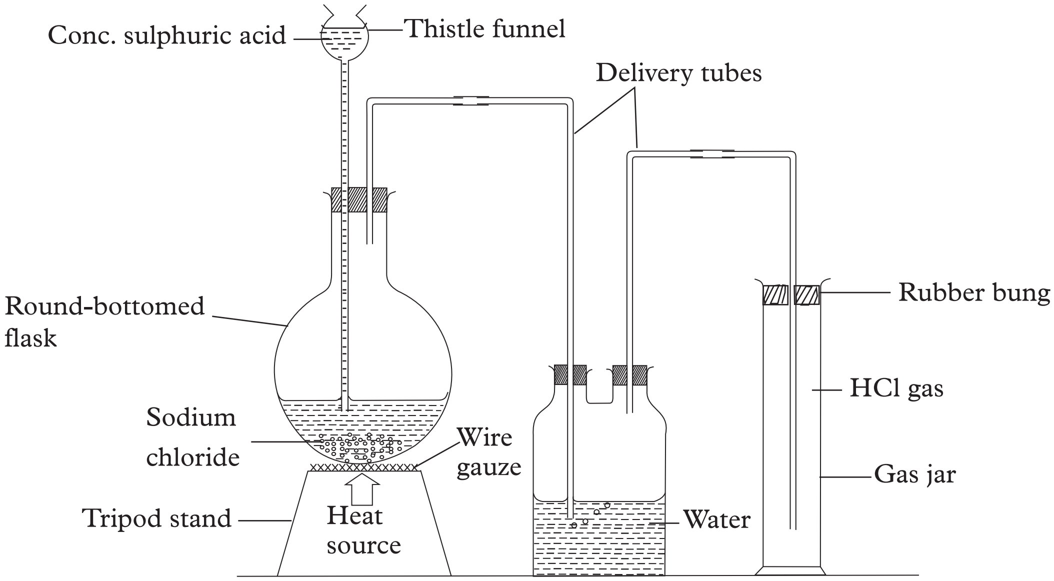


Fig 6.2: Laboratory preparation of hydrogen chloride gas

3. Slowly add concentrated sulphuric acid to the sodium chloride. What do you observe?
4. Heat the mixture gently.

Study questions

1. What do you observe when a mixture of sodium chloride and concentrated sulphuric acid is heated gently?
2. Write an equation for the reaction that occurs in the flask.
3. Why is an inverted funnel necessary in this experiment?

Discussion corner!

1. Work in your study groups to discuss the observations made during the experiment and the study question above.
2. Make a report and present it to the class.

The facts

When the mixture of sodium chloride and concentrated sulphuric acid is heated, effervescence occurs and a colourless gas is produced. This gas is hydrogen chloride.

$$NaCl(s) + H_2SO_4(l) \longrightarrow NaHSO_4(aq) + HCl(g)$$

In this reaction, the less volatile concentrated sulphuric acid displaces the more volatile hydrochloric acid from sodium chloride.

Hydrogen chloride gas is highly soluble in water hence it is dissolved in distilled water using an inverted funnel arrangement to avoid "sucking back". A solution of hydrogen chloride gas in water is hydrochloric acid. Enough hydrogen chloride has to be dissolved in water to give a more concentrated hydrochloric acid.

The less volatile sulphuric acid also displaces the more volatile nitric acid from nitrates (sodium nitrate or potassium nitrate) according to the following equation.

$$NaNO_3(s) + H_2SO_4(l) \longrightarrow NaHSO_4(aq) + HNO_3(aq)$$

In this preparation, all glass apparatus is used since nitric acid attacks cork and rubber fittings. The nitric acid vapour is then condensed to give nitric acid.

b) Preparation of a base by addition of water to soluble metal oxide

Activity 6.2

Apparatus and reagents

Sodium oxide, distilled water, stirrer, beaker, conical flask or beaker.

Procedure

1. Place some sodium oxide in a conical flask.

2. Add distilled water to the sodium oxide and stir the mixture. What do you observe?
3. Test the resulting solution using blue and red litmus paper. What do you observe?

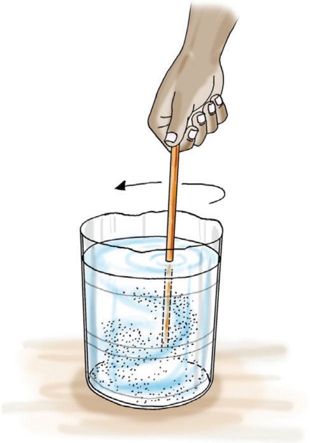


Fig 6.3: Stirring sodium oxide and water mixture

Study question

1. What is the purpose of stirring?
2. Do you think all bases can be prepared using the oxides of their metals? Explain.
3. What do you expect on the effect of litmus paper?

Discussion corner!

1. In your study groups discuss the observations in the experiment and the study questions above.
2. Present your findings or answers to the class.

The facts

Sodium oxide is a white solid formed when sodium metal burns in air. On addition of water, the sodium oxide dissolves forming a colourless solution. The solution turns red litmus paper blue showing it is an alkali.

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

When potassium oxide is dissolved in water, it also forms an alkali called potassium hydroxide.

$K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$

Self-evaluation Test 6.1

1. a) Describe how hydrochloric acid can be prepared in the laboratory.
   b) Nitric acid can be prepared in the laboratory by heating a mixture of potassium nitrate and concentrated sulphuric acid.
      i) What type of reaction occurs?
      ii) Which of the two acids mentioned is more volatile? Explain.
2. Starting with sodium oxide, describe how a solution of sodium hydroxide can be prepared in the laboratory.

6.2 Reaction between water and sodium metal

Activity 6.3

Apparatus and reagents

Sodium metal, trough, water, splint, Bunsen burner, boiling tube, aluminium foil, knife, universal indicator, a pair of tongs.

Procedure

1. Place water in a trough up to half its volume.
2. Add three drops of universal indicator solution to the water.
3. Cut a small piece of sodium (size of rice grain) and drop into the water in the trough. Record all the observations.

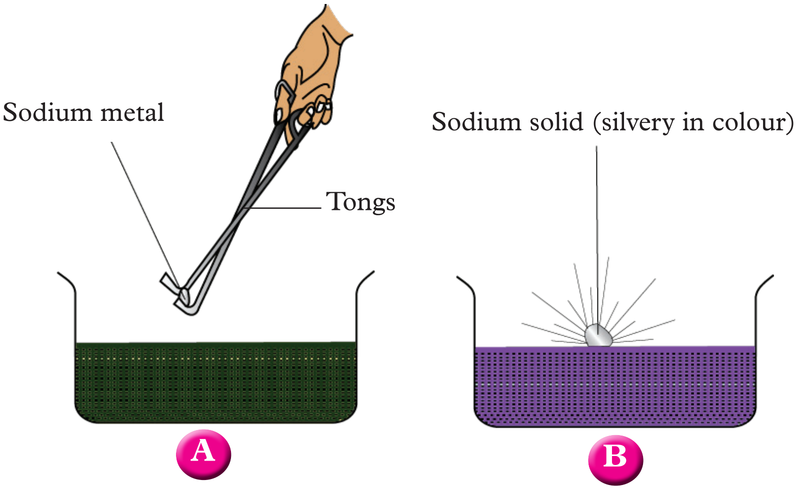


Fig 6.4: Reaction of sodium with water

4. Wrap another piece of sodium in wire gauze. Place it in a trough of water and insert on it a test tube full of water as shown. Collect the gas in the test tube.

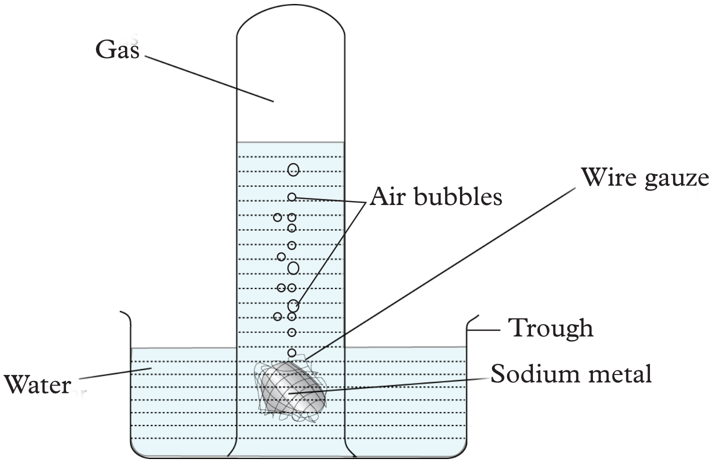


Fig 6.5: Collecting produced gas

5. Test the collected gas in the test tube with a burning splint. Record your observations.

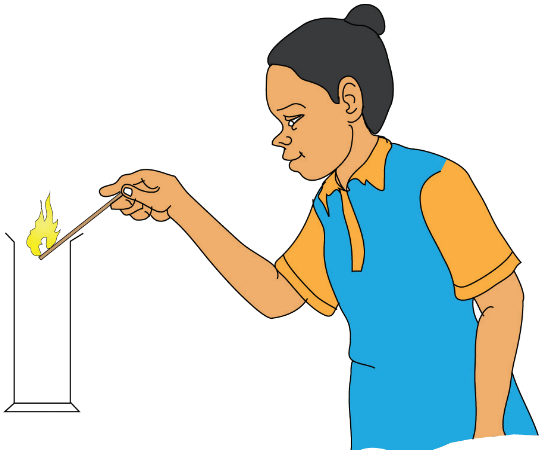


Fig 6.6: Testing for collected gas

Study questions

1. What observations are made when sodium metal is dropped in water?
2. Write the equation for the reaction between sodium metal and water.
3. What happens to the burning splint inserted in a test tube containing the gas collected?
4. Suggest the properties of the gas.

Discussion corner!

In pairs discuss the observations and study questions above and present your findings in the class.

The facts

When sodium metal is dropped in water, the following observations are made.

- Sodium darts around on the surface of the water (floats on water).
- A colourless gas is given off.
- Universal indicator turns purple.

Sodium floats on water because it is less dense than water. It reacts with water to produce a gas that burns with a 'pop' sound. The gas is hydrogen. It darts on the water surface since the hydrogen gas evolved is trying to escape from all directions. A hissing sound is produced due to the heat generated during the reaction.

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

The universal indicator turns purple indicating the solution formed is alkaline (sodium hydroxide). The confirmatory test for hydrogen gas is that it burns with a 'pop' sound when a burning splint is introduced into the gas.

Self-evaluation Test 6.2

1. How is sodium metal stored in the laboratory? Explain.
2. A small piece of sodium metal was dropped in a trough of water, state and explain the observations made.
3. Write an equation for the reaction that occurs in(2) above.
4. Describe a test for hydrogen gas.

6.3 Reactions of acids and bases

a) Reactions of acids

Activity 6.4

Apparatus and reagents

Dilute hydrochloric acid, magnesium metal, calcium carbonate, sodium hydrogen carbonate, magnesium oxide, sodium hydroxide, phenolphthalein indicator, test tubes, splint, Bunsen burner and beakers.

Procedure

1. Place 5 cm^3 of dilute hydrochloric acid in a test tube. To the acid, add a piece of magnesium ribbon. Record your observations. Introduce a burning splint in the test tube. What do you observe?
2. Place 5 cm^3 of dilute hydrochloric acid into a test tube. Add half spatula endful of calcium carbonate. What do you observe? Insert a burning splint into the test tube where the reaction is occurring. Record your observations.

3. Repeat procedure 2 using sodium hydrogen carbonate in place of calcium carbonate.
4. To 5 cm^3 of dilute hydrochloric acid in a test tube, add magnesium oxide and heat. What do you observe?
5. Place 5 cm^3 of sodium hydroxide in a test tube and add 2 drops of phenolphthalein indicator. What do you observe? To this solution, add dilute hydrochloric acid slowly until the colour of the solution just turns colourless.

Study questions

1. What is the purpose of the burning splint in step 1 of the procedure above?
2. Why was there heating in step 4 of the procedure?
3. How do you expect the changes in phenolphthalein?
4. What would happen if methyl orange was used in place of phenolphthalein?
5. Suggest the equations for the reaction in step 1-5 of the procedure?
6. Describe all the possible observations in the procedure above.

Discussion corner!

1. In pairs discuss the observations of the experiments and the study questions above. Write your findings including the equations in a report to be presented to the entire class.

The facts

1. When a piece of magnesium ribbon is dropped in dilute hydrochloric acid, effervescence occurs and a colourless gas that burns with a 'pop' sound is evolved. The gas is hydrogen.

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

   A similar reaction occurs when dilute sulphuric acid is reacted with magnesium metal.

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

   Generally metals above hydrogen in the reactivity series react with dilute acids to liberate hydrogen.

   NB: If nitric acid is used, the hydrogen liberated is oxidised to water since nitric acid is a powerful oxidising agent. Generally when a metal reacts with a dilute acid, the products are salt and hydrogen gas.

   Metal + dilute acid ⟶ salt + hydrogen gas
2. When calcium carbonate and sodium hydrogen carbonate are added to dilute hydrochloric acid in separate test tubes, effervescence occurs and a colourless gas is liberated. The gas puts off a burning splint. The gas is carbon dioxide.

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

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

Generally, dilute acids react with carbonates to form salt, carbon dioxide and water.

Carbonate + dilute acid ⟶ salt + carbon dioxide + water

However, the reaction between dilute sulphuric acid and calcium carbonate starts and then soon stops. This is due to the formation of insoluble calcium sulphate which coats the carbonate preventing any further contact between the reactants hence the reaction stops.

- Explain what happens when dilute hydrochloric acid is reacted with lead (II) carbonate.

3. When a mixture of magnesium oxide and hydrochloric acid is heated, the magnesium oxide dissolves reacting with the acid to form magnesium chloride and water. If universal indicator is placed in the dilute hydrochloric acid, it turns red, if magnesium oxide is continuously added a little at a time with heating, the indicator eventually turns green showing the solution has become neutral. This is a neutralisation reaction.

   $MgO(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$

   Generally when an acid reacts with a metal oxide, salt and water are formed.

   Dilute acid + metal oxide ⟶ salt + water.

   Dilute sulphuric acid when warmed with copper (II) oxide, the black copper (II) oxide reacts with the acid to form blue copper (II) sulphate and water.

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

   Note: The acid used determine the type of salt formed. Hydrochloric acid gives chlorides, sulphuric acid gives sulphates whereas nitric acid gives nitrates.

4. When dilute hydrochloric acid was slowly added to sodium hydroxide solution to which phenolphthalein indicator was added, the colour of the solution turns from pink to colourless at some point. Dilute hydrochloric acid reacts with sodium hydroxide to form a salt, sodium chloride, and water only.

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

   This is a neutralisation reaction. Generally acids react with alkalis to form salt and water only. This is called **neutralisation reaction**.

   Dilute acid + alkali ⟶ salt + water

   Dilute sulphuric acid neutralises potassium hydroxide solution to form potassium sulphate and water.

   $2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$

Self-evaluation Test 6.3

1. Copy and complete the reaction below.
   (a) Magnesium carbonate + hydrochloric acid ⟶ ____+_____+______
   (b) Sodium carbonate + nitric acid ⟶ ____+_____+______
   (c) Calcium carbonate + sulphuric acid ⟶ ____+_____+______
   (d) Potassium carbonate + hydrochloric acid ⟶ ____+_____+______
2. When acid is added to an alkali, what products are formed?
3. Complete the chemical equation $K_2O + HCl \longrightarrow$
4. What is the product of the reaction between nitric acid and ammonia?

b) Reaction of bases

Activity 6.5

Apparatus and reagents

Sodium hydroxide, burette, teat pipette, dilute hydrochloric acid, pipette, conical flask, methyl orange indicator.

Procedure

1. Pipette 25cm^3 of 1M sodium hydroxide solution into a conical flask and add three drops of methyl orange indicator. What do you observe?
2. Run 1M hydrochloric acid to the sodium hydroxide solution until the end point. How do you tell when the end point is reached?

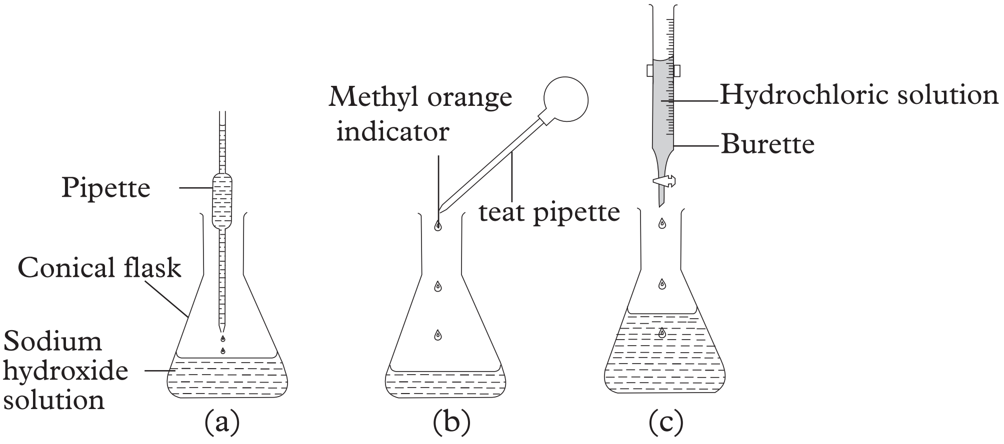


Fig 6.7: Experimental setup for the reaction between sodium hydroxide and dilute hydrochloric acid

Study questions

1. What name is given to the reaction that occurs?
2. Why was methyl orange added?

3. Is the concentration the same? Explain.

Discussion corner!

1. Work in your study groups and discuss the observations of the experiment and the study questions above.
2. Present your findings to the class.

The facts

When methyl orange indicator is added to sodium hydroxide solution, the solution turns yellow. On addition of hydrochloric acid, neutralisation takes place and sodium chloride and water are formed.

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

The end point is reached when one drop of the hydrochloric acid changes the colour of the solution from yellow to orange. Some bases such as zinc oxide, lead (II) oxide and aluminium oxide react with bases to form complex salts.

$$ZnO(s) + 2NaOH(aq) + H_2O(l) \longrightarrow Na_2Zn(OH)_4(aq)$$

$$PbO(s) + 2NaOH(g) + H_2O(l) \longrightarrow Na_2Pb(OH)_4(aq)$$

$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \longrightarrow 2NaAl(OH)_4(aq)$$

These oxides which react with both alkalis and acids are known as **amphoteric oxides**.

Self-evaluation Test 6.4

1. Give four chemical properties of acids.
2 a) What is neutralisation?
  b) Write an equation to show the reaction between calcium oxide and dilute hydrochloric acid.
3 a) State one chemical property of bases.
  b) Potassium hydroxide is neutralised by dilute sulphuric acid. Write an equation for the reaction that takes place.
4 a) What is an amphoteric oxide?
  b) Name two amphoteric oxides.

6.4 Uses and dangers of acids and bases

Uses of acids and bases

Activity 6.6

1. State some applications of acids and bases in our daily life.
2. Explain the role of hydrochloric acid in the human stomach.
3. What are some of the dangers paused by acids and bases to the environment.

Discussion corner!

In pairs discuss the questions above. Write a report and present it to the class.

The facts

a) Acids find a variety of uses in our daily lives

1. Tannic acid found in the bark of certain trees is used in processing leather. The gelatin in untanned animal skins become insoluble after the skin is treated with tannin. The process is called **tanning**.
2. Hydrochloric acid is naturally secreted in the human stomach and is a component of the gastric juice. The role of hydrochloric acid in the stomach along with other gastric juices is to provide suitable conditions for enzymes involved in food digestion. Hydrochloric acid makes up about 5% of gastric juice. During protein digestion in the human stomach, hydrochloric acid activates pepsinogen and convert it to the pepsin. Pepsin then begins to break the peptide bonds in proteins forming smaller peptide chains. Smaller peptide chains formed with pepsin leave the stomach and enter the small intestine where they are further broken down by enzymes secreted by pancreases.
3. Sulphuric acid is used in lead –acid accumulators (batteries) for motor vehicles and solar batteries.
4. Ascorbic acid found in citrus fruits is an important source of vitamin C for the body.
5. Ethanoic acid is the main component of vinegar which is used in food seasoning.
6. Acids are also used in the making of fertilizers such as ammonium nitrate, ammonium sulphate among others.

b) Some of the uses of bases

1. Ammonia is used in production of fertilisers such as ammonium sulphate. It is also used in household cleaning agents.
2. Aluminium hydroxide is used in the manufacture of other aluminium compounds and to make gastric medicine (anti – acid).

3. Calcium hydroxide is used in making cement, neutralise excess acidity in the soil and sewage treatment.
4. Sodium hydroxide is used in manufacture of soaps, detergents and other cleaning agents.
5. Magnesium hydroxide is used to make anti acid tablets. It is also used as an anti-perspirant deodorant.
6. Calicum oxide is also used to reduce acidity of soils.

Dangers of acids and bases

In senior 1, you learnt about dangers of acids and bases in our daily lives. We learnt that acids and bases cause a number of undesirable effects in the environment.

1. Acid rain

Some industrial processes release acidic gases such as sulphur dioxide, carbon dioxide and oxides of nitrogen into the atmosphere. When these gases accumulate in the atmosphere, they combine with rain water making it acidic. This is called acid rain. Acid rain affects the environment in a number of ways.

- Acid rain peels paints off building walls and structures made of marble (limestone). It also corrodes metallic structures and statues.
- When acid rain through surface run off ends in rivers and lake, it eventually lead to the death of fish and other aquatic species.
- Acid rain causes important nutrients to leach out of the soil. This causes trees and other plants to be deprived of the nutrients.
- Due to acid rain, aluminium ions are free from clay as aluminium sulphate resulting in damage of tree roots. The trees are unable to draw enough water through the damaged roots leading to wilting.

2. Tooth decay

Bacteria convert foods that stick between the teeth into acids. These acids dissolve the enamel leading to teeth decay; that eventually causes tooth aches. Tooth decay starts when the pH in the mouth falls below 5.8.

Note: Tooth pastes are basic in nature to help neutralise the acids in the mouth.

3. Soil pH

The acidity or alkalinity of the soil affect plants growth. The level of acidity is determined by the amount of hydrogen ions in the soil. When the concentration of hydrogen ions (H^+) is very high, the soil is described as being acidic. On the other hand when concentration of hydrogen ions is low, the soil is said to be basic. Soil acidity is caused by excessive use of fertilisers and acid rain. Acidic soil is not suitable for the healthy growth of many crops.

Note: If the soil is too acidic, it is treated (corrected) by addition of lime. Lime is a mixture of calcium oxide and calcium hydroxide. This mixture neutralises the effects of excess acid in the soil.

4. Stomach indigestion

Dilute hydrochloric acid secreted in our stomachs help in food digestion. When the stomach secretes excess acid than is necessary, indigestion occurs. The excess acid burns the stomach walls causing ulcers. Excessive food consumption or too much stress leads to excessive acid secretion in the stomach.

Note: Anti-acid tablets contain mild bases such as magnesium hydroxide. This neutralizes excess acidity in the stomach hence relieving the pain.

Self-evaluation Test 6.5

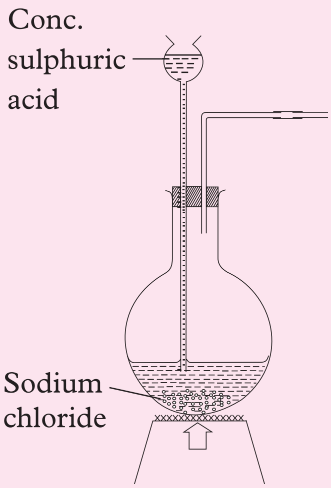
1. a) State the uses of some acids in our daily life.
   b) State the uses of some bases in our daily life.
2. Give three dangers caused by acids and bases

Unit summary 6

- Acid is any substance that tastes sour, changes the colour of certain indicators (e.g. reddens blue litmus paper), reacts with some metals (e.g. iron) to liberate hydrogen, reacts with bases to form salts, and promotes certain chemical reactions.
- Bases are substances that are slippery to the touch, taste astringent, change the colour of indicators (e.g. turn red litmus paper blue), react with acids to form salts, promote certain chemical reactions.
- An acid may be prepared in the laboratory through displacement of a volatile acid by a less volatile one.
- A base may be prepared in the laboratory through the addition of water to a soluble metal oxide.
- Generally, acids react with alkali to form salt and water only. This type of reaction is known as neutralisation reaction.
- Oxides that react with both alkalis and acids are known as amphoteric oxides.
- Acids are used in lead –acid batteries, making of fertilisers, in tanning factories, is found in the human stomach where it helps during digestion, among other use.
- Bases are used as fertilisers, in the manufacture of anti-acids, cement, soaps and detergents among others.
- Acids and bases can cause undesirable effects on the environment such as acid rain, tooth decay, soil pH change and stomach indigestion.

Test your Competence 6

1. a) Name the method used to prepare hydrochloric acid in the laboratory.
   b) The following diagram gives an incomplete set up of the apparatus that can be used to prepare a sample of hydrochloric acid in the laboratory. Study it and answer the questions that follow.

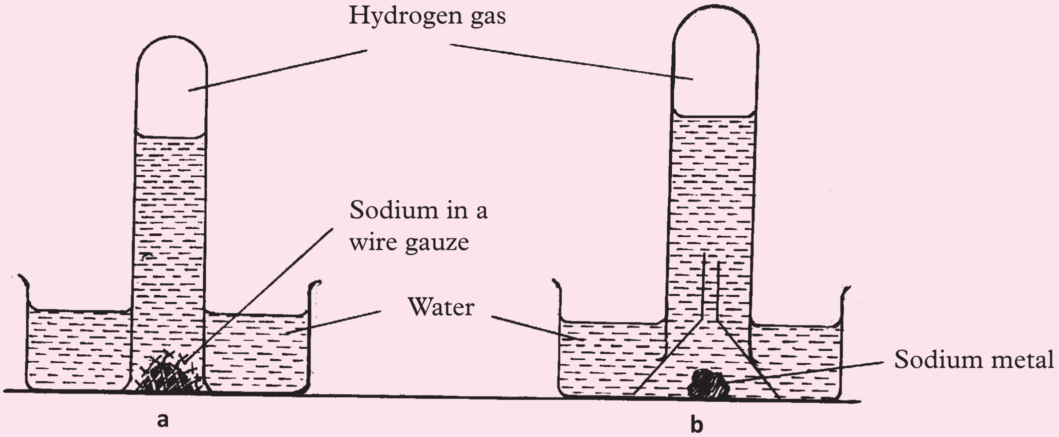


   i) Complete the diagram to show how a sample of hydrochloric acid may be collected.
   ii) Write an equation for the reaction that occurs.
   iii) Name another substance which can be used in place of sodium chloride.

   c) Dilute hydrochloric acid was added to zinc metal.
   i) State and explain the observations made.
   ii) Describe a test that can be used to confirm the gaseous product.
   iii) Write an equation for the reaction that occurs.

2. Sodium oxide was added to 5cm^3 of water in a boiling tube.
   (a) (i) What observations were made?
   (ii) Name the product formed in this reaction.
   (iii) Write an equation for the reaction that occurs.
   (b) A sample of the solution formed was placed in a test tube. 1 drop of methyl orange indicator was added; followed by addition of hydrochloric acid drop by drop until a change in colour just occurred.
   (i) What type of reaction took place?
   (ii) State two observations made on addition of methyl orange indicator up to the colour change.
   (iii) Write an equation for the reaction that occurs.

3. Sodium metal was placed in water in a trough.



i) State and explain any three observations made.

ii) Write an equation for the reaction that occurs.

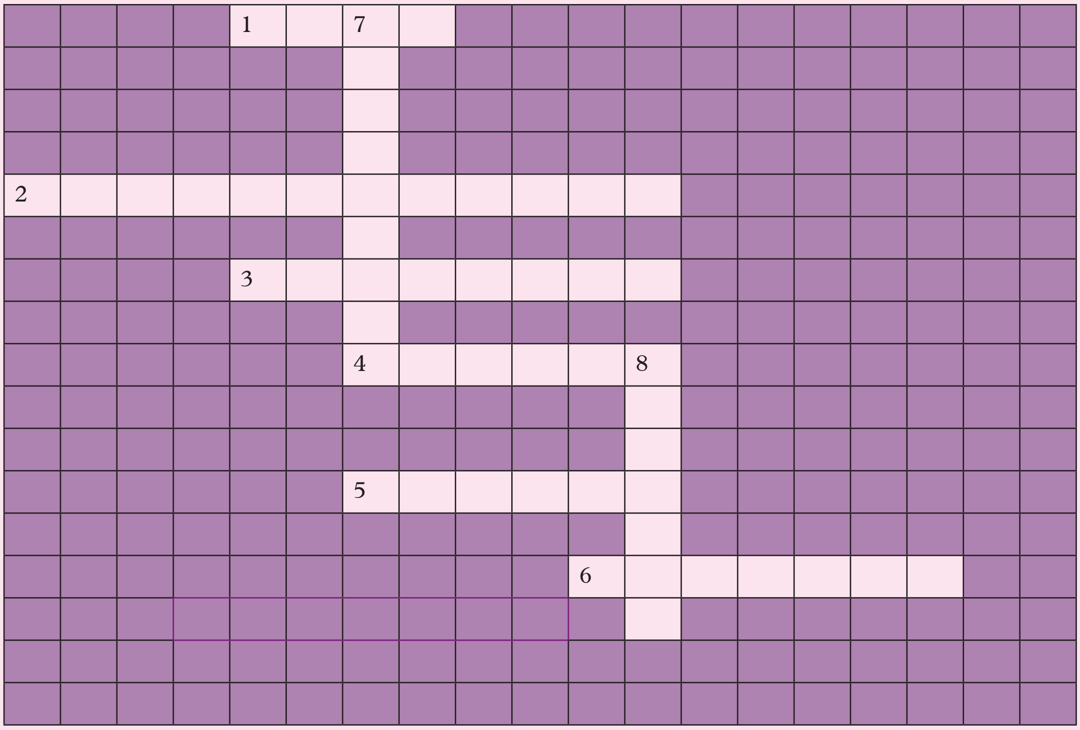
iii) If lead (II) oxide was added to a sample of the solution formed, write an equation for the reaction that occurs.

4. a) What is acid rain?

b) Describe the formation of acid rain.

c) State two environmental effects of acid rain.

5. Use the questions that follow to fill the crossword puzzle below.



Across

1. Feels slippery or soapy to touch.
2. Acid found in stomach of human beings.
3. Acid present in fizzy drinks.
4. A lemon contains _______ acid.
5. It turns red in acid and blue in base.
6. A sour tasting solution containing acetic acid.

Down

7. Acid found in car battery.
8. Common name for sodium hydroxide is ______ soda.

6. State and explain the observations made when dilute hydrochloric acid is added to lead (II) oxide in a test tube.
7. State the uses of the following bases:
   a) Magnesium hydroxide
   b) Sodium hydroxide
8. a) What is the role of hydrochloric acid in the human stomach?
   b) (i) What is indigestion?
      (ii) What causes indigestion?
      (iii) Explain how indigestion is treated.
9. The substance whose aqueous solution change the blue litmus paper to red is ___________.
   a) acid
   b) base
   c) neutral solution
   d) salt
10. A substance which has the tendency of loosing one or more protons is called_____.
11. When equal amounts of acid and base are mixed together, salt and water are formed. This reaction is known as___________.

UNIT 7

Concentration of solutions

Key unit competency

To be able to determine the concentrations of solutions from data obtained by simple acid-base titrations.

Learning objectives

By the end of this unit, I should be able to:

- Explain the difference between end point and equivalence point.
- Describe the process of titration and perform simple acid-base titrations.
- Carry out experiments in groups to prepare solutions of different concentrations.
- Use data obtained from the titration of an acid and a base to calculate the concentration of either acid or base.
- Appreciate the importance of accuracy in measurements.
- Develop a culture of working in a team.
- Appreciate the importance of repeating procedures to obtain more accurate results.
- Develop self confidence during presentations and discussion in groups.

Mind teaser

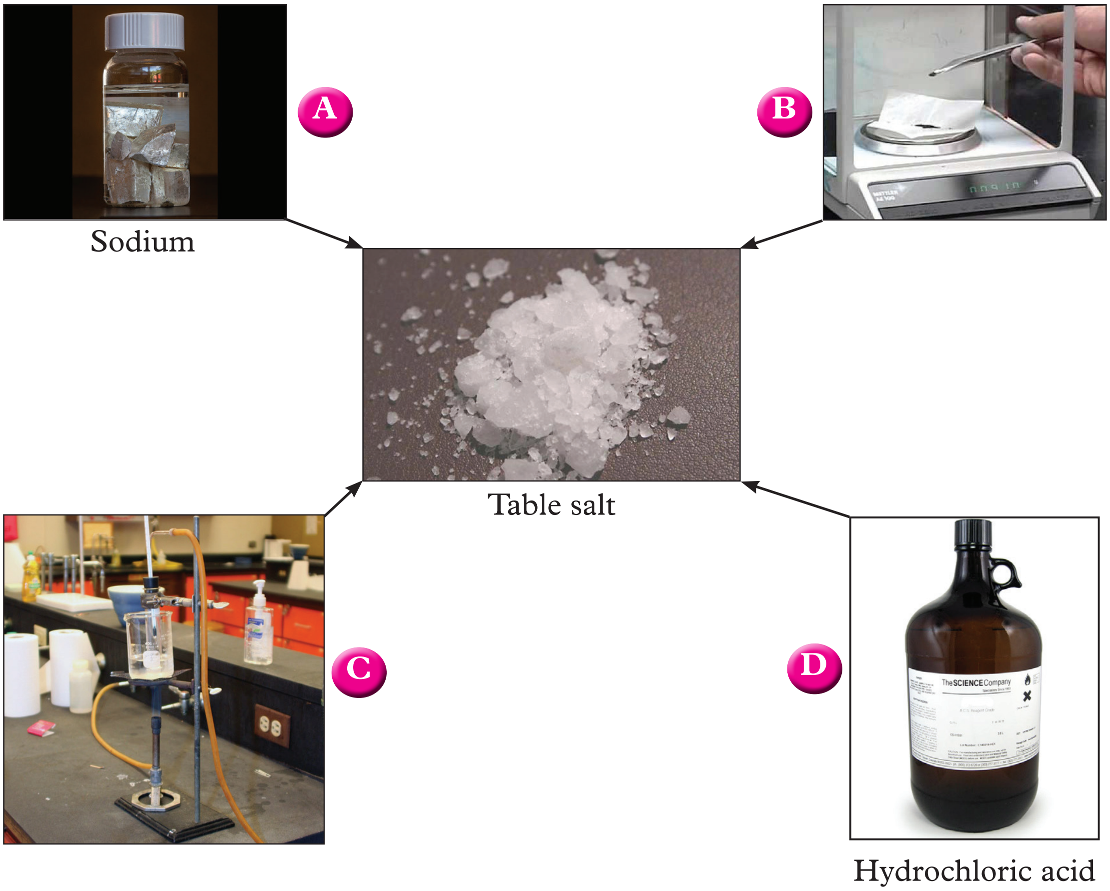


Fig 7.1

What is the relationship between the items in the photos? Put the activities in the circled letters into two categories that is raw materials and processes. Identify each.

Introduction

Some experiments in Chemistry require that we use solutions whose concentrations are known. They are called **standard solutions**. A standard solution is prepared by accurately measuring the mass of solute and dissolving it in a known volume of solvent to make the required volume of the solution. Using various formulae, the concentration of such solutions can be calculated. For example, when an acid reacts with a base, the reaction continues and stops after some time. The point at which the reaction stops is called **end point**. If we know the concentration of the acid, we can use the reacting volume to determine the concentration of the base. Titration is a process that can be used to determine the concentration of solutions. **Equivalent point** in any chemical reaction is reached when the amount of each chemical needed to react is present. At this point, more of the reactants is in excess. It should not be confused with the **end point**.

7.1 Concept of concentration

Determining concentration of solutions

Activity 7.1

Materials

Sugar, teaspoons, 3 cups, water.

Procedure

1. Put equal amounts of drinking or potable water in each of the three cups and label them A, B and C.
2. Add sugar as follows:
    - In cup A, add half teaspoonful of sugar.
    - In cup B, add one spoonful of sugar.
    - In cup C, add two spoonfuls of sugar.
3. Stir and taste each sugar solutions prepared in step 2 above.

Study questions

1. In which cup was the solution sweetest?
2. How can we tell the cup with a lot of sugar.
3. How can we decrease the taste of sugar in each cup?
4. What would happen when the water used was warmed first before step 2 in the procedure above?
5. How can concentrations of solutions be measured?

I have discovered that…

Cup C had the sweetest taste because it contained most sugar particles. By adding more water that is diluting the solution, we reduce the taste of sugar.

The facts

Many chemical reactions are performed with reagents in solution. It is necessary to know the amount of solute dissolved in a given amount of solution. If a large volume of water is added to a small amount of solute, the solution is said to be **dilute**. Similarly if the solution contains a larger amount of solute, the solution is said to be **concentrated.**

Concentration of solutions

The term concentration refers to the *amount of solute dissolved in a specific volume of solvent*. In a given amount of water, the more solute dissolved, the more concentrated the solution. If the solution contains a little solute, it is described as a dilute solution.

Similarly, if the solution contains large amount of solute, then the solution becomes more concentrated.

How concentrations are expressed

In consumer and industrial world, the most common method of expressing the concentration is based on the quantity of solute in a fixed quantity of solution. The quantities referred to here can be expressed in weight (w), volume (v) or both (that is the weight of a solute in a given volume of solution). In order to distinguish among these possibilities the abbreviations (w/w, v/v and w/v) are used. There are a number of ways to express the relative amounts of solute and solvent in a solution.

Units of concentration

Once you have identified the solute and solvent in a solution, you are ready to determine its concentration. Concentration may be expressed in several ways, using percent composition by mass, volume percent, mole fraction, molarity, molality or normality.

a) Percent composition (by mass)

We can consider percent by mass (or weight percent, as it is sometimes called) in two ways:

- The parts of solute per 100 parts of solution.
- The fraction of a solute in a solution multiplied by 100.

You need two pieces of information to calculate the percent by mass of a solute in a solution:

- The mass of the solute in the solution.
- The mass of the solution.

Percent by mass is the mass of the solute divided by the mass of the solution (mass of solute plus mass of solvent), multiplied by 100.

Use the following equation to calculate percent by mass:

$$\text{Percent mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

Worked example7.1

Determine the percent composition by mass of a 100 g salt solution which contains 20 g salt.

Solution

$$\frac{20\text{g}}{100\text{g solution}} \times 100\% = 20\%$$

b) Molarity (M)

Molarity (M) tells us the number of moles of a solute in exactly one litre of a solution.

Note: Molarity is spelled with an "r" and is represented by a capital M.

We need two pieces of information to calculate the molarity of a solute in a solution:

- The moles of solute present in the solution.
- The volume of solution (in litres) containing the solute.

To calculate molarity we use the equation:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in litres}}$$

Molarity is probably the most commonly used unit of concentration. It is the number of moles of solute per litre of solution (not necessarily the same as the volume of solvent).

Worked example 7.2

What is the molarity of a solution made when water is added to 11g calcium chloride($CaCl_2$) to make 100 ml of solution? (Ca=20, Cl=35)

Solution

RMM of $CaCl_2$ = 40 + (35 × 2)=110

$$\text{Moles} = \frac{\text{mass of } CaCl_2}{\text{RMM } CaCl_2} \text{ given } = \frac{11}{110} = 0.10 \text{ moles}$$

$110 \text{ g} \longrightarrow 1 \text{ mole}$

$11 \text{ g} \longrightarrow x \text{ mole}$

$$x = \frac{11 \times 1}{110} = 0.1 \text{ mole}$$

$100 \text{ ml} \longrightarrow 0.1 \text{ mole}$

$1000 \text{ ml} \longrightarrow x$

$$x = \frac{1000 \times 0.1}{100} = 1 \text{ M}$$

$\therefore$ molarity = 1 M

c) Volume percent (%)

Volume percent (or volume/volume percent) is most often used when preparing solutions of liquids. Volume percent is given by the formula:

$$\% = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$$

Note: Volume percent is relative to volume of solution, not volume of solvent. For example, if wine is 12% ethanol, it means that there are 12 ml of ethanol for every 100 ml of wine solution. In addittion it is important to realise that liquid and gas volumes are not necessarily additive. For example, if you mix 12 cm^3 of ethanol with 100 ml of wine, you will get a solution which is less than 112 cm^3 in volume.

Self-evaluation Test 7.1

1. A normal saline solution used in medicine for wound cleaning is 0.91% (w/v) solution of sodium chloride (NaCl) in water. How would you prepare 1.5 litres of this solution?
2. Describe how you would prepare 30 g of 20 %(w/w) solution of potassium chloride (KCl) in water.
3. How would you make 120 cm^3 of 0.10 M potassium hydroxide (KOH) in water?

7.2 Preparation of solutions

How to prepare molar solutions

Activity 7.2

To prepare 1 mol dm^{-3} sodium chloride solution (1M NaCl solution).

Apparatus and chemicals

- Weighing balance
- 1 dm^3 volumetric flask
- Stirrer/glass rod
- Wash bottle
- Distilled water
- Beaker
- Funnel
- Sodium chloride

Procedure

1. Weigh 58.5 g of sodium chloride crystals.
2. Place 400 cm^3 of distilled water in a beaker.
3. Add a little salt into the water, stirring continuously until all the salt dissolves.
4. Using a filter funnel transfer the salt solution into a 1 dm^3 volumetric flask Fig. 7.2(b).
5. Rinse the beaker with distilled water carefully and add the washing into the flask.
6. Add more distilled water and shake the flask well. Add more water until the solution level is just below the calibration mark. Using a wash bottle add more distilled water drop by drop until the bottom of the meniscus is at the same level with calibration mark Fig. 7.2(c).
7. Use a stopper to cover the flask and invert it several times to make sure the solution mixes thoroughly.
8. Transfer the solution into a reagent bottle and label it 1M NaCl, see Fig. 7.2 (d).

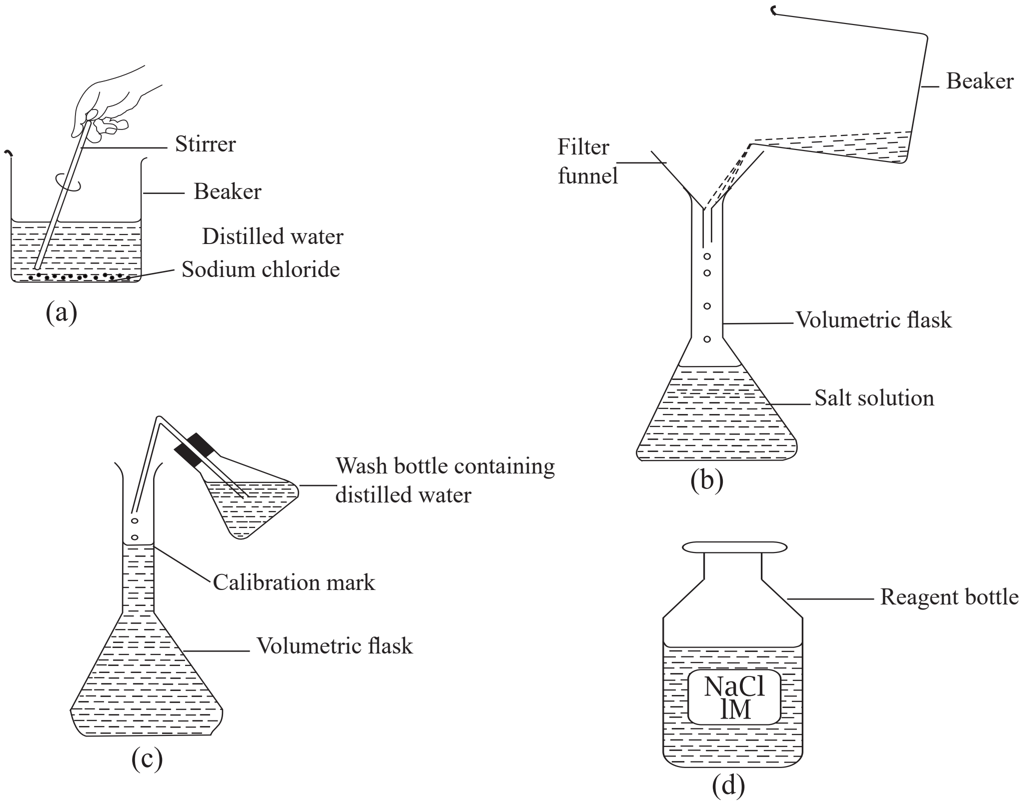


Fig 7.2: How to make 1M solution of sodium chloride

Study questions

1. Did all the salt dissolve in the water?
2. Write the formula of sodium chloride.
3. Using the relative atomic masses, find the molar mass of sodium chloride.
4. What conclusion can we make from this experiment?

The term concentration refers to the amount of solute dissolved in a specific volume of solvent. In a given amount of water, the more solute dissolved, the more concentrated the solution. If the solution contains a little solute, it is described as a dilute solution. Similarly, if a large volume of water is added to a small amount of solute, then the solution is dilute.

Molar solutions

The concentration of a solution is usually expressed in number of moles dissolved in 1000 cm^3, that is 1 cubic decimetre (dm^3) of solution.
1000 cm^3 = 1 dm^3 = 1 litre.

Note: The total volume of the solute plus the solvent must be equal to 1000 cm^3. This is why a 1 dm^3 volumetric flask is used.
When one mole of a solute is dissolved in solvent(water) and the volume of

the solution is made up to 1 dm^3 the solution is said to be a molar solution. A concentration of one mole per cubic decimetre is often written **as 1 mol/dm^3** or **1 mol dm^{-3}** or 1M. Concentration expressed this way is sometimes referred to as the molarity of a solution, abbreviated as M. Therefore, a molar solution contains 1 mole of solute per dm^3 solution.

A solution containing 2 moles in 1000 cm^3 (1 dm^3), has a concentration of 2 mol dm^{-3} or simply 2 M. A solution which contains 0.5 moles of solute dissolved in 1 dm^3 solution would be written as 0.5 mol dm^{-3} or 0.5 M. This shows the concentration is known. A solution whose concentration is known is called a **standard solution**.

I have discovered that…

To prepare 1M solution of sodium chloride, 58.5g of NaCl is added to water and the volume made up to 1 litre.

The facts

Finding the molar mass of sodium chloride

One mole of sodium chloride is 58.5 g.

Therefore when we dissolve this amount of salt in distilled water and the volume is made up to 1000 cm^3 of solution, the molarity or molar concentration of the solution is 1 mol dm^{-3} or 1M.

Suppose we did not wish to prepare so much solution that is 1000 cm^3 (1 dm^3), we needed only 100 cm^3, how much salt should we dissolve to make 1M solution? This is what we need to understand.

1M means that 1 mole of a substance, in this case salt, is contained in 1 dm^3 (1000 cm^3).

But 1 mole of NaCl = 58.5 g

It means 58.5 g were contained in 1000 cm^3 of solution

$\therefore$ x g are dissolved in 100 cm^3

Cross multiply and solve for x

$$\therefore x = \frac{(58.5 \text{ x } 100 \text{ g})'}{1000}$$

$$= 5.85 \text{ g}$$

Therefore for us to make 1M NaCl of 100 cm^3; we add 5.85 g of NaCl in distilled water and the volume made up to 100 cm^3.

How to calculate dilutions

A solution is diluted whenever a solvent is added to the solution. Adding solvent results in a solution of lower concentration. You can calculate the concentration of a solution following a dilution by applying this equation:

$M_iV_i = M_fV_f$

Where M is molarity, V is volume, and the subscripts i and f refer to the initial and final values.

Worked example 7.3

How many milliliters of 5.5 M NaOH are needed to prepare 300 cm³ of 1.2 M NaOH?

Solution

$5.5\ \text{M} \times V_1 = 1.2\ \text{M} \times 0.3\ \text{L}$

$$V_1 = 1.2\ \text{M} \times \frac{0.3\ \text{L}}{5.5\ \text{M}}$$

$V_1 = 0.065\ \text{L}$

$V_1 = 65\ \text{ml}$

To prepare the 1.2 M NaOH solution, you pour 65 cm³ of 5.5 M NaOH into your container and add water up to 300 cm³ final volume.

Note: Most laboratories keep stock solutions of common or frequently used solutions of high concentration. These stock solutions are diluted to obtain the working solutions required.

Worked example 7.4

Calculate the volume of 1M sodium hydroxide solution needed to make 100 ml of 0.5 M solution of sodium hydroxide.

Solution

Use the formula: $M_1V_1 = M_fV_f$

where

M_1 = Initial concentration; V_1 = Initial volume

M_f = Final concentration; V_f = Final volume

$$\therefore V_1 = \frac{M_fV_f}{M_1};\ M_1 = 1;\ M_f = 0.5, V_f = 100\ \text{ml} = 0.11$$

$$V_1 = \frac{0.5 \times 0.1}{1} = 0.05\ \text{L or } 50\ \text{ml}$$

Therefore, 50 cm³ of 1 M sodium hydroxide solution is needed to make 100 ml of 0.5 M solution.

Worked example 7.5

Suppose your teacher asks you to prepare 250 cm³ of 1 mol dm^{-3} (1M) sodium hydroxide for a chemistry lesson. How would you prepare this solution?

Solution

Step 1:

Work out the moles needed:

1 M means 1000 cm^3 contain 1 mole

250 cm^3 contain x moles

Cross multiply then solve for x

$$\therefore x = \frac{250 \times 1}{1000} = 0.25 \text{ moles}$$

Step 2:

Write the formula of sodium hydroxide and work out the relative formula mass (R.F.M).

Given Na = 23, O = 16; H = 1

Na = 1 × 23 = 23

O = 1 × 16 = 16

H = 1 × 1 = 1

∴R.F.M is 23 + 16 + 1= 40

This means that

1 moles of NaOH $\longrightarrow$ 40 g

0.25 moles of NaOH $\longrightarrow$ y g

Cross multiply and solve for y

$$y = \frac{40 \text{ g} \times 0.25 \text{ moles}}{1} = 10$$

∴ y = 10 g

This means that you weigh 10 g of sodium hydroxide pellets and put in a 250 cm^3 volumetric flask, pour in distilled water and shake to dissolve. Make it up to 250 cm^3 mark with more distilled water.

Concentration and molarity

Determine the molarity of a solution made by dissolving 20.0 g of NaOH in sufficient water to yield a 482 cm^3 solution.

Solution

Molarity is an expression of the moles of solute (NaOH) per litre of solution.

To work this problem, you need to be able to calculate the number of moles of sodium hydroxide (NaOH) and be able to convert cubic centimetres into litres.

Step 1:

Calculate the number of moles of NaOH that are in 20.0 grams.

Look up the atomic masses for the elements in NaOH from the Periodic Table. The atomic masses are found to be:

Na = 23.0

H = 1.0

O = 16.0

1 mol NaOH weighs = 23.0 g + 16.0 g + 1.0 g = 40.0 g

So the number of moles in 20.0 g is:

$$\text{moles NaOH} = \frac{20.0\ \text{g} \times 1\ \text{mol}}{40.0\ \text{g}} = 0.500\ \text{mol}$$

Step 2:

Determine the volume of solution in litres

1 litre = 1000 cm^3, Volume of solution in litres = $\frac{482\ \text{cm}^3 \times 1\ \text{litre}}{1000\ \text{cm}^3}$

= 0.482 litre

Step 3:

Determine the molarity of the solution.

Simply divide the number of moles by the volume of solution to get the molarity:

$$\text{Molarity} = \frac{0.500\ \text{mol}}{0.482\ \text{litre}}$$

Molarity = 1.04 mol/litre = 1.04 M

Self-evaluation Test 7.2

1. How many grams of beryllium chloride are needed to make 125 cm^3 of a 0.050 M solution?
2. The density of ethanol is 0.789 g/cm^3. How many grams of ethanol should be mixed with 225 cm^3 of water to make a 4.5% (v/v) mixture?
3. Explain how to make at least one litre of a 1.25 molar ammonium hydroxide solution.
4. What is the molarity of a solution in which 0.45 grams of sodium nitrate is dissolved in 265 cm^3 of solution?
5. What will the volume of a 0.50 M solution be if it contains 25 grams of calcium hydroxide?
6. How many grams of ammonia are present in 5.0 L of a 0.050 M solution?

7.3 Acid-base titrations

Research activity

1. Using the internet and textbooks from the library, research about the following:
    - Acid-base indicators
    - Apparatus used during titration
    - Handling titration apparatus
    - Conducting simple titration
    - Calculations involving acid-base titration.
2. Prepare a report and present it in class.
3. Compare your finding with the contents outlined below.

We have learnt that when an acid react with a base or alkali; neutralisation reaction takes place. Neutralisation reactions can be done more accurately by a process known as titration. Titration is the process of adding a solution from a burette into a standard solution usually in a conical flask until the reaction between the two solutions is complete. An indicator is used to show the point at which the reaction is complete. The point at which the reaction is complete is called the **end-point**. The volume of the solution added from the burette is known as the **titre**. The volume transferred by the pipette into the conical flask is known as an **aliquot**.

Apparatus used when titrating

When we titrate, we use pipettes and burettes. This is because we deal with accurate volumes and also involves finding exact amounts. This type of experiment is also known as **volumetric analysis**. A pipette delivers fixed and exact volumes of liquid or solution, while a burette is used to add accurate amount of liquid or solution for complete neutralisation. Table 7.1 shows the apparatus commonly used in volumetric analysis and their uses.

Table 7.1: Apparatus commonly used in volumetric analysis and their uses

Apparatus	**Use**
Balance	To weigh solids to be dissolved.
Measuring cylinder	Measure volume of solution to be diluted.
Volumetric flask	To make solution of known concentration.
Pipette	To deliver fixed volumes of liquid or solution into conical flasks.
Burette	To release small amounts of solutions into the conical flasks.
Beaker	To collect and pour solutions into the burette.
Funnel	To facilitate pouring solutions into the burette.
White tile or paper	To help in the observation of colour change in the flask more clearly.
Dropper	To add 2-3 drops of an indicator into the conical flask.

Handling titration apparatus

Burettes and pipettes are expensive apparatus and must be handled with a lot of care to avoid breakage.

Filling a pipette

- Wash the pipette thoroughly with water and then rinse it with the solution given i.e. the one you are supposed to fill the pipette with.
- Suck the solution up the pipette well above the calibration mark. It is recommended to suck the solution from a tilted conical flask to avoid air bubbles entering into the pipette as shown in Fig. 7.3 and to use a pipette filler to fill the pipette.

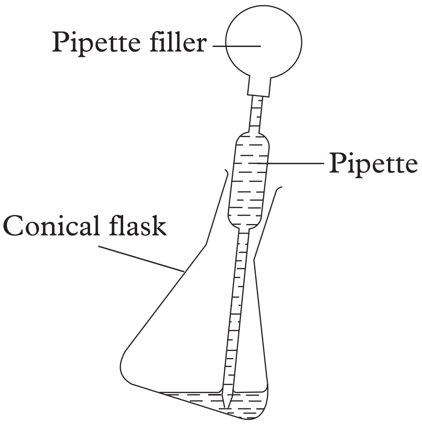


Fig. 7.3: Correct method of filling a pipette with a filler

- Note that when the solution falls below the mark, you have to re-fill all over again. This wastes time. To avoid this hold the pipette upright and read below the meniscus. See Fig. 7.4.
- Allow the solution to run out into the conical flask. Do not force out the little amount of solution left in the pipette as the pipette is graduated with this allowance.

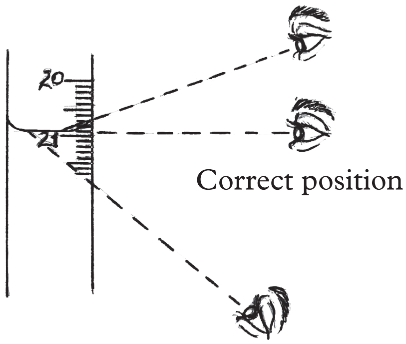


Fig. 7.4: Correct way of reading below the meniscus

- Add 2 - 3 drops of the indicator as shown in Fig. 7.5.

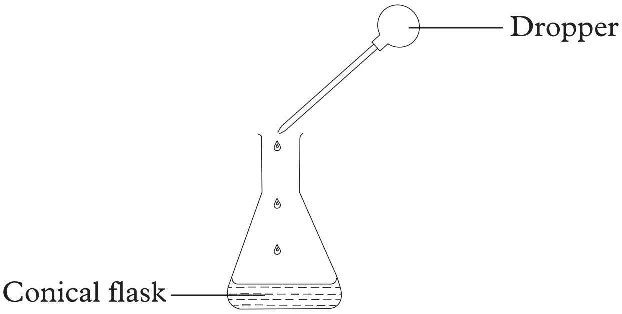


Fig. 7.5: Adding an indicator to a solution

Filling a burrette

A burette is calibrated from top downwards that is 0 cm^3 at the top and 50 cm^3 at the bottom near the tap. Examine the burette and note that the divisions are 1 cm^3.

- Wash the burette with water, then rinse it with the given acid solution or the solution to be added to the burette.
- Clamp the burette to the stand in an upright position as shown alongside.
- Use a filter funnel and a beaker to fill the burette with the acid or solution to above the 0 cm^3 graduation mark.

Note: When filling the burette, hold the funnel with one hand so that air can escape as you fill the burette otherwise air will hold the acid in the funnel and when you lift the funnel, the acid quickly fills the burette and spills over.

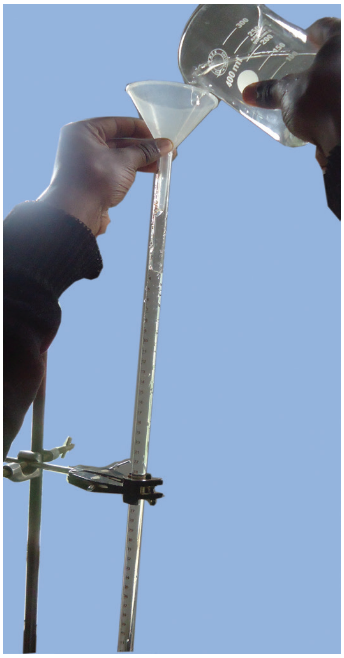


Fig 7.6: Correct method of filling a burette

Titration procedure

Do you remember the procedure for preparation of a salt by reacting an acid with a base (alkali) learnt in Senior 2? The method used was titration. Titration can also be used to find the concentration of the alkali used.

In the school laboratory, titration of sodium hydroxide with hydrochloric acid is the most common. It is carried out in the following steps:

1. 25.0 cm^3 of sodium hydroxide is pipetted into a conical flask to which a few drops of phenolphthalein indicator are added (phenolphthalein indicator is pink in alkaline and colourless in acid solution.

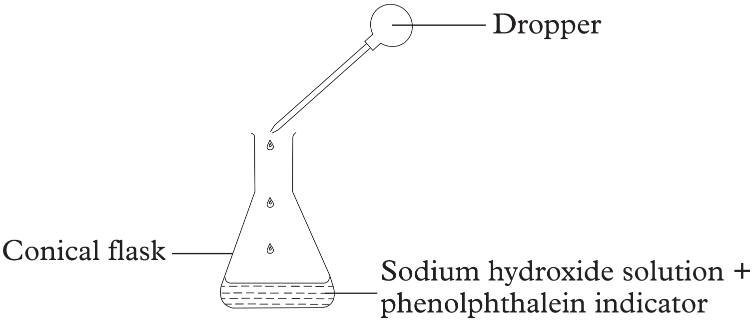


Fig 7.7: Dropping phenolphthalein indicator in a conical flask

2. A solution of hydrochloric acid is then filled in the burette using a filter funnel until it is up exactly to zero mark.

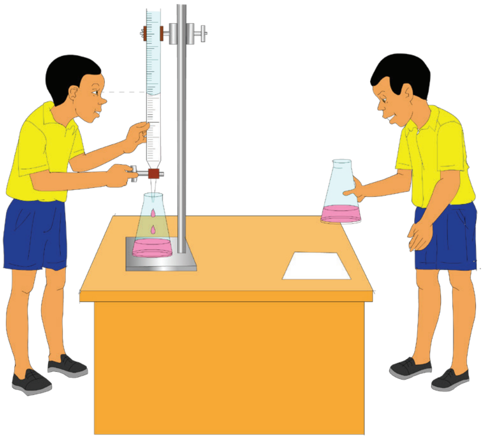


Fig. 7.8: Titration process

3. The filter funnel is then removed.
4. Place a white tile or paper under the conical flask to clearly see the colour changes.
5. The hydrochloric acid is added to the sodium hydroxide solution dropwise usually no more than 0.5cm^3 at a time.

   Note: The contents of the flask is swirled after each addition of acid for thorough mixing while at the same time keeping an eye on the solution in the flask to notice any colour change.
6. The acid is added until the alkali has been completely neutralised. This is shown by the disappearance of the pink colour of the indicator.

   Note: When you swirl the flask and the new colour developing persists for a while, the end point is near. Add one drop of the acid at a time until you get a permanent colour. This is the **end-point** of the titration.

- Read and record the volume of the acid to the nearest 0.1 cm^3. Record your results in a table like the one shown below.

Table 7.2: An example of recording of volume of titration results

Burette reading	**1**	**2**	**3**
Final burette reading (cm^3)	21.50	43.20	21.10
Initial burette reading (cm^3)	0.00	22.00	0.00
Volume of acid used (cm^3)	21.50	21.20	21.10

- Repeat the titration two more times using a different clean conical flask. Add the same number of drops of the indicator as used in the first titration. However, if you use the same flask, you must wash it thoroughly with distilled water and then rinse with a little of the solution you intend to suck into the pipette.

  Note: The first volume obtained should guide you to get the second volume fairly quickly as follows:

"If your first titre volume was 25.0 cm^3, you may add 21 cm^3 quickly while swirling the flask. Then add the acid dropwise while swirling the conical flask until you get the same permanent colour change as in previous titration."

- Repeat the titration using the second volume as a guide to the volume required. You can add the acid until you are about 1 cm^3 from the end point, then you add the acid drop by drop.
- Get the average of any two values for the volume of the acid added, which differ by not more than ± 0.2 cm^3. In other words, the three titrations should be fairly consistent and accurate. Accuracy refers to how close the measurements (that is volumes) are to the "correct or teacher's value.

From Table 7.2, the average volume of acid used

$$= \frac{21.2 + 21.1}{2} = 21.15 \text{ cm}^3$$

Note: The value 21.5 in the table is omitted because it is off the limit of ± 0.2 from the other readings. For example, if you got values such as 24.2 cm^3, 24.5 cm^3 and 24.6 cm^3, which two values would you use for getting the averages? Answer: 24.5 and 24.6 because the difference between them is not more than 0.2 cm^3.

Sample data and treatment of results

After titration, you should be able to carry out calculations involving concentrations of the various solutions, their molarity and masses of substances used. To do this, you should know:

- the reagents involved in the reactions and their relative formular masses.
- the type of reaction taking place and a balanced equation.
- the volumes of various solutions involved.
- the formula:

$M_aV_a = M_bV_b$ where

M_a - concentration of acid; V_a - volume of acid

M_b - concentration of base; V_b - volume of base

Using this formula, the unknown value can be obtained by making it the subject of the formula for cases involving a monobase and a monoacid. For example; considering a neutralisation reaction between 21cm^3 of 0.1M hydrochloric acid and 25 cm^3 of unknown concentration of sodium hydroxide.

$M_a = 0.1\text{mol/dm}^3$ or 0.1 M

$V_a = 21\text{cm}^3$

M_b = unknown

$V_b = 25.0\text{cm}^3$

To calculate M_b; make it the subject of formulae thus $M_b = \frac{M_a V_a}{V_b}$

Substituting in expression above

$$\frac{0.1 \times 21}{25} = 0.084 \text{ M}$$

Note:1dm^3 = 1 litre = 1000ml = 1000 cm^3

Moles = molarity (mol/dm^3) × volume ($\text{dm}^3 = \text{cm}^3/1000$)

$$\text{Molarity (mol/dm}^3\text{)} = \frac{\text{Moles}}{\text{Volume}}$$

1 mole = formula mass in grams.

In most volumetric calculations of this type, you first calculate the known moles of one reactant from a volume and its molarity.

Then, from the equation, you relate this to the number of moles of the other reactant, and then with the volume of the unknown concentration, you work out its molarity as well.

Worked example 7.6

Given the equation: $NaOH (aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

25.0 cm^3 of a sodium hydroxide solution was pipetted into a conical flask and titrated with a standard solution of 0.200 mol dm^{-3} (0.2M) hydrochloric acid.

Using phenolphthalein indicator for the titration, it was found that 15.0 cm^3 of the acid was required to neutralise the alkali.

Calculate the molarity of the sodium hydroxide and its concentration in g/dm^3.

Solution

Moles that reacted

1000 cm^3 of HCl $\longrightarrow$ 0.2 mole

15 cm^3 of HCl used $\longrightarrow$ x mole

$$x = \frac{15 \times 0.2}{1000} = 0.003 \text{ mole}$$

∴ 0.003 mole of HCl reacted.

From the equation, the reacting mole ratio of HCl: NaOH = 1:1,

Hence 0.003 mole of NaOH reacted.

Therefore:

25 cm^3 of NaOH = 0.003 mole

1000 cm^3 of NaOH = y mole

$$\therefore y = \frac{1000 \times 0.003}{25}$$

$$= 0.12 \text{ mole}$$

Hence molarity of NaOH is 0.12 M.

In g/dm^3
1 mole of NaOH = 40 g
0.12 mole of NaOH = 2 g

$$z = \frac{0.12 \times 40}{1}$$

= 4.80 g
Hence concentration of NaOH is = 4.8 g/dm^3

Worked example 7.7

Given the equation: $2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4 + 2H_2O(l)$
20.0 cm^3 of a sulphuric acid solution was titrated with a standardised solution of 0.0500 mol/dm^3 (0.05M) potassium hydroxide.Using phenolphthalein indicator for the titration, the acid required 36.0 cm^3 of the alkali (KOH) for complete neutralisation. What was the concentration of the acid in mol/dm^3 and in g/ dm^3.

Solution

Moles of KOH reacting:

1000 cm^3 of KOH $\longrightarrow$ 0.05 mole

36 cm^3 of KOH $\longrightarrow$ x mole

$$x = \frac{36 \times 0.05}{1000} = 0.0018 \text{ mole}$$

Therefore, the moles of potassium hydroxide reacting is 0.0018 mole.
From the equation of the reaction, the reacting mole ratios is KOH: H_2SO_4 = 2:1
Therefore, the number of moles of H_2SO_4 reacting is $\frac{0.0018}{2}$ = 0.0009 mole
Therefore 20 cm^3 of $H_2SO_4 \longrightarrow$ 0.0009 mole
1000 cm^3 of $H_2SO_4 \longrightarrow$ y mole

$$y = \frac{1000 \times 0.0009}{20} = 0.045 \text{ mole}$$

Molarity of H_2SO_4 = 0.0450 mol dm^{-3} (0.045M)
Concentration in g/dm^3
Since, mass = moles × formula mass
RFM of $H_2SO_4 = 2 + 32 + (4 \times 16) = 98$
1 mole $\longrightarrow$ 98 g

0.045 mole $\longrightarrow$ 2 g

$$2 = \frac{0.045 \times 98}{1}$$

= 4.41g
Therefore, the concentration in g/dm^3 of H_2SO_4 = 4.41 g/dm^3.

Worked example 7.8

Given the equation: $NaOH\ (aq) + HCl\ (aq) \longrightarrow NaCl(aq) + H_2O(l)$

Twenty five (25.00 cm^3) portions of a dilute hydrochloric acid solution were titrated with a standard solution of sodium hydroxide of concentration 0.250 mol/dm^3 (0.25M).

Using phenolphthalein indicator for titration, it was found that the average titre was 18.50 cm^3 of sodium hydroxide. Calculate:

(a) the molarity of the hydrochloric acid

(b) its concentration in g/dm^3.

Solution

Moles of NaOH = molarity NaOH × volume of NaOH (in $dm^3 = cm^3/1000$)

$$\text{Moles of NaOH} = \frac{0.250 \times 18.5}{1000}$$

$$= 0.004625$$

In the equation, 1 mole of NaOH reacts with 1 mole of HCl

Therefore 0.004625 moles of HCl reacted (in 25.00 cm^3).

$$\text{Molarity} = \frac{\text{moles}}{\text{volume in dm}^3} \quad (1\ dm^3 = 1000\ cm^3)$$

$$\text{(a) Molarity HCl} = \frac{0.004625 \times 1000}{25}$$

$$= 0.185\ mol/dm^3$$

(b) Concentration (in g/dm^3) = molarity × formula mass

Formula mass of HCl = 1 + 35.5 = 36.5

$$= 0.185 \times 36.5$$

$$= 6.75\ g/dm^3$$

Worked example 7.9

Given the equation: $2NaOH\ (aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4 + 2H_2O(l)$

25.00 cm^3 portions of a sodium hydroxide solution were titrated with a standardised solution of 0.75 mol/dm^3 sulphuric acid solution using phenolphthalein indicator. If the average titre was 17.70 cm^3 of sulphuric acid, what is the molar concentration of the sodium hydroxide?

Solution

Moles of H_2SO_4 used = Molarity of H_2SO_4 × Volume in dm^3

$$= \frac{0.75 \times 17.70}{1000}$$

$$= 0.013275\ \text{mole}$$

From the balanced equation, for every mole of H_2SO_4, two moles of NaOH react

Therefore, moles of NaOH = 2 × moles of H_2SO_4

Moles of NaOH = 0.013275 × 2 = 0.02655 mol

$$\text{Molarity of NaOH} = \frac{\text{Moles NaOH}}{\text{volume in dm}^3} = \left(\frac{25}{1000} = 0.025\right)$$

$$\text{Molarity of NaOH} = \frac{0.02655}{0.025}$$

$$= 1.062 \text{ mol/dm}^3$$

Worked example 7.10

Given the equation: $NH_3(aq) + HCl(aq) \longrightarrow NH_4Cl(aq)$

5.00 cm³ portions of household ammonia were titrated with a standard hydrochloric acid solution of 1.00 mol/dm³ (1M). If the average titration using methyl orange indicator was 22.5 cm³ of hydrochloric acid, calculate:

(a) the molarity of the ammonia solution, and its concentration in

(b) g/dm³ and

(c) g/cm³

Solution

(a) Moles of HCl in titrated = molarity of HCl × volume of HCl in dm³

$$\frac{1 \times 22.5}{1000} = 0.0225$$

Mole of HCl = 0.0225 moles

From the equation, 1 mole of NH_3 reacts with 1 mole of HCl

Therefore moles of HCl = moles NH_3

= 0.0225 mol

$$\text{Molarity of } NH_3 = \frac{\text{mol } NH_3}{\text{Volume of } NH_3 \text{ in dm}^3} \Rightarrow \left(\frac{5}{1000} = 0.005\right)$$

$$\text{Molarity of } NH_3 = \frac{0.0225}{0.005}$$

$$= 4.50 \text{ M}$$

(b) Concentration in g/dm³= formula mass × molarity

Formula mass NH_3 = 14 + (3 × 1)

= 17g

Concentration of NH_3 = 17 × 4.50

= 76.5 g/dm³

(c) Since there are 1000 cm³ in 1 dm³

$$\text{Concentration of } NH_3 = \frac{76.5}{1000}$$

$$= 0.0765 \text{ g/cm}^3$$

7.4 Titration of an acid with a carbonate (Back titration)

Activity 7.3

You are provided with

- Solution M - which is 0.5 M hydrochloric acid.
- Solution N - containing 5.3 g of a metal carbonate, ($X_2 CO_3$) dissolved in 250 cm^3 distilled water in a volumetric flask.
- Methyl orange indicator.
- Titration apparatus.

You are required to determine the:

(a) Concentration (molarity) of solution N in
  (i) moles per dm^3
  (ii) g dm^{-3}

(b) Relative atomic mass (R.A.M) of X.

(Na = 23; C = 12; O = 16)

You will solve these questions by conducting a titration and filling your results in a table shown below.

Burette reading	1	2	3
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of acid used (cm^3)			

Self-evaluation Test 7.3

1. 24.2 cm^3 of a solution containing 0.2mol/dm^3 of hydrochloric acid neutralised 25 cm^3 of potassium hydroxide solution. What is the concentration of potassium hydroxide solution?
2. 22.4 cm^3 of a solution containing 0.1mol/dm^3 of sulphuric acid neutralised 25 cm^3 of sodium hydroxide. What is the concentration of sodium hydroxide?
3. In a titration involving 24 cm^3 KOH solution against a solution containing 1mol/dm^3 of H_2SO_4, 28 cm^3 of the acid was found to neutralise the alkali completely.
   (a) Write a word and chemical balanced equation for the reaction.
   (b) Name a suitable indicator for the titration and state the colour change you would observe at the end point.
   (c) Calculate the concentration of the alkali in mol/dm^3.
   (d) Describe a chemical test which you would use to identify the type of salt produced during the reaction.

4. 5.95g of potassium bromide was dissolved in 400cm^3 of water. Calculate its molarity.
5. 100 g of sodium chloride (NaCl) is dissolved in 450 cm^3 of water.
   (a) How many moles of NaCl are present in the solution?
   (b) What is the volume of water (in dm^3)?
   (c) Calculate the concentration of the solution.
   (d) What mass of sodium chloride would need to be added for the concentration to become 5.7 mol dm^{-3}?
6. What is the molarity of the solution formed by dissolving 80 g of sodium hydroxide (NaOH) in 500 cm^3 of water?
7. What mass (g) of hydrogen chloride (HCl) is needed to make up 1000 cm^3 of a solution of concentration 1 mol.dm^{-3}?
8. How many moles of H_2SO_4 are there in 250 cm^3 of a 0.8M sulphuric acid solution? What mass of acid is in this solution?
9. 25 cm^3 of sodium hydroxide solution was pipetted into a conical flask and titrated with 0.2M hydrochloric acid. Using a suitable indicator, it was found that 15 cm^3 of acid was needed to neutralise the alkali. Calculate the molarity of the sodium hydroxide.
10. 4.9 g of sulphuric acid is dissolved in water and the final solution has a volume of 220 cm^3. Using titration, it was found that 20 cm^3 of this solution was able to completely neutralise 10 cm^3 of a sodium hydroxide solution. Calculate the concentration of the sodium hydroxide in mol.dm^{-3}.

Unit summary 7

- A standard solution contains a precisely known concentration of an element or a substance.
- End point is the point in a titration at which a reaction is complete, often marked by a colour change.
- Concentration is the amount of solute per unit volume of solution.
- Molarity (M) is the concentration of a solution expressed as the number of moles of solute per litre of solution.
- Titration is a technique where a solution of known concentration is used to determine an unknown concentration of a given solution.
- An acid–base titration enables determination of the concentration of an acid or base by exactly neutralising it with an acid or base of known concentration. This allows for quantitative analysis of the concentration of an unknown acid or base solution.

Test your Competence 7

1. Which one of the following pieces of apparatus is not used during titration?
   - A. Liebig condenser
   - B. Burette
   - C. Pipette filler
   - D. Conical flask
2. Which one of the following equations is correct for the reaction between hydrochloric acid and sodium carbonate?
   - A. $Na_2CO_3(s) + HCl\ (s) \longrightarrow NaCl\ (aq) + H_2O\ (l) + CO_2\ (g)$
   - B. $Na_2CO_3\ (s) + 2HCl(aq) \longrightarrow 2NaCl\ (aq) + H_2O\ (l) + CO_2\ (g)$
   - C. $Na_2CO_3\ (s) + 2HCl\ (aq) \longrightarrow NaCl(aq) + H_2O\ (l)$
   - D. $Na_2CO_3\ (s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + CO_2\ (g)$
3. Before starting titration, the pipette, burette and conical flask are rinsed with deionised water and not tap water. Why is this?
   - A. Ions in the tap water may 'clog up' the tap of the burette.
   - B. The acid, when placed in the burette will become more dilute if tap water rather than deionised water is used to rinse the burette.
   - C. Tap water contains ions that may affect the titration result.
   - D. The colour change at the end-point of the titration is more noticeable if deionised water is used.
4. When placing the acid in the burette, which one of the following procedures is incorrect?
   - A. The top of the meniscus is read when doing the titration.
   - B. Any air bubble below the tap is removed before the titration.
   - C. The funnel is removed from the top of the burette before the titration.
   - D. The burette is filled to the zero mark when filling it up.
5. When pipetting sodium carbonate solution into the conical flask, which one of the following procedures is always carried out?
   - A. The sodium carbonate solution must be pipetted directly from the volumetric flask into the conical flask.
   - B. When emptying the contents of the pipette into the conical flask, the last few drops are blown out.
   - C. A pipette filler is used instead of using one's mouth to suck up the solution.
   - D. The pipette is always rinsed with sodium carbonate solution first and then with deionised water.
6. Why is the sodium carbonate mixed with the acid in a conical flask rather than a beaker?
   - A. Its easier to see the colour change in a conical flask.
   - B. The conical flask will be less likely to break during the titration.

C. To prevent any of the solutions splashing out when swirling the flask.
D. The conical flask is more accurate for measuring volume.

7. When adding the indicator to the conical flask, only two or three drops are added because
   - (i) too much indicator may make it difficult to see the colour change at the end-point clearly,
   - (ii) the indicator is corrosive, and adding too much poses a greater risk,
   - (iii) the indicator may be a weak acid or base and may interfere with the titration result.

   Which of the above statements is/are correct?
   A. (i) and (ii) only
   B. (ii) and (iii) only
   C. (iii) only
   D. All of them
8. When adding the acid from the burette into the conical flask
   - (i) the acid is initially added drop by drop,
   - (ii) the conical flask is swirled while adding the acid,
   - (iii) the conical flask is placed over a white tile or white sheet of paper,
   - (iv) the inside of the conical flask is washed down from time to time with deionised water.

   Which of the above procedures is/are not carried out during the titration?
   A. (i) and (iii) only
   B. (i) only
   C. (iv) only
   D. (ii) and (iii) only
9. If the colour of the indicator at the end-point changes from yellow to peach/pink, then which one of the following indicators was most likely to have been used?
   A. Litmus indicator
   B. Universal indicator
   C. Phenolphthalein
   D. Methyl orange
10. How many titrations should be carried out in order to give a valid result?
   A. One rough titration and two accurate titrations that agree to within 0.1 cm^3 of each other.
   B. Three titrations which are then averaged to get a result.
   C. One rough titration and one accurate titration.
   D. Two rough titrations and one accurate titration that agrees to within 0.5 cm^3 of the rough titrations.
11. The equation for the reaction taking place in this titration is

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

If **A** represents the acid, and **B** represents the base, then what are the values of N_A and N_B in calculating the concentration of the hydrochloric acid solution?

A. $N_A = 1$ and $N_B = 1$
B. $N_A = 1$ and $N_B = 2$
C. $N_A = 2$ and $N_B = 2$
D. $N_A = 2$ and $N_B = 1$

12. Using the correct values of N_A and N_B from the previous question, and given that 25 cm^3 of 0.1 M sodium carbonate solution was used in neutralising 21 cm^3 of hydrochloric acid solution, what is the concentration of the acid in moles per litre?

A. 0.024 mol/litre
B. 0.06 mol/litre
C. 0.12 mol/litre
D. 0.24 mol/litre

13. Given that the concentration of a hydrochloric acid solution is 0.24 mol/litre, what is this concentration in grams per litre? The molar mass of hydrochloric acid is 36.5 grams.

A. 8.52 grams/litre
B. 0.0066 grams/litre
C. 8.76 grams/litre
D. 36.74 grams/litre

14. Fill in the question below (use the following words where appropriate: indicator, volume, unknown, conical, titration, pipette acid, orange, burette, neutralised, end-point)

The technique used to find the concentration of an ________ solution is called a ________. In an acid-base titration, a certain ________ of the base is measured out using a ________ of a certain value, for example 20 cm^3, into a ________ flask and an ________ is added. The ________ is slowly and carefully added to the measured volume of the base using a ________ until the indicator, for example, methyl ________, just changes colour. This is called the ________. The indicator is used to show when the base has been completely ________ by the acid.

UNIT 8

Electrolysis and its applications

Key unit competency

To be able to examine and explain the electrolysis of different electrolytes and state their application in daily life.

Learning objectives

By the end of this unit, I should be able to:

- Explain the concept of electrolysis.
- Predict the products of electrolysis of: molten lead bromide, concentrated sodium chloride (brine), molten sodium chloride and electrolysis of water (dilute sulphuric acid) using platinum and copper electrodes.
- Carry out an experiment to investigate the electrolysis of copper (II) sulphate solution.
- Write balanced equations to represent reactions that take place at each electrode during electrolysis.
- Write balanced redox equations in terms of loss of oxygen or gain of hydrogen and viceversa.
- Perform an experiment to electroplate a silver coin using copper (II) sulphate or nitrate solution.
- Develop a culture of working in a team and respect the procedures of an experiment.
- Develop orderliness especially during experiments and presentations.
- Develop self-confidence in handling electrical apparatus and chemicals.
- Appreciate the application of electrolysis in daily life.

Mind teaser

A Senior 3 student set up the apparatus as shown alongside. To his surprise, the balloons started expanding after sometime. Balloon Y at the end of the experiment was twice the size of balloon X.

What could be the possible explanation of the observations in this experiment? Can you make a conclusion regarding the behaviour of ionic solutions such as sodium chloride (salt) when an electric current is passed through them?

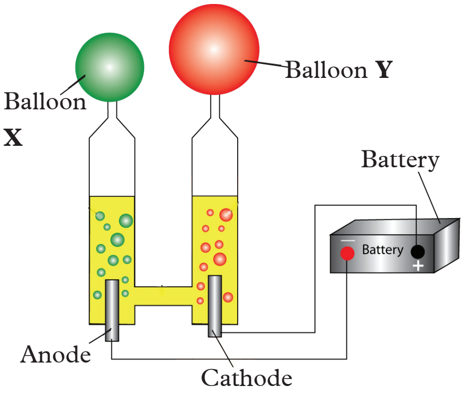


Fig 8.1

Introduction

Electrolysis is the decomposition of an electrolyte by an electric current that is passed through it. The products of electrolysis vary depending on certain factors. Among such factors include the nature of the electrodes used and the concentration of the electrolyte. Electrolysis has found application in various fields, such as electroplating and in purification and extraction of metals.

8.1 Introduction to oxidation-reduction reactions

Activity 8.1

Materials and apparatus

- 100 cm^3 beakers (2)
- 1 small spoonful of copper chloride crystals
- Stirring rod
- 8 cm × 8 cm piece of aluminium foil
- Filter paper
- Ring stand with iron ring
- Funnel

Procedure

1. Put a spoonful of copper(II) chloride in one beaker.
2. Add enough water to make a solution to about 1000 cm^3 in volume.
3. Stir the mixture until all the solid dissolves. Record your observations.
4. Cut the aluminium foil into tiny pieces.
5. Collect the aluminium pieces and place them all in the beaker with the solution.
    - Record your observations.
6. When the reaction appears to be nearly complete, stir the mixture to ensure a complete reaction.
7. Set up a filter apparatus and filter the mixture. Collect the filtrate in the other beaker.
8. Make observations of the residue and the filtrate.

Study questions

1. What happened to the aluminium foil and what is the reddish-brown substance?
2. Why did it get so hot?
3. Why did the green colour disappear?

Discussion corner!

1. Discuss in groups the observations made in the experiment and the study questions above.
2. Present your finding to the class.

I have discovered that…

Aluminium foil reacts vigorously and exothermically with the copper chloride solution. The aluminium foil first turns black and then reddish-brown solid forms as the reaction proceeds(copper). The solution soon loses its green colour (Cu^{2+} gone; Al^{3+} formed).

The facts

Aluminium metal replaces Cu^{2+} from the solution as shown in the equation.

$$3CuCl_2(aq) + 2Al(s) \longrightarrow 2AlCl_3(aq) + 3Cu(s)$$

a) Oxidation and reduction in terms of electron transfer

Oxidation is loss of electrons while reduction is gain of electrons.

Remembering these definitions is essential. It can be easily done using this convenient acronym:

O I L
Oxidation is loss

R I G
reduction is gain

b) Oxidation Number

Reasearch activity

1. Using the internet and textbooks from the library, research on:
   - Changes of oxidation numbers during redox reactions.
   - Determination of oxidation numbers.
   - How to calculate oxidation states.
2. Write a report and present it in class.

When dealing with ions **oxidation number** can be viewed as being equal to the net charge on the ion. Oxidation is then viewed as a reaction which increases the oxidation number and reduction as one which reduces the oxidation number. This view of oxidation and reduction helps us deal with the fact that "oxidation" can occur even when there is no oxygen! The definition of redox reactions is extended to include other reactions with non-metals such as chlorine and bromine, for example, in the reaction below.

$$Mg(s) + Cl_2(g) \longrightarrow Mg^{2+}(aq) + 2Cl^-(aq)$$

The Mg is seen to increase in oxidation number from **0 to 2** (oxidation) while the chlorine atoms experience a decrease in oxidation number from **0 to -1** (reduction). An atom is considered to have oxidation number zero(0).

Determination of oxidation numbers

To determine the oxidation number of elements in a molecule, compound or a complex ion, the following rules are applied:

- The oxidation number of an atom of any free (uncombined) element is **zero,** for example K, Na, Zn etc have zero as their oxidation numbers.
- The oxidation number of any simple one ion is equal to its **charge**. Thus the oxidation number of Na^+, Ca^{2+}, Cl^-, etc are Na = +1, Ca = +2 and Cl^- = -1 respectively.
- The oxidation number of hydrogen in non- ionic compounds is **+1**. This applies to hydrogen compounds such as HCl, CH_4, H_2O, among others except in **metal hydrides** such as NaH where its oxidation number is **-1**.
- The oxidation number of oxygen is **-2** in all its compounds such as H_2O, NO, SO_2 etc except in peroxides such as H_2O_2 where it is **-1**.
- The algebraic sum of the oxidation numbers of all atoms in the formula of a neutral compound must be **zero**.
- The algebraic sum of the oxidation numbers of all the atoms in a complex ion is equal to the **charge** on the ion.

Examples

1. What is the oxidation number of **N** in NH_4^+?

Solution

Let the oxidation number of N be represented by letter 'y', hence

$y + (4 \times 1) = +1$

$y + 4 = +1$

Therefore,

$y = +1 - 4 = -3$

Thus, the oxidation number of **N** in NH_4^+ is -3

2. What is the oxidation number of **S** in SO_2 ?

Solution

Let the oxidation number of S be represented by letter 'z', hence

$z + (2 \times -2)=0$

$z +(-4) = 0$

$z = +4$

Thus, the oxidation number of **S** in SO_2 is +4

3. Determine the oxidation number of chlorine in ClO_3^-

Solution

Let the oxidation number of Cl be represented by letter ' n', hence

$n + (3 \times -2)= -1$

$n + (-6) = -1$

$n = -1 + 6 = +5$

Thus, the oxidation number of **Cl** in ClO_3^- is +5.

c) Oxidation and reduction in terms of oxygen transfer

Discussion corner!

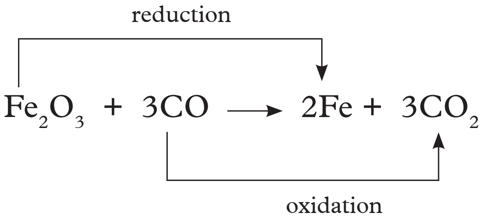
1. Discuss with a friend the following:
   a. What are oxidising and reducing agents?
   b. Define oxidation and reduction in terms of oxygen transfer and hydrogen transfer.
2. Compare your findings with the rest of the class.

The terms oxidation and reduction can be defined in terms of adding or removing oxygen to a compound. While this is the easiest to remember it is the least accepted definition.

Oxidation is the gain of oxygen.

Reduction is the loss of oxygen.

Example is the extraction of iron from its ore:



$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

Because both reduction and oxidation are occurring simultaneously, this is known as a redox reaction.

An oxidising agent is substance which increase the oxidation state of another substance during a reaction. In the above example, the iron (III) oxide is the oxidising agent. **A reducing agent** lowers the oxidation state or removes oxygen of another substance. In the equation above, the carbon monoxide is the reducing agent.

- Oxidizing agents
  (i) give up their oxygen to another substance.
  (ii) accepts electrons or gain electrons.
- Reducing agents
  (i) remove oxygen from another substance.
  (ii) donate or lose their electrons.

d) Oxidation and reduction in terms of hydrogen transfer

Oxidation is the loss of hydrogen while reduction is the gain of hydrogen.
For example, ethanol can be oxidised to ethanal:

$$CH_3CH_2OH \longrightarrow CH_3CHO$$

oxidation by loss of hydrogen

An oxidising agent is required to remove hydrogen from the ethanol. A commonly used oxidising agent is potassium dichromate (VI) solution acidified with dilute sulphuric acid. Ethanal can also be reduced back to ethanol by adding hydrogen. A possible reducing agent is sodium tetrahydridoborate, $NaBH_4$. Again the equation is too complicated to consider at this point.

reduction by gain of hydrogen

$$CH_3CHO \longrightarrow CH_3CH_2OH$$

Oxidizing and reducing agents

A more precise definition of oxidising and reducing agents

- Oxidising agents add oxygen to another substance or remove hydrogen from it.
- Reducing agents remove oxygen from another substance or add hydrogen to it.

8.2 Electrolysis of solutions

Reasearch activity

Research and present on electrolysis of solutions under the following sub headings:

1. Definitions of electrolysis, electrode, anode and cathode.
2. Role of water in electrolysis.
3. Factors affecting preferential discharge.

Arrhenius Theory of electrolysis

Arrhenius put forth, in 1884, a comprehensive theory which is known as theory of electrolytic dissociation or ionic theory.

An electrolyte, when dissolved in water, breaks up into two types of charged particles, one carrying a positive charge and the other a negative charge. These charged particles are called **ions**. Positively charged ions are termed **cations** and negatively charged as **anions**.

Fig 8.2 Arrhenius August

For example

$NaCl(aq) \longrightarrow Na^+(aq) + Cl^-(aq)$

$K_2SO_4(aq) \longrightarrow 2K^+(aq) + SO_4^{2-}(aq)$

Electrolytes ions

In its modern form, the theory assumes that solid electrolytes are composed of ions which are held together by electrostatic forces of attraction. When an electrolyte is dissolved in a solvent, these forces are weakened and the electrolyte undergoes dissociation into separate ions. The ions are solvated. The process of splitting of the molecules into ions of an electrolyte is called **ionisation**.

Electrolysis is the decomposition of a compound by passing an electric current through it. The compound being electrolysed is called **electrolyte**. The electrolyte is a compound which when in solution or molten state, conducts an electric current and is decomposed by it. In preparing electrolytes for electrolysis, we either melt solid compounds or dissolve them in water.

During electrolysis, electrodes are used. **Electrodes** are electric conductors usually metals or carbon used as the two terminals during electrolysis. The electrode can either be positively charged called **anode** or negatively charged called **cathode**. The anode is the electrode through which electrons leave the electrolyte and current enters the electrolyte. The cathode is the electrode through which electrons enter the electrolyte and current leave the electrolyte.

When two or more cations, e.g. Na^+ and H^+ or anions such as OH^- and SO_4^{2-} migrate to the same electrode one is selected and is discharged preferentially. There are factors that influence ions to be discharged. They include:

1. concentration of ions in solution.
2. nature (types) of electrode
3. position of the ions in the electrochemical series.

Electrolysis of acidified water or (dilute sulphuric acid)

Activity 8.2

Material and apparatus

- 6-volt or 9-volt battery
- Two alligator clip leads or insulated wire
- Beaker or glass
- Piece of thin cardboard or cardstock
- Two sharpened pencils

Procedure

1. Fill the beaker or glass with warm water.

2. Carefully remove the erasers and metal sleeves so you can sharpen both ends of each pencil. These pencils are your electrodes. The graphite in them will conduct electricity, but will not dissolve into the water.

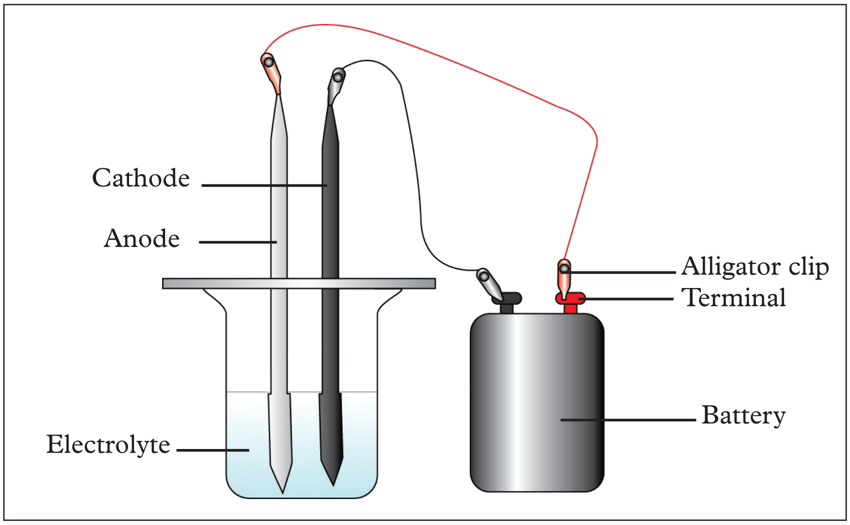


Fig 8.3 Electrolysis of water

3. Cut a piece of the cardboard to fit over the beaker, then punch two holes in the centre of the cardboard about one inch apart. Push the pencils through the holes and set them in the glass. They should extend into the water, but not touch the bottom of the glass. The cardboard will hold them in place.
4. Connect each pencil to the battery with a crocodile clip lead attached to the exposed graphite (pencil lead). If you do not have crocodile clip leads, use two lengths of wire and strip an inch of insulation off each end. Wrap the wire around the graphite of each pencil and connect the wires to the battery. You may need to use tape to hold the wires in place.

Study questions

1. Does one pencil collect more bubbles than the other? Which one? Why do you think this is?
2. Why is electrolysis of dilute sulphuric acid referred electrolysis of acidified water?
3. Write equations taking place at each electrode.
4. Identify which electrode produces more gas bubbles.

Discussion corner!

1. In groups discuss the observations made in the experiment and the study questions above.
2. Write a report of your findings and present it to the class.

I have discovered that…

As soon as you connect the wires to the battery, you will see bubbles appearing around each tip of the pencil inside the water and moving upward. Those bubbles are the hydrogen gas and oxygen gas that have been split apart by the electricity as it travels through the water from one pencil to the other. The pencil attached to the negative terminal of the battery collects hydrogen gas while the one connected to the positive terminal collects oxygen.

The facts

Water is a good solvent. This is because its molecules are polar i.e. its oxygen molecule is partially negative whereas hydrogen atom is partially positive. Because of this, the oxygen atoms are attracted by the positive ions while the hydrogen atoms are attracted by the negative ions during dissolution. The solid in this case is the ionic compound being dissolved. The attraction between the water molecules and the ions in the solid may be so intense that it overcomes the forces of attraction holding the ions in the solid together. When this happens, the solid dissolves. The ions then become free to move in the aqueous solution and therefore, can conduct an electric current during electrolysis. Water is a weak electrolyte. It is only slightly dissociated. Its molecules dissociate into hydrogen ions (H^+) and hydroxide ions (OH^-) i.e.

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

During electroysis, the hydrogen ions (H^+) are attracted to the cathode and hydroxide ions (OH^-) to the anode.

Reaction at cathode (negative electrode)

Hydrogen ions + $2e^-$ ⟶ Hydrogen gas.

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

Hydrogen gas liberated at cathode.

Reaction at anode (positive electrode)

Hydroxide ions ⟶ oxygen gas + water + 4 electrons

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

Oxygen gas liberated at anode

Overall reaction

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

Electrolysis of dilute sodium chloride

Activity 8.3

Apparatus and chemicals

- Beaker
- Carbon (graphite) rods
- Wooden splint
- 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 chloride

Procedure

1. Set up the apparatus as shown in Fig 8.4.

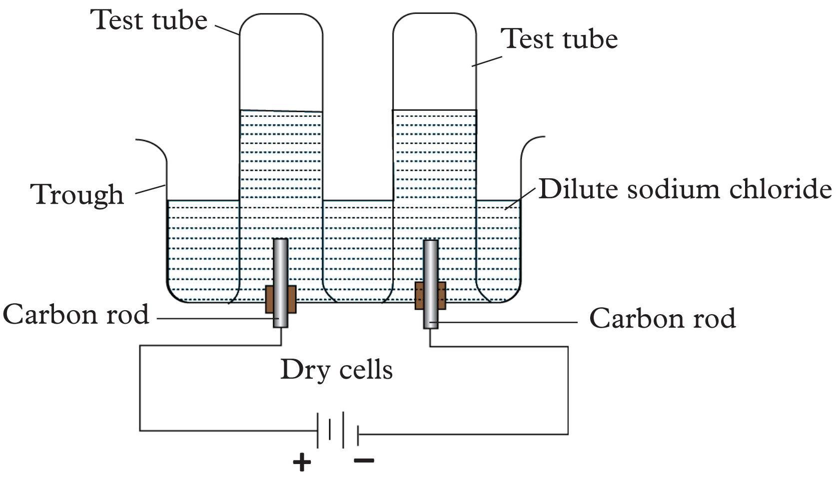


Fig. 8.4: Electrolysis of dilute sodium chloride solution

2. Half-fill the trough with a dilute solution of sodium chloride.
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 a glowing and burning splints.
6. Copy Table 8.1 in your notebook and record your observations.

Table 8.1: Electrolysis of dilute sodium chloride solution

Gas collected	Colour	Test with glowing / burning splint
Anode		
Cathode		

I have discovered that…

A colourless gas, which re-lights a glowing splint, is collected at the anode. The gas is therefore oxygen. At the cathode, a colourless gas is collected which burns with a 'pop' sound. This gas is therefore hydrogen.

The facts

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 chloride	$Na^+(aq)$, Cl^- (aq)
From water	$H^+(aq)$, OH^- (aq)

The positive ions (cations), Na^+ and H^+ migrate to the cathode. The negative ions, (anions) Cl^- and OH^-, migrate to the anode. This is because of the attraction of opposite charges.

At the cathode	**At the anode**
Ions present, $Na^+(aq)$, $H^+(aq)$	Ions present, $Cl^-(aq)$, $OH^-(aq)$
Both ions move to the cathode. H^+ ion is discharged in preference to Na^+ ion because it is lower than Na^+ ion in the electrochemical series. It gains an electron to form H atom. $H^+(aq) + e^- \longrightarrow H(g)$ $H(g) + H(g) \longrightarrow H_2(g)$ $2H^+(aq) + 2e^- \longrightarrow H_2(g)$	Both ions move to the anode. OH^- ion is discharged because it is lower in electrochemical series OH^- is discharged by losing an electron. $OH^-(aq) \longrightarrow OH(g) + e^-$ 4OH^- ions interact to form one mole of oxygen and two moles of water as shown below, releasing 4 electrons which move to the cathode. Overall reaction $4OH^-(aq) \longrightarrow O_2(g) + 2H_2O(l) + 4e^-$

But the $4e^-$ produced at the anode must react hence overall equation is $4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$	

Note: The $4e^-$ 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 Cl^- ions remain in solution, resulting in sodium chloride solution of higher concentration.

Electrolysis of brine (concentrated sodium chloride)

Activity 8.4

Apparatus and chemicals

- The apparatus is the same as those in experiment 8.2. You can also use glass cell fitted with inert electrode of platinum.
- 1 M sodium chloride solution.

Procedure

1. Set up the apparatus as shown in Fig. 8.5.

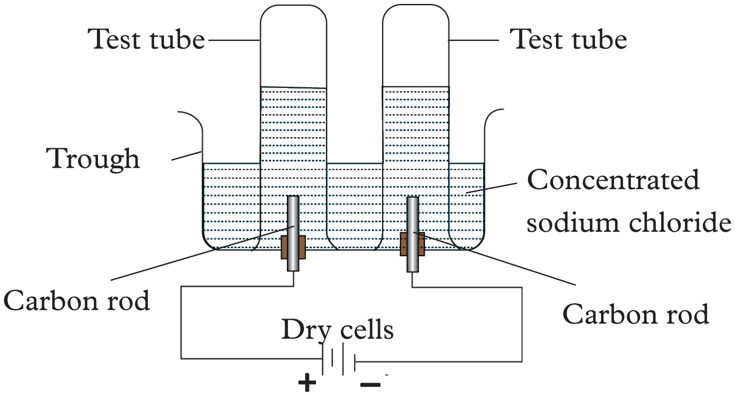


Fig. 8.5: Electrolysis of brine

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 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?

The facts

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 a 'pop' sound is collected. This gas is therefore hydrogen.

Ions present

From sodium chloride	$Na^+(aq), Cl^-(aq)$
From water	$H^+(aq), OH^-$

At the cathode	At the anode
$Na^+(aq), H^+$ H^+ ions are discharged in preference to Na^+ ions because H^+ ion is lower in electrochemical series than Na^+ ion. $2H^+(aq) + 2e^- \longrightarrow H_2(g)$	$Cl^-(aq), OH^-(aq)$ Although OH^- ion is lower in the electrochemical series than Cl^-, the concentration of Cl^- ions at the cathode is greater than the OH^- ions. Therefore, Cl^- ions are discharged in preference to OH^- ions $2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$

Note

(a) The concentration has determined the product at the cathode.

(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

$$Cl_2(g) + 2OH^-(aq) \longrightarrow Cl^-(aq) + OCl^-(aq) + H_2O(l)$$

Role play

1. Two points can be chosen in the classroom. For example, two benches, one with a positive sign and one with a negative sign. This represents the electrodes in the system.
2 Create several A4 signs, two with Na^+, two with Cl^-, eight with H, four with O and one with a picture of a bulb.

3. Choose two pairs of students, with each pair linking arms.
4. Within each pair, one student holds the Na^+ sign while the other student holds the Cl^- sign.
5. These two pairs show an ionically bonded sodium chloride compound.
6. Another student can hold the sign of a bulb, showing that the electricity is flowing by holding it up.
7. The students should then consider what happens when the bond is broken and the Na^+ and the Cl^- ions move to the cathode and anode respectively.

Self-evaluation Test 8.1

1. Complete the paragraph using the words below:
   (*Anode*, *cathode*, *gas*, *hydrogen*, *lose*, *oxygen*, *positive*)
   During electrolysis _________ ions move towards the _________ and negative ions move towards the _________. At the cathode the metal or _________ ions gain electrons and form metal atoms or hydrogen _________. At the anode the non-metal ions _________ electrons and form halogen or _________.
2. Explain the differences in the electrolysis products of:
   (a) Molten sodium chloride
   (b) Dilute aqueous solution of sodium chloride.
3. Write half equations for the following electrolysis reactions:
   (a) Lead ions (Pb^{2+}) reacting to form lead (Pb)
   (b) Hydrogen ions reacting to form hydrogen gas
   (c) Hydroxide ions reacting to form oxygen gas.
   (d) Chloride ions reacting to form chlorine gas.

Electrolysis of copper (II) sulphate solution (using inert electrodes)

Inert electrodes are substances like carbon (graphite) or platinum. They take no part in electrolysis except to carry electrons to and from the electrolyte.

Activity 8.5

Apparatus and chemicals

- Battery/dry cells
- Carbon rods
- Connecting wires fitted with crocodile clips
- Beaker
- Bulb (6V) or ammeter
- Copper (II) sulphate solution

Procedure

1. Half-fill a beaker with copper (II) sulphate solution.
2. Set up the apparatus as shown in Fig. 8.6 or 8.7.

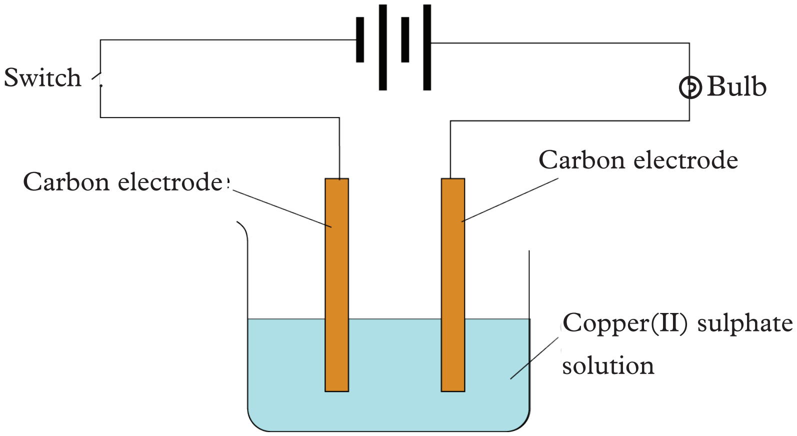


Fig. 8.6: Electrolysis of copper (II) sulphate using graphite electrodes

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?

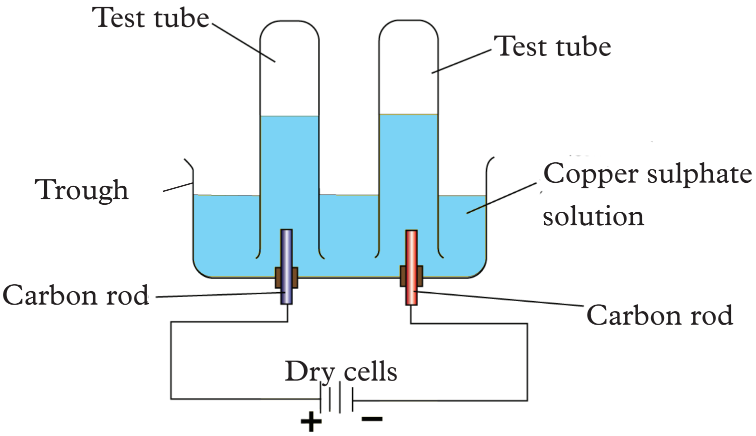


Fig. 8.7: Apparatus showing electrolysis of copper (II) sulphate using graphite electrodes.

Study question

1. What is the purpose of the bulb?
2. What is the purpose of the switch?
3. What did you observe at each of the electrodes?
4. What happens to the colour of the solution at the end of the experiment?

The facts

Ions present

From copper (II) sulphate, $Cu^{2+}(aq)$, $SO_4^{2-}(aq)$
From water $H^+(aq)$, $OH^-(aq)$

At the cathode	At the anode
• $Cu^{2+}(aq)$, $H^+(aq)$ both move to the cathode. • 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. Ionic equation: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$ Account for the $4e^-$ from the anode.	• SO_4^{2-} (aq), OH^- (aq) both move to the anode. • OH^- ions preferentially lose electrons being lower in the electrochemical series. They combine to form water and oxygen Ionic equation: $4OH^-(aq) \longrightarrow 2H_2O(l) + O_2\,(g) + 4e^-$ Bubbles of oxygen gas are observed at the anode.

Copper(II), Cu^{2+}, ions, give copper (II) sulphate solution its blue colour. As they are being discharged and deposited at the cathode as copper metal, the colour of the solution fades and if we continue with electrolysis the solution finally turns colourless and acidic.

- If we continue passing electricity in the colourless solution what would be the product at:
  (i) cathode?
  (ii) anode?

Note

Further electrolysis of the colourless solution would give the same products as electrolysis of dilute sulphuric acid.

Electrolysis of copper (II) sulphate solution (using copper electrodes)

Activity 8.6

Apparatus and chemicals

- As in experiment 8.5 but using copper electrodes
- Balance (sensitive)
- Distilled water
- Propanone

Procedure

1. Put copper (II) sulphate solution in a beaker.

2. Clean the copper electrodes until they shine.
3. Weigh the electrodes and record the mass in your notebook as in Table 8.5 below.
4. Set up the apparatus as shown in Fig. 8.8 and switch on electric current.

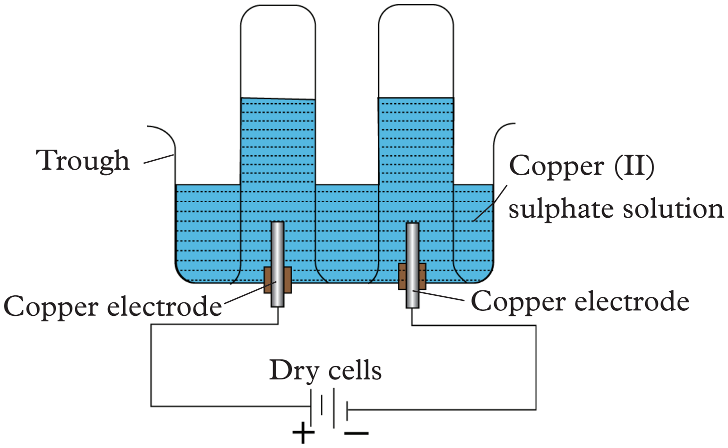


Fig. 8.8: Electrolysis of copper (II) sulphate using copper electrodes

- 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 them dry off.
7. When dry, re-weigh the electrodes.
8. Record the mass obtained after the experiment.
   - Why must electrodes be cleaned before electrolysis?
   - Why must the electrodes be washed with distilled water and propanone after electrolysis?

Table 8.5: Results of electrolysis of copper(II) sulphate solution using copper electrodes

	Before passing electricity		**After passing electricity**	
	Anode	**Cathode**	**Anode**	**Cathode**
Colour of copper (II) sulphate solution				
Mass of electrode				
Change in mass of electrode				

Study questions

1. Observe and describe what happens at each electrode?
2. What happens to the colour of the solution at the end of the experiment?

Discussion corner!

Discuss the observations made during the experiment in groups and the study questions above. Present to the class your findings.

I have discovered that…

The masses of both electrodes change during electrolysis. The anode loses mass while the cathode gains mass.

The facts

The electrodes must be cleaned to remove any oxide that might have formed on the surface. The electrodes are washed with water and propanone to ensure they are completely dry.

Ions present:

From copper (II) sulphate, $Cu^{2+}(aq)$, SO_4^{2-} (aq)

From water $H^+(aq)$, OH^- (aq)

At the cathode	**At the anode**
$Cu^{2+}(aq)$, $H^+(aq)$ These ions move to the cathode. Cu^{2+} ions are lower in 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 as shown in Fig. 8.9 and it increases the mass of cathode.	$SO_4^{2-}(aq)$, $OH^-(aq)$ 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 gets dissolved. This process is called **electrode ionisation** and it causes a decrease in anode mass.

Note: Copper (II), Cu^{2+}, ions make copper (II) sulphate solution blue. For each ion removed at the cathode, it is replaced when the anode copper dissolves at the anode. Therefore, the solution remains blue and the concentration of the solution remains constant.

When we electrolyse copper (II) 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. 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.

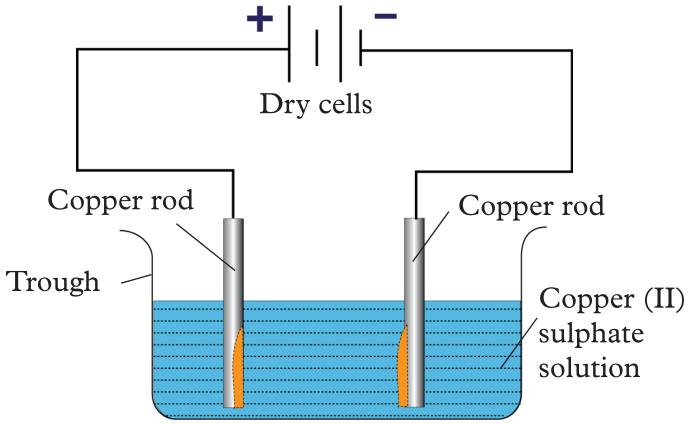


Fig. 8.9 : Loss in mass at the anode is equal to gain in mass at the cathode

Self-evaluation Test 8.2

1. Draw a table to summarise the similarities and differences when copper (II) sulphate solution is electrolysed using:
   (a) Platinum electrodes
   (b) Copper electrodes
2. Complete the following passage using the words below:
   (*Cathode*, *copper*, *electrolysed*, *impure*, *solution*)
   When a solution of copper (II) sulphate is ________ using ________ electrodes the copper atoms at the ________ anode go into ________ as copper ions. At the ________ the copper ions turn into copper atoms.

8.3 Electrolysis of molten compounds

Discussion corner!

Carry out a research in your study group and present to the class about the electrolysis of:

- Molten lead (II) bromide
- Molten aluminium oxide

Electrolysis of molten lead (II) bromide

The facts

Electrolyte: Molten lead (II) bromide ($PbBr_2$)

Electrode: carbon (inert)
Ions present: Pb^{2+} + $2Br^-$

Reaction at cathode (negative electrode)

Lead ions + 2 electrons ⟶ Lead atoms

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

Molten lead will collect at the bottom near cathode.

Reaction at anode (positive electrode)

Bromide ions ⟶ bromine gas + 2 electrons

$2Br^- \longrightarrow Br_2 + 2e^-$

Bromine gas will be discharged at anode as a red brown gas.

Overall reaction:

Lead bromide ⟶ molten lead + Bromine gas

$PbBr_2\ (l) \longrightarrow Pb\ (l) + Br_2\ (g)$

8.4 Applications of electrolysis

Research activity

1. Using the internet and texbooks from the library, research on the applications of electrolysis. Use the following subheadings:
   (i) Extraction of metal
   (ii) Refining metals
   (iii) Electroplating metal
2. Write a report and present it in class.
3. Compare your findings with the contents below.

Electrolysis has many applications.

1. Extraction of metals

Extraction of metals by the process of electrolysis is known as electro-metallurgy. This process is used to extract highly reactive metals such as sodium, magnesium and aluminium from their molten ionic compounds, e.g.

- Sodium is extracted from molten sodium chloride.
- Aluminium is extracted from molten bauxite (hydrated aluminium oxide, $Al_2O_3.2H_2O$).

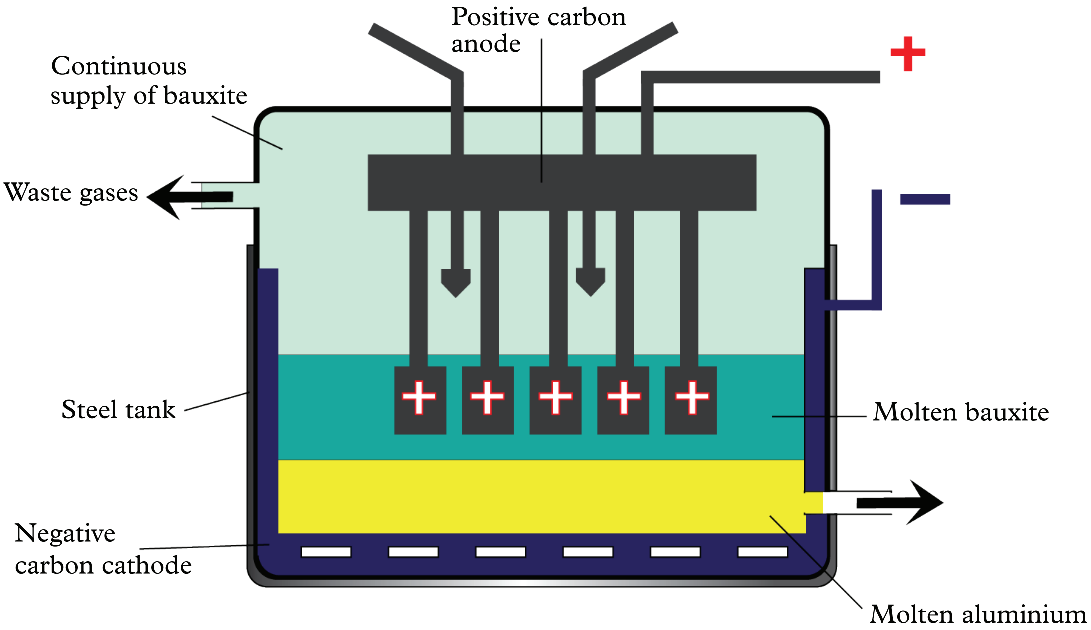


Fig 8.10 Hall's cell for electrolysis of aluminium oxide

Reaction at cathode (negative electrode)

Aluminium ions + 3 electrons ⟶ Molten aluminium

$4Al^{3+}(aq) + 12e^{-} \longrightarrow 4Al\ (l)$

Molten aluminium will collect at the bottom near cathode.

Reaction at anode (negative electrode)

Oxide ions ⟶ oxygen gas + 12 electrons

$6O^{2-} \longrightarrow 3O_2 + 12e^{-}$

Oxygen gas will be discharged at anode.

Overall reaction:

Aluminium oxide ⟶ Molten aluminium + oxygen gas

$4Al_2O_3\ (l) \longrightarrow 8Al\ (l) + 6O_2\ (g)$

2. Refining of metals

Copper is an excellent conductor of electricity. It is therefore used widely as a conductor in various electrical appliances and electrical lines. Copper extracted from its ores is not pure enough to be used for electrical wiring. It must first be purified.

It is purified (refined) using electrolysis as shown in Fig. 8.11 before use.

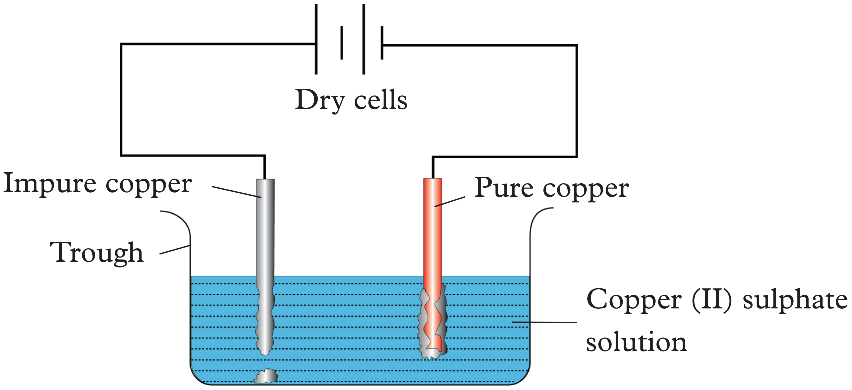


Fig 8.11: Refining of impure copper

During refining of copper, the following must be done:

- the cathode must be made of thin sheet of pure copper.
- the anode is made of impure copper.
- the electrolytic solution must contain copper (II) ions e.g. copper (II) 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 (–ve), accept two electrons and get discharged. They then get deposited on the pure cathode as solid copper.

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

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. This sludge contains valuable metals like silver and gold.

3. Electroplating of metals

Electroplating is a process whereby a thin layer of desired material is applied on a required metal of interest. Electroplating is mostly done on stainless steel to prevent rusting. It is also done on some objects to make them look shiny and attractive.

Examples

(i) Most metal objects are made of iron or steel. To prevent rusting, they are nickel–chromium plated e.g. bicycle handlebars, car rims, among others.

(ii) Silver or gold plating is done e.g, on spoons, plates and jewellery items to improve their appearance.

(iii) Tin cans – Cans made of steel are electroplated using a thin layer of tin to prevent them from rusting.

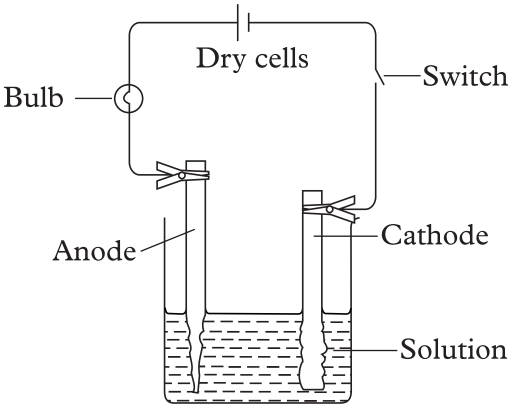


Fig. 8.12: Conventional electrolytic cell that can be used for electroplating

To coat the item, the following procedure is used.

- First, the item to be electroplated is thoroughly cleaned.
- An electrolyte whose ions are required to be deposited is selected.
- Direct current (DC) is used instead of alternating current (AC).
- The item to be coated must be made the cathode. Metal for coating must be made the anode.

Manufacture of sodium hydroxide, chlorine and hydrogen

The main raw material is concentrated sodium chloride (brine). Electrolysis of brine yields sodium hydroxide, chlorine and hydrogen. Fig. 8.13 shows the set-up for the electrolysis process.

In this industrial process, mercury is made to flow along the cell as shown in Fig. 8.13. This mercury acts as the cathode from where sodium is discharged.

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

Graphite is made the anode because it is an inert electrode hence resists corrosion by chlorine which is discharged and collected at the anode i.e.

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

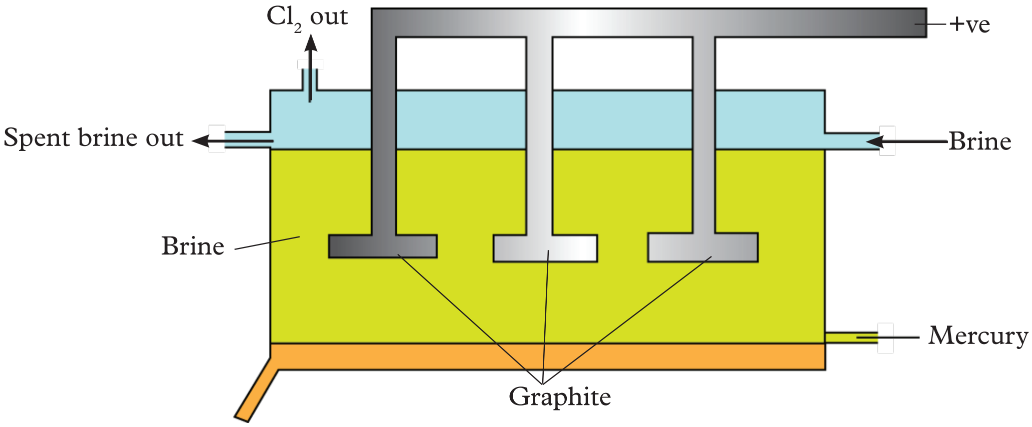


Fig. 8.13: The flowing mercury cell for the manufacture of sodium hydroxide, chlorine and hydrogen

Sodium formed at the cathode dissolves in mercury forming sodium-mercury amalgam (NaHg). The amalgam is reacted with water to form sodium hydroxide solution and hydrogen gas.

$$\underset{\text{sodium-mercury amalgam}}{2NaHg\,(l)} + 2H_2O\,(l) \longrightarrow 2NaOH\,(aq) + H_2\,(g) + 2Hg(l)$$

The mercury produced is recycled. Chlorine, sodium hydroxide and hydrogen are packaged in readiness for use.

Note: Mercury is a heavy toxic metal. It can cause heavy metal poisoning to animals and humans. We should therefore ensure that all measures are taken to avoid any traces of mercury finds its way to the environment through the effluents from industries or any other means.

Self-evaluation Test 8.3

1. Complete the passage using the words below.
   (*Anode*, *cathode*, *current*, *electrodes* and *electrolyte*)
   Electrolysis occurs when an electric _________ passes through a molten _________. The two rods dipping into the electrolyte are called _________. Metals are deposited at the _________.
2. Complete the passage using appropriate words.
   Aluminium is extracted by electrolysis of molten aluminium oxide _________ in _________. This mixture melts at much _________ temperature than _________ aluminium oxide. Therefore a lot of _________ is saved and the electrical _________ of the electrolyte is improved.
3. Predict the products formed at the anode and cathode when the following molten compounds are electrolysed.
   (a) Sodium fluoride
   (b) Copper (II) iodide
   (c) Zinc bromide
4. What type of compounds can be electrolysed? What conditions are needed for this electrolysis to take place?

Electroplating

Activity 8.7

Material and apparatus

- 1.5-volt D battery with battery holder
- Two crocodile clip leads or insulated wire
- Beaker or glass
- Copper (II) sulphate
- Copper electrode (or coil of copper wire)
- Spoon
- Safety equipment

Procedure

1. Prepare the spoon for copper-plating by cleaning it with toothpaste or soap and water. Dry it off on a paper towel.
2. Stir copper sulphate into some hot water in a beaker until no more will dissolve. Your solution should be dark blue. Let it cool.
3. Use one alligator clip to attach the copper electrode to the positive terminal of the battery (this is now the anode) and the other to attach the spoon to the negative terminal (now called the cathode).
4. Partially suspend the spoon in the solution by wrapping the wire loosely around it and placing it across the mouth of the beaker. The wrapping wire should not touch the solution.
5. Place the copper electrode into the solution, making sure it does not touch the spoon and the solution level is below the alligator clip. An electrical circuit has now formed and current is flowing.
6. Leave the circuit running for 20-30 minutes.

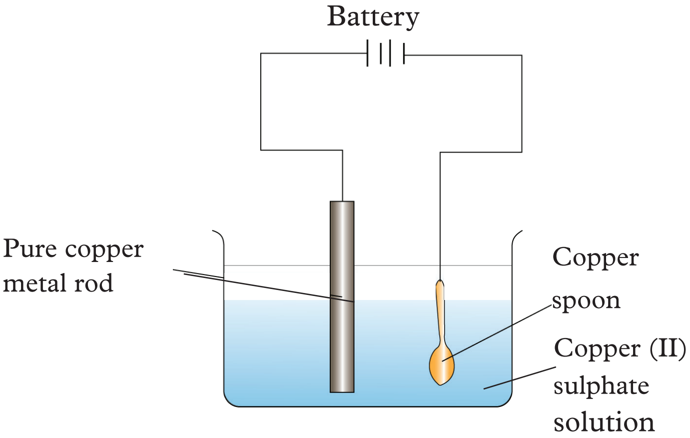


Fig 8.14: Electroplating using copper plated spoon

Study questions

1. Explain what happens at each electrode.
2. Write equations for the reaction that take place at each electrode.

3. What is the colour of the spoon at the end of the experiment?
4. Explain why the electrolyte process was continuous?

Discussion corner!

1. Discuss in pairs the observations made during the experiment and the study questions above.
2. Write a report and present your finding in class.

I have discovered that…

Copper sulphate solution is an electrolyte that conducts electricity. When electric current is flowing, oxidation (loss of electrons) takes place at the copper anode, adding copper ions to the solution. Those ions move to the cathode, where reduction (gain of electrons) takes place, plating or depositing the copper metal onto the spoon. There were already copper ions present in the copper sulphate solution before the reaction, but the oxidation reaction at the anode kept replacing them in the solution as they were plated onto the key, keeping the reaction going.

The facts

The coating of a metal object with another metal object is called **electroplating**. It is carried out in a cell called plating bath. It contains an electrolyte. For silver plating electrolyte is a solution of silver salt. The article to be plated is made the cathode in the cell so that metal ions move to it when inserted in electrolyte.

Reaction at cathode (negative electrode)

Copper ions + $2e^-$ $\longrightarrow$ Copper atoms.

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$ Copper metal deposit at cathode.

Reaction at anode (positive electrode)

Copper atoms dissolve from anode $\longrightarrow$ Copper ions in the solution + $2e^-$

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

Copper metal of anode will dissolve and added in electrolyte.

Self-evaluation Test 8.4

1. Complete the passage using the words below:
   (*Anode, cathode, electrolyte, plate, solution*)
   To electroplate an object you make the object to be electrolysed the _______.
   The _______ is the metal that will _______ the surface.
   The _______ is a _______ of the salt compound of the plating metal.
2. Using half equations explain how silver can be used to electroplate an object.
3. You want to electroplate a piece of copper with chromium. Draw a labelled diagram of the apparatus you could use.

Unit summary 8

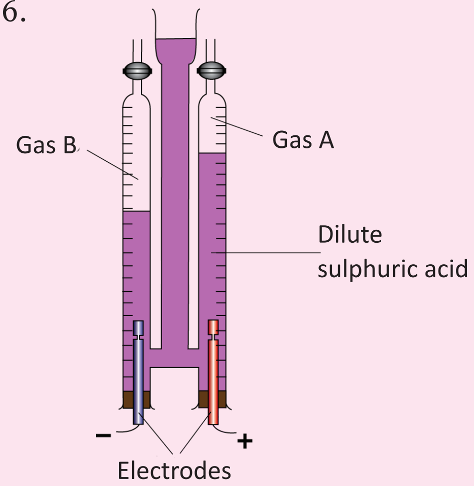
- Electrolysis is the process in which a chemical change, especially decomposition, is brought about by passing an electric current through a molten or solution of electrolytes and the electrolyte's ions move toward the negative and positive electrodes and are selectively discharged.
- Oxidation is the loss of electrons during a reaction by a molecule, atom or ion.
- Reduction is gaining of electrons during a reaction by a molecule, an atom or an ion.
- An oxidation-reduction reaction is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron.
- Oxidation number, also called oxidation state, is the net change that an atom should have if the compound was ionic.
- Oxidising agent is a substance that tends to bring about oxidation by being reduced and gaining electrons.
- Reducing agent is a substance that tends to bring about reduction by being oxidized and losing electrons.
- Cathode is the negatively charged electrode by which electrons enter an electrical device.
- Anode is the positively charged electrode by which the electrons leave the conducting device.
- Electroplating is the process of coating layer of a metal onto another by electrolysis. It is most commonly used for decorative purposes or to prevent corrosion of a metal.

Test your Competence 8

1. Which of the following statements concerning electrolysis is incorrect?
   A. During electrolysis, chemical energy is converted into electrical energy.
   B. Electrolysis involves decomposition of an electrolyte.
   C. An energy supply is required for electrolysis.
   D. Graphite is commonly used as the inert electrodes in electrolysis.
2. Which of the following is not an essential component of the set-up for electrolysis?
   A. Electrolyte
   B. Electrodes
   C. Ammeter
   D. Energy supply
3. Which of the following substances would not be decomposed by electricity?
   A. Molten silver
   B. Dilute sulphuric acid
   C. Concentrated sodium chloride solution
   D. Aqueous silver nitrate
4. Which of the following ions are reduced during electrolysis of molten silver chloride?
   A. Silver ions
   B. Chloride ions
   C. Hydrogen ions
   D. Hydroxide ions
5. Which of the following sets of changes would occur during the electrolysis of dilute sulphuric acid using platinum electrodes?

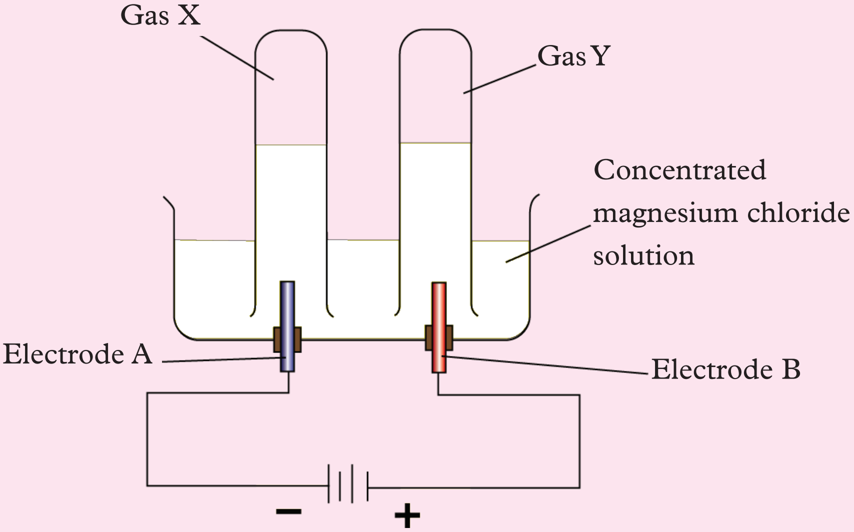
Product at cathode	Product at anode	Concentration of sulphuric acid
A Oxygen	Hydrogen	Increases
B Hydrogen	Oxygen	Increases
C Oxygen	Hydrogen	Decreases
D Hydrogen	Oxygen	Decreases

6.



Which of the following statements concerning the set-up alongside is incorrect?
   A. Only hydroxide ions migrate to the anode.
   B. The negative electrode is the cathode.
   C. Gas A is oxygen.
   D. The electrodes can be made of platinum.

7.



Which of the following combinations is correct?

	Gas X	**Gas Y**
A.	Hydrogen	Oxygen
B.	Oxygen	Hydrogen
C.	Hydrogen	Chlorine
D.	Chlorine	Hydrogen

8. The products of the electrolysis of dilute magnesium iodide solution using carbon electrodes are

A. Hydrogen and iodine
B. Hydrogen and oxygen
C. Magnesium and iodine
D. Magnesium and oxygen

9. Which of the following substances, when electrolyzed using graphite electrodes, will liberate gaseous products at both electrodes?

A. Molten sodium chloride
B. Sodium bromide solution
C. Zinc sulphate solution
D. Silver nitrate solution

10. (i) The (higher/lower) ________ the concentration of an ion in a solution, the greater is the probability of its being discharged at its appropriate electrode.

(ii) Cations are formed by ________ (Loss/gain) of electrons and anions are formed by ________ (Loss/gain) of electrons.

(iii) Electrolysis is the passage of ________ (electricity/electrons) through a liquid or a solution accompanied by a ________ (physical/chemical) change.

11. (a) Explain why solid sodium chloride does not allow electricity to pass through.

(b) What kind of particles will be found in a liquid compound that is a non-electrolyte?

(c) If HX is a weak acid, what particles will be present in its dilute solution apart from those of water?

(d) Which ions must be present in a solution used for electroplating a particular metal?

(e) Explain how electrolysis is an example of redox-reaction?

12. Choose A, B, C or D to match the descriptions (i) to (v) below. Some alphabets may be repeated. [A – non-electrolyte; B – strong electrolyte; C – weak electrolyte; D – metallic conductor]

(i) Molten ionic compound

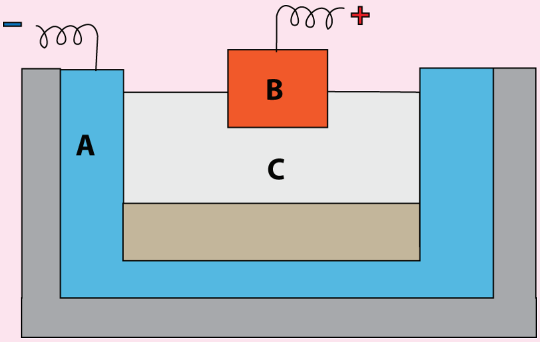
(ii) Carbon tetrachloride

(iii) An aluminium wire

(iv) A solution containing solvent molecules, solute molecules and ions formed by the dissociation of solute molecules.

(v) A sugar solution with sugar molecules and water molecules.

13. (a) The following is a sketch of electrolytic cell used in extraction of aluminium.



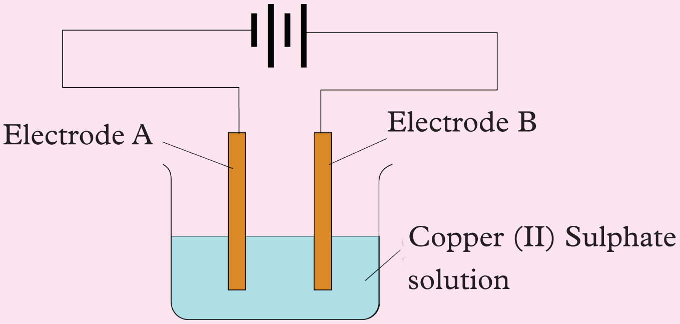
(a) What is the substance of which the electrodes of A and B are made?

(b) At which electrode (A or B) is the aluminium formed?

(c) What are the two aluminium compounds in the electrolyte C?

(d) Why is it necessary for electrode B to be continuously replaced?

14. Study the diagram given and answer the questions that follow:



(a) Give the names of the electrodes A and B.

(b) Which electrode is the oxidizing electrode?

15. (a) Electrons are getting added to an element Y.

(i) Is Y getting oxidized or reduced?

(ii) What charge will Y have after the addition of electrons?

(iii) Which electrode will Y migrate to during the process of electrolysis?

(b) (i) Elements Q and S react together to form an ionic compound. Under normal conditions, which physical state will the compound QS exist in?

(ii) Can Q and S, both be metals? Justify your answer.

16. (a) What is an electrolyte?

(b) Classify the following substances under three headings: Strong electrolytes, weak electrolytes, non electrolytes

Acetic acid, ammonium chloride, ammonium hydroxide, carbon tetrachloride, dilute hydrochloric acid, sodium acetate, dilute sulphuric acid

17. Choosing only words from the following list, write down the appropriate words to fill in the blanks: From (a) to (b) below: [*anions, anode, cathode, cations, electrode, electrolyte, nickel, and voltammeter*.]

To electroplate an object with nickel requires an (a) _________ which must be a solution containing (b) _________ ions. The article to be plated is placed as the (c) _________ of the cell in which the plating is carried out. The (d) _________ of the cell is made from pure nickel.

UNIT 9

Structure and properties of alkenes and alcohols

Key unit competency

To be able to relate the properties of alkenes and alcohols to their functional groups.

Learning objectives

By the end of this unit, I should be able to:

- Describe the structures of alkenes, alcohols and their physical properties.
- Name alkenes and alcohols using the IUPAC system of naming - up to C5 and state their chemical properties.
- Describe the uses and misuse of alcohols and alkenes in daily life.
- Describe the fermentation process.
- Describe the hydration of alkenes as a method of preparing alcohols.
- Prepare alcohol by a fermentation process.
- Draw the structures of alkenes and alcohols.
- Carry out an experiment to confirm the presence of alkenes.
- Develop observation skills in testing for alkenes with bromine water.
- Develop a culture of working in a team during discussions and experiments as well as develop orderliness in work to present results.
- Respect procedures in an experiment and develop self confidence in the presentation of research work and presentations.
- Appreciate the social effects of the misuse of alcohol.

Mind teaser

Study the pictures below. Describe how you have interacted with them in your day today activities.

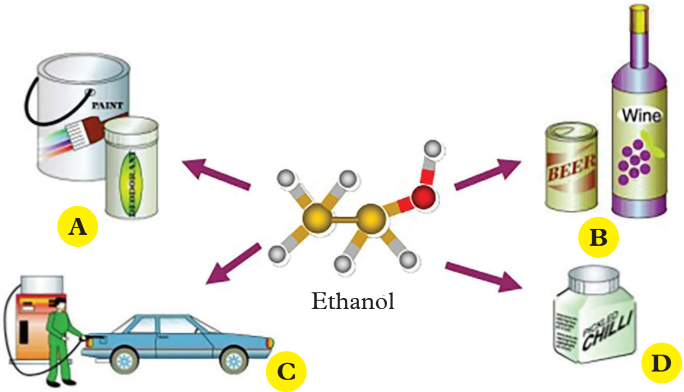


Fig 9.1

How are the things in the picture related to ethanol?

What category of compounds does ethanol belong to? Can you name other compounds related to ethanol and their uses?

Introduction

Ethene is an example of alkene compounds. Alkenes are a type of hydrocarbon characterised by the presence of a double bond or several of them in their structure. The double bond(s) gives alkenes their characteristic properties. It is an example of a homologous series of hydrocarbons. Another homologous series is **alcohols** (or alkanols). Alcohols are characterised by presence of hydroxy (–OH) group in its structure, which is also responsible for alcoholic properties.

9.1 Nomenclature of alkenes and their structures

Activity 9.1

Requirements

- Molecular model kits
- Diagrams
- Pictures

Procedure

1. In groups model the structures of alkenes using the molecular model kits provided.
2. Refer to the charts provided for names of the alkene modelled.
3. Compare your findings with those of other groups.

Study questions

1. What is the general formula for alkenes?
2. Draw the structure of alkenes up to carbon-5.
3. How are alkenes prepared?

Discussion corner!

1. Discuss the structures of alkenes modelled above?
2. Present your drawn structure of alkene to the teacher for assessment.

I have discovered that…

The general formula of alkene is C_nH_{2n}

The facts

Alkenes are in a family of hydrocarbons containing covalent carbon-carbon double bond (>C=C<) as well as the single bonds formed between carbon and carbon (C-C) and carbon and hydrogen (C-H).

The C=C double bond is called the functional group of alkenes.

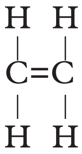
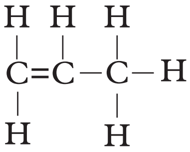
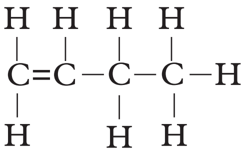
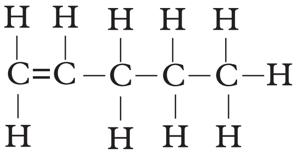
Alkenes are homologous series of organic compounds with the general formula C_nH_{2n} where n = 2, 3, 4, etc. For example: C_2H_4, C_3H_6, C_4H_8.

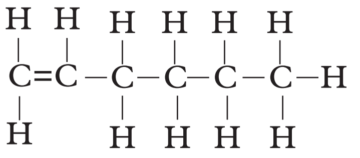
In the general formula n = number of carbon atoms in the alkene molecule (n = 2, 3, 4 etc.) and from the general formula you can deduce the number of hydrogen atoms, hence the complete molecular formula for alkene with a carbon chain containing one double bond only. This general formula only applies to hydrocarbon alkenes with one C=C double bond.

An easy method of naming alkenes may be done by replacing the 'a' in alkanes with an 'e'.

Note: There are no alkene where n = 1. Why do you think so?

Table 9.1 The first five members of alkenes

Number of carbon	**Name**	**Formula C_nH2_n**	**Structural formula**	**Condensed formula of the structure**
2	Ethene	C_2H_4	 $\begin{matrix} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & H \end{matrix}$	$CH_2 = CH_2$
3	Propene	C_3H_6	 $\begin{matrix} H & & H & & H & & \\ \vert & & \vert & & \vert & & \\ C & = & C & - & C & - & H \\ \vert & & & & \vert & & \\ H & & & & H & & \end{matrix}$	$CH_2 = CHCH_3$
4	But-1-ene	C_4H_8	 $\begin{matrix} H & & H & & H & & H & & \\ \vert & & \vert & & \vert & & \vert & & \\ C & = & C & - & C & - & C & - & H \\ \vert & & & & \vert & & \vert & & \\ H & & & & H & & H & & \end{matrix}$	$CH_2 = CHCH_2CH_3$
5	Pent-1-ene	C_5H_{10}	 $\begin{matrix} H & & H & & H & & H & & H & & \\ \vert & & \vert & & \vert & & \vert & & \vert & & \\ C & = & C & - & C & - & C & - & C & - & H \\ \vert & & & & \vert & & \vert & & \vert & & \\ H & & & & H & & H & & H & & \end{matrix}$	$CH_2 = CHCH_2CH_2CH_3$

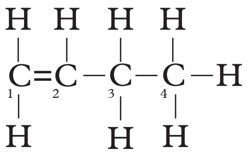
6	Hex-1-ene	C_6H_{12}		$CH_2 = CHCH_2CH_2CH_2CH_3$

Look at Table 9.1. Check on the structural formulae of alkenes. What difference do you notice from one formula to the next?

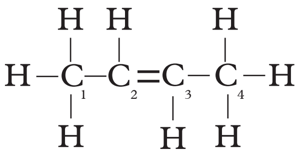
They differ by a–CH_2– group, hence form a homologous series.

IUPAC naming of alkenes

The IUPAC rules used in naming alkanes can also be used in systematic naming for straight chain alkenes. The rules demand that the position of the double bond must be included in the name. This is why number 1 is included in some alkenes in Table 9.1. It tells us where the double bond is placed in that particular alkene as shown in the examples below.

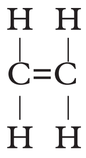


but–1–ene

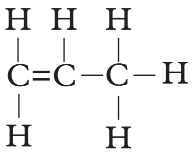


but – 2 – ene

The number shows the position of the double bond, and we write the lower number of the two carbon atoms bearing the double bond. In case of an alkene like ethene or propene the number is not necessary because if you number the carbon chain from left or right hand side the double bond will be on the 1st carbon.



ethene



propene

Note: In alkenes, you must count the number of bonds around each carbon atom properly to ensure each carbon atom is surrounded by maximum of 4 bonds.

Now let us discuss the other IUPAC rules of naming alkenes and see how we apply them.

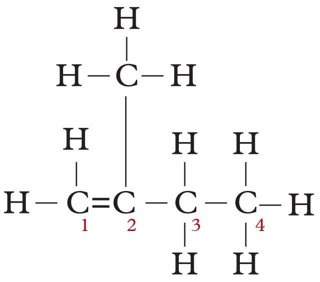
Rules for naming alkenes

Research activity

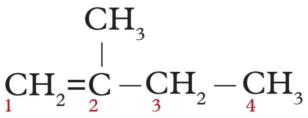
1. Using the internet and books from the library, research on the rules that are applied when naming alkenes.
2. In groups, use the findings to name some alkenes.
3. Compare your findings with the discussion below.

1. Determine the longest continuous chain of carbon atoms that has a double bond between two carbon-atoms. This will give us the parent or ending name.
2. Number the carbon atoms in the chain so that the double bond is between the two carbon atoms assigned the possible lowest numbers. This means that you have to decide whether to number beginning on the right or left end of the longest identified chain. Always start from the side that gives the double bond the lowest number.
3. Identify the various branching groups or substituent groups attached to this continuous chain of carbon atoms.
4. If more than one member of a substituents groups is attached to the same carbon atom, we write the position number to the number of times substituent group appears separated by a comma then a hyphen then a prefix denoting the number of times the group is attached such as di, tri, followed by the substituent group name and lastly the name of the parent alkene, for example, 3,3-dimethylbutene etc.
5. Write the lowest number of the carbon atoms bearing the double bond.

Example of systematic naming of some alkene compounds.



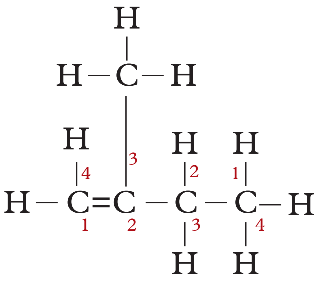
or



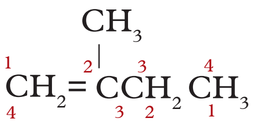
$$\underset{1}{CH_2}=\underset{2}{\overset{\overset{\displaystyle CH_3}{|}}{C}}-\underset{3}{CH_2}-\underset{4}{CH_3}$$

1. Identify the longest continuous chain of carbon atoms with the double bond. The longest carbon chain has four carbon atoms. Therefore the parent name is butene.

2. Number the carbon chain so that the double bond is between the lowest numbered carbon atoms.



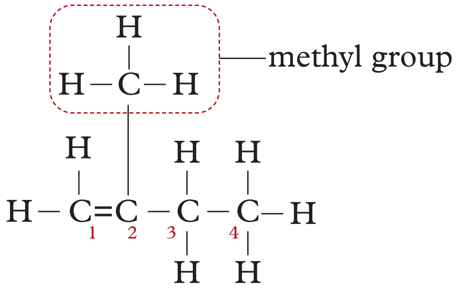
or



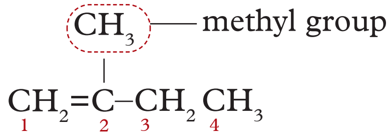
$$\underset{4}{\overset{1}{CH_2}}=\overset{\overset{\displaystyle CH_3}{2|}}{C}\underset{3\ \ 2}{\overset{3}{CH_2}}\underset{1}{\overset{4}{CH_3}}$$

- Numbering from left to right would place the double bond between carbon number 1 and carbon number 2. This means that we must number the carbon chain from left to right to assign the possible lowest number.

3. Identify and locate all branched groups, attaching a prefix number to which the branch is attached.



or



$CH_2=\overset{CH_3}{C}-CH_2CH_3$

- Here, we have a methyl group attached to carbon number 2.

4. Write the number of the carbon atoms bearing alkyl group first followed by a hyphen then the name of the alkyl group.
5. We must also include the lowest number in the alkene name to indicate where the double bond is placed. In this case, we insert 1 in between the parent name as follows:

   2 –methylbut –1– ene

Therefore, 2 tells us the position of the substituent group, i.e the methyl group is attached on the second carbon atom. 1 tells us that the double bond starts on the 1st carbon, so it must be between the first and second carbon.

Isomerism

Research activity

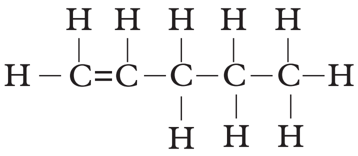
1. In your study groups, review isomerism in alkanes.
2. Use the working knowledge in (1) above to try write isomerism in alkenes.
3. Compare your findings with the content below.

In previous classes you learnt about chain isomerism in alkanes. Here you are going to learn about two types of structural isomerism in alkenes.

They are:

- chain isomerism.
- position isomerism.

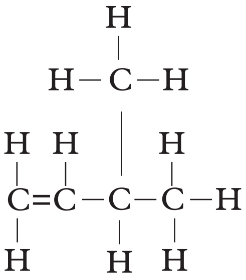
Let us consider the structure of pentene to study the two types of isomerism.



1. Chain isomerism

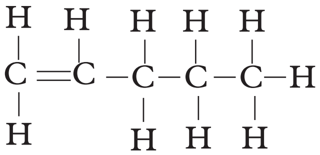
It is possible to draw other different structural formulae of pent-1-ene. The atoms of carbon are arranged in a different way in each structure as shown below.

(a)



3 – methylbut–1–ene

(b)

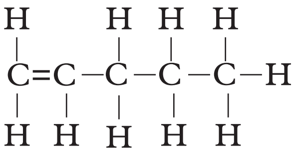


pent–1–ene

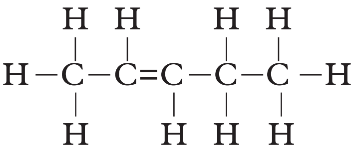
The above compounds (a) and (b) have the same molecular formula (i.e. C_5H_{10}), but they have different structural formulae. The general name of such structures is **isomers**. In 3 – methylbut–1–ene and pent–1–ene, the position of the double bond remains the same but the length of the chain is different. Such isomers are known as **chain isomers**.

2. Position isomerism

Sometimes the position of a double bond in an alkene can be shifted to a different position within the molecule. This type of isomerism is known as **position isomerism**. This is shown in the following examples.



Pent-1-ene

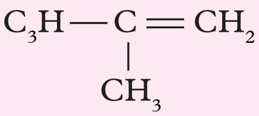


Pent-2-ene

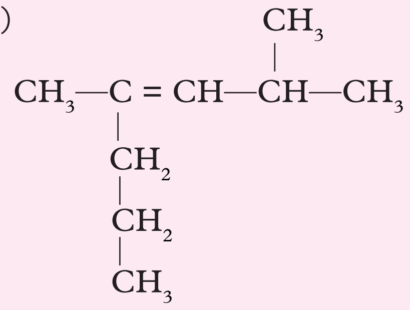
Note that in the above isomers, the position of the double bond shifts while the position of the carbon atoms remains the same.
Such isomers are known as **position isomers**.

Self-evaluation Test 9.1

1. Name the following alkenes.
   (a) $CH_3CH_2CH = CH_2$

   (b) 
   $C_3H—\underset{\displaystyle CH_3}{\underset{|}{C}}=CH_2$

2. Write the structural formulae for the following alkenes.
   (a) hex – 2 – ene
   (b) 2,3 – dimethylbut – 2 –ene
3. (a) What are isomers?
   (b) Is there any difference between structural isomers and position isomers?
   (c) Draw and name all structural and position isomers of hexene.
4. Name the following compounds.
   (a)



$$CH_3-\underset{\substack{|\\CH_2\\|\\CH_2\\|\\CH_3}}{C}=CH-\overset{\substack{CH_3\\|}}{CH}-CH_3$$

   (b) $CH_3CH = CHCH_3$
   (c) $CH_3(CH_2)_2CH = CH_2$
5. What is the main feature in the bonding of an alkene? Why is the first member of this series not methene?
6. Draw the structural formulae of the following compounds.
   (a) But – 2 –ene
   (b) 2 – methylprop-1-ene
   (c) 3,3 – dimethylbut –1 –ene
7. What is the difference between saturated and unsaturated hydrocarbons?

9.2 Sources of alkenes

There are two sources of alkenes.

(a) Dehydration of alcohols

(b) Cracking of hydrocarbons

Preparation and properties of alkenes

(a) Dehydration of ethanol

Laboratory preparation of ethene

Activity 9.2

Apparatus and chemicals

- Thermometer
- Round-bottomed flask
- Sand bath
- Boiling tube
- Cork rubber stopper
- Bunsen burner
- Delivery tubes
- Trough
- Beehive shelf
- Gas jars
- Concentrated sulphuric acid
- Ethanol
- Bromine water
- Acidified potassium permanganate
- Broken porcelain / dry sand.

Caution: Concentrated sulphuric acid is corrosive. Ethanol and ethene are flammable. There is substantial production of sulphur dioxide in this reaction which can cause irritation and even trigger an asthmatic attack. This experiment should therefore be done in a fume chamber or in open space.

Procedure

1. Place about 20 cm^3 of ethanol in a round-bottomed flask and slowly add about $40cm^3$ concentrated sulphuric acid, while cooling and shaking the flask.

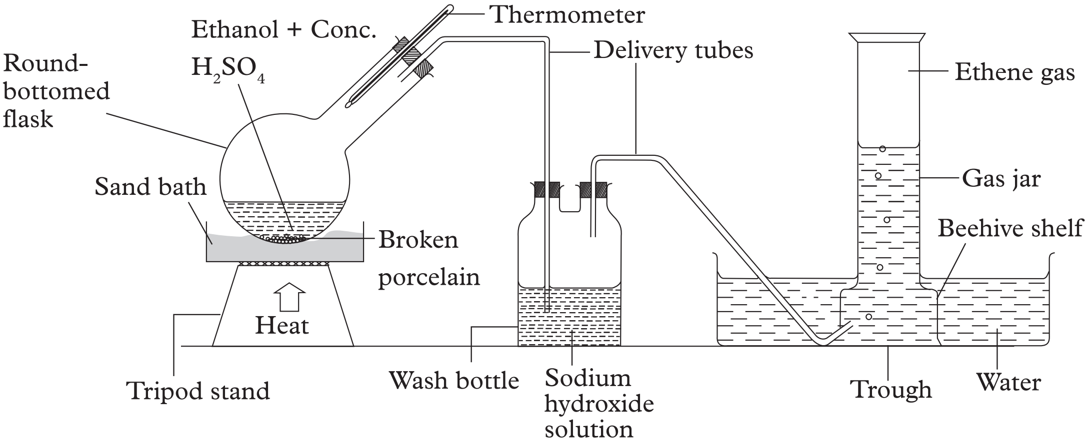


Fig. 9.2: Laboratory preparation of ethene gas

2. To the round-bottomed flask, add 2-3 g of clean dry broken pieces of porcelain or sand and assemble the apparatus as shown in Fig 9.2
3. Heat the flask gently over a sand bath. What do you observe?
4. Allow the gas produced to escape from the delivery tube for a few minutes then collect samples of this gas in a boiling tube then over water as shown. Close the tube with a cork, then remove the cork and light the gas.

   If the tube contains ethene, it burns with a blue flame; if ethene is mixed with air, the mixture ignites with an explosion. Allow more gas to escape from the delivery tube until you prove that it is pure ethene. Collect several gas jars of ethene.
   - The first portion of the gas collected is a mixture of displaced hot air from apparatus and gas evolved.
   - Ethene is flammable and it is important not to ignite the gas directly from the delivery tube.
5. Disconnect the delivery tube connecting the round-bottomed flask wash bottle before you stop heating.
   - Why must the tube be disconnected before you stop heating?
6. Bubble portion of the gas collected through the test tubes containing:
   (a) bromine water
   (b) acidified potassium permanganate. Record your observations.
      - What do you conclude?
7. The residue in the flask should be disposed off by first allowing it to cool, and then pour it into a large volume of cold water.

Study questions

1. Why is sulphuric acid added into the flask while cooling and shaking the glass?
2. Why should clean, dry broken porcelain or sand be added into the mixture of conc. sulphuric acid and ethanol?
3. Why is it advisable to heat the acid gently?
4. Why does the gas ignite with an explosion if it is not pure ethene?
5. Why must the delivery tube be disconnected from the round-bottomed flask before the heat is turned off?
6. Why should the residue in the flask be first poured into a large volume of cold water when disposing it?

Discussion corner!

1. In your study groups discuss the observations made during the experiment and the study questions above.
2. Write a report of your findings and present in class.

The facts

When a mixture of ethanol and concentrated sulphuric acid is heated gently an oily liquid called ethyl hydrogensulphate and water are formed.

Ethanol + conc. Sulphuric acid ⟶ ethyl hydrogensulphate + water

$$C_2H_5OH(l) + H_2SO_4(aq) \longrightarrow C_2H_5HSO_3H(l) + H_2O(l)$$

When the mixture is heated to 170°C, a steady stream of ethene is produced. The ethyl hydrogensulphate formed is unstable. It decomposes to sulphuric acid and ethene gas.

Ethyl hydrogensulphate ⟶ sulphuric acid + ethene

$$C_2H_5HSO_3H(l) \longrightarrow H_2SO_4(aq) + C_2H_4(g)$$

The above chemical reactions are equivalent to the dehydration of ethanol by concentrated sulphuric acid and can be summarised as shown below.

$$\text{Ethanol} \xrightarrow[170°\text{C}]{\text{Conc. } H_2SO_4} \text{ethene + water}$$

$$CH_3CH_2OH(l) \xrightarrow[170°\text{C}]{\text{Conc. } H_2SO_4} CH_2 = CH_2(g) + H_2O\ (l)$$

Ethanol contains elements that make up water i.e. hydrogen and oxygen hence water is removed in the above reaction. When concentrated sulphuric acid reacts with this water, a lot of heat is produced i.e the reaction is extremely exothermic.

It is therefore necessary to cool the flask while this reaction is being carried out. Ethanol and concentrated sulphuric acid are miscible although they have different densities. Shaking the flask helps to mix the two liquids properly. Dry broken porcelain increase the surface area on which gas bubbles can form. They also ensure smooth boiling of the mixture.

The mixture of ethanol and the acid is heated gently to prevent a violent reaction which can cause large amounts of sulphuric acid to jump off the flask into the delivery tube. Pure ethene burns with a blue flame, otherwise if mixed with air, it burns with an explosion. That is why the first portion of the gas collected is allowed to escape.

The delivery tube should first be disconnected from the round bottomed flask before we stop heating to prevent water from being sucked back into the flask. The sulphuric acid remaining in the flask as a residue should be diluted by pouring it into a large volume of cold water. This is to dilute the acid so as to make it less corrosive.

Note: By dehydrating different alcohols, different alkenes are produced. The general equation for the production of alkenes through dehydration of alcohols is:

$$C_nH_{2n+1}OH \xrightarrow{\text{Conc. } H_2SO_4} C_nH_{2n} + H_2O$$

Cracking hydrocarbons in industry

Alkenes can be obtained from alkanes by heating crude oil after refining. This process is called **cracking**. Cracking can be done in two ways i.e thermal cracking (heating) or catalytic cracking (heating with a catalyst).

Cracking involves breaking long-chain alkane molecules into a shorter alkane and an alkene. For example, when hexane is cracked butane and ethene are obtained. After distilling the crude oil, too many large hydrocarbons are left behind. The smaller hydrocarbons like petrol, are in great demand. Scientists have found a way to change the larger less useful molecules into smaller, more useful ones.

Cracking helps produce more petrol. The petrol obtained this way is of a better quality than obtained by distillation of crude oil. Cracked petrol is used to blend other petrol to improve the quality. When a long carbon chain alkane is cracked (split) at high temperature it produces a shorter alkane and an alkene.

$$\text{Long alkane} \xrightarrow[\text{heating}]{\text{thermal}} \text{Shorter alkane} + \text{Alkene}$$

Example

When a long molecule e.g. decane, $C_{10}H_{22}$, is cracked it produces octane and ethene.

$$\underset{\text{decane}}{C_{10}H_{22}} \longrightarrow \underset{\text{octane}}{C_8H_{18}} + \underset{\text{ethene}}{C_2H_4}$$

When a catalyst is used, cracking can be made to occur at fairly low temperatures. This is known as **catalytic cracking**.

Note: Cracking produces extra petrol and is as well a source of alkenes.

9.3 Properties of alkenes

Discussion corner!

1. You will be provided with charts on trends of physical and chemical properties of alkenes. Study the charts provided on trends. Alternatively, study Table 9.2 below.

Table 9.2: Trends in physical properties of alkenes

Number of alkene	Formula	Melting point (°C)	Boiling point (°C)	Density (g/cm³)	Physical state at room temperature	Solubility in water	Solubility in organic solvent
Ethene	C_2H_4	–169	–104	____	gas	insoluble	soluble
Propene	C_3H_6	–185	–47.7	____	gas	insoluble	soluble
But-1-ene	C_4H_8	–185	–6.2	____	gas	insoluble	soluble
Pent-1-ene	C_5H_{10}	–138	30.0	0.640	liquid	insoluble	soluble
Hex-1-ene	C_6H_{12}	–98	63.9	0.674	liquid	insoluble	soluble

2. In pairs discuss the trends in properties of alkene.
3. Compare your findings with those of other groups.

I have discovered that…

Alkenes have varying physical properties. However most of them are soluble in organic solvents but insoluble in water. Those with fewer crabon atoms have lower melting and boiling points while their densities also depend on the length of the carbon atoms.

The facts

Alkenes have similar chemical properties and show a gradual change in physical properties.

- There is increased melting and boiling points as the number of carbon atoms increase. This is because with increasing carbon atoms the molecular mass increases. This also causes an increase in the intermolecular forces of attraction. To break these intermolecular forces, more energy is required hence the increase in melting and boiling points.

- The first three alkenes, i.e ethene, propene and but –1–ene are gases at room temperature while pent-1-ene and hex-1-ene are liquids. This also has to do with increase in intermolecular forces of attractions.
- Alkenes are organic compounds hence they are insoluble in water but soluble in organic solvents.

Physical properties of ethene

i. Colourless gas.
ii. Has a pleasant odour (sweet smell).
iii. Completely insoluble in water but soluble in organic solvents.
iv. It boils at –102°C and melts at –169°C.
v. It has anaesthetic effect on inhalation.

Alkenes readily burn, just like alkanes, to give carbon dioxide and water if combustion is complete e.g.

Alkene + oxygen ⟶ carbon dioxide + water

- Ethene + oxygen ⟶ carbon dioxide + water

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$$

- Propene + oxygen ⟶ carbon dioxide + water

$$C_3H_6(g) + 4\tfrac{1}{2}O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(l)$$

or $2C_3H_6(g) + 9O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$

However, they are not used as fuels for two reasons.

1. They are far too valuable for use to make plastics, anti–freeze and numerous other useful compounds.
2. They burn with a more smoky flame than alkanes.

Chemical properties of ethene

A double bond between carbon atoms is a stronger link than a single bond. However, the second bond is weaker than a single bond i.e in terms of energy needed to break bonds.

As such, the second bond is easier to break. This bond easily opens up and takes additional atoms for example, those of chlorine or bromine, making the unsaturated carbon atoms become saturated.

Let us now consider the some reactions of ethene.

1. Combustion of ethene

Ethene is highly flammable and burns readily in air to form carbon (IV) oxide and water.

Ethene + oxygen ⟶ carbon(IV) oxide + water

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$$

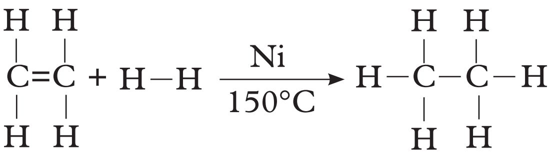
2. Addition reactions

(i) Addition of hydrogen to ethene

Ethene reacts with hydrogen in the presence of a nickel catalyst to form ethane. A temperature of 150°C is required.

$$\text{Ethene} + \text{hydrogen} \xrightarrow[150^\circ\text{C}]{\text{Nickel catalyst}} \text{ethane}$$

$$C_2H_4(g) + H_2(g) \xrightarrow[150^\circ\text{C}]{\text{Nickel catalyst}} C_2H_6(g)$$



$$\begin{array}{c}H\quad H\\ |\quad\ |\\ C{=}C\\ |\quad\ |\\ H\quad H\end{array} + H{-}H \xrightarrow[150^\circ\text{C}]{\text{Ni}} H{-}\begin{array}{c}H\\ |\\ C\\ |\\ H\end{array}{-}\begin{array}{c}H\\ |\\ C\\ |\\ H\end{array}{-}H$$

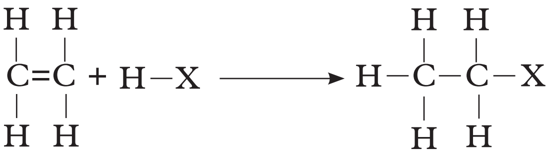
This process is also called **hydrogenation**.

(ii) Addition of hydrogen halides (HX) to ethene

The general formula of hydrogen halides is HX.

X represents any atom of a halogen such as chlorine, bromine or iodine.

The reaction of ethene with a hydrogen halide is as follows:



$$\begin{array}{c}H\quad H\\ |\quad\ |\\ C{=}C\\ |\quad\ |\\ H\quad H\end{array} + H{-}X \longrightarrow H{-}\begin{array}{c}H\\ |\\ C\\ |\\ H\end{array}{-}\begin{array}{c}H\\ |\\ C\\ |\\ H\end{array}{-}X$$

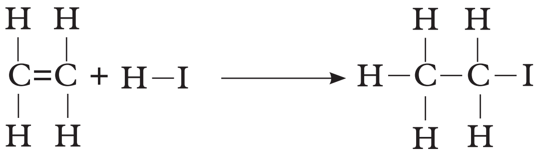
Hydrogen halides readily add across the double bond of alkenes to form alkyl halides. The reactivity of ethene with the hydrogen halides is as follows: HI > HBr > HCl. Ethene reacts readily with hydrogen iodide and with hydrogen bromide at room temperature to form iodoethane and bromoethane respectively.

Note: '>' denotes more reactive than

$$\text{Ethene} + \text{hydrogen iodide} \longrightarrow \text{iodoethane}$$

$$C_2H_4(g) + HI(aq) \longrightarrow C_2H_5I(g)$$

or

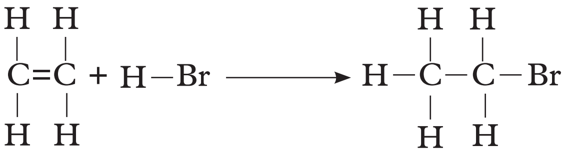


$$\begin{array}{c}H\quad H\\ |\quad\ |\\ C{=}C\\ |\quad\ |\\ H\quad H\end{array} + H{-}I \longrightarrow H{-}\begin{array}{c}H\\ |\\ C\\ |\\ H\end{array}{-}\begin{array}{c}H\\ |\\ C\\ |\\ H\end{array}{-}I$$

Ethene + hydrogen bromide $\longrightarrow$ bromoethane

$C_2H_4(g)$ + $HBr(aq)$ $\longrightarrow$ $C_2H_5Br(g)$

or



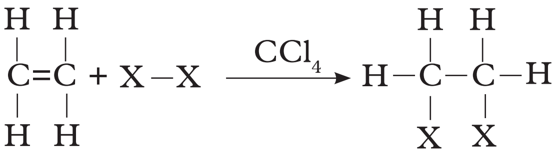
$$\underset{H}{\overset{H}{C}}{=}\underset{H}{\overset{H}{C}} + H{-}Br \longrightarrow H{-}\underset{H}{\overset{H}{C}}{-}\underset{H}{\overset{H}{C}}{-}Br$$

Note: With ethene, the hydrogen atom of the hydrogen iodide can add to either carbon atom to yield iodoethane. Follow the above procedure to write the equations and structures of the reaction between ethene with hydrogen chloride.

(iii) Addition of halogens to alkenes

Halogens readily add across a double bond of alkenes to form compounds called **dihalides**. This process is called **halogenation**. We may represent a halogen molecule with X_2 or $X - X$, for example Cl_2 or Br_2.

A general reaction of ethene with a halogen is shown below.



$$\underset{H}{\overset{H}{C}}{=}\underset{H}{\overset{H}{C}} + X{-}X \xrightarrow{CCl_4} H{-}\underset{X}{\overset{H}{C}}{-}\underset{X}{\overset{H}{C}}{-}H$$

Note: When naming the product formed, in which a hydrogen atom is substituted with a halogen atom, the prefix is borrowed from the name of the halogen where **ine** part is replaced with "o" as follows:

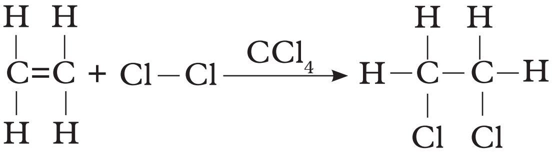
- from chlorine we get **chloro**
- bromine we get **bromo**
- iodine we get **iodo**.

Chlorine and bromine react with ethene to form 1,2-dichloroethane and 1,2–dibromoethane respectively as shown below. Both the products are colourless. Ethene instantly decolourises both bromine and chlorine. The decolourization of bromine is a test for unsaturated hydrocarbons.

Ethene + chlorine $\xrightarrow{CCl_4}$ 1,2 – dichloroethane

$C_2H_4(g)$ + $Cl_2(g)$ $\xrightarrow{CCl_4}$ $C_2H_4Cl_2(g)$

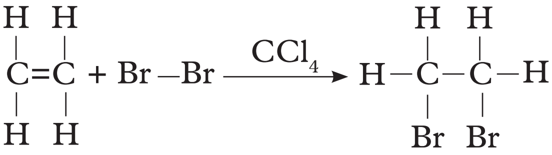
(green-yellow) *(colourless)*



$$\underset{H}{\overset{H}{C}}=\underset{H}{\overset{H}{C}} + Cl-Cl \xrightarrow{CCl_4} H-\underset{Cl}{\overset{H}{C}}-\underset{Cl}{\overset{H}{C}}-H$$

$$\text{Ethene} + \text{bromine} \xrightarrow{CCl_4} 1,2-\text{dibromoethane}$$

$$C_2H_4(g) + \underset{(brown)}{Br_2(l)} \xrightarrow{CCl_4} \underset{(colourless)}{C_2H_4Br_2(l)}$$



$$\underset{H}{\overset{H}{C}}=\underset{H}{\overset{H}{C}} + Br-Br \xrightarrow{CCl_4} H-\underset{Br}{\overset{H}{C}}-\underset{Br}{\overset{H}{C}}-H$$

Work to do: Follow the example above and write the reaction of ethene with iodine.

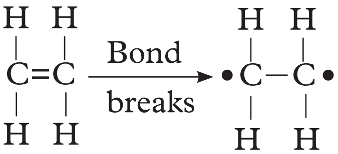
3. Polymerisation reactions (self-addition reactions)

Research activity

Research and present on polymerisation reactions under the following subheadings.

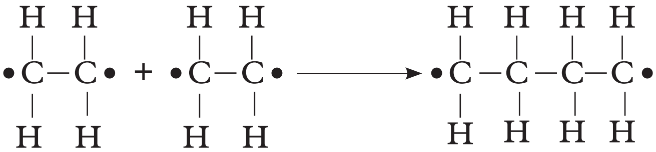
- Definition of monomers, polymers and polymerisation.
- Self-addition reactions and additional polymerisation

When ethene is heated under high pressure and high temperatures, about 200°C, it leads to the formation of low density polymers (polythene). First, one of the double bond in each molecule breaks, forming a monomer with free electrons which form bonds with next neighbours



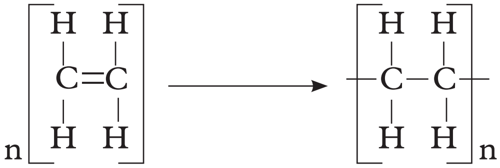
$$\underset{H}{\overset{H}{C}}=\underset{H}{\overset{H}{C}} \xrightarrow[\text{breaks}]{\text{Bond}} \bullet\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}\bullet$$

A monomer in this case is an individual unit of ethene. Then, the monomers add up in a self-reaction.



$$\bullet\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}\bullet + \bullet\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}\bullet \longrightarrow \bullet\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}\bullet$$

Many monomers can join together to form a long chain consisting of thousands of carbon atoms. The process by which many small molecules called **monomers** join to form a large molecule is called **polymerisation**. The large molecule formed is called a **polymer**. The process can be represented by the general equation shown below.



It can also be represented as:

$$\underbrace{n(CH_2 = CH_2) + (CH_2 = CH_2) + (CH_2 = CH_2)}_{\text{ethene (monomers)}} \longrightarrow \underbrace{[CH_2 - CH_2 - CH_2 - CH_2 - CH_2 - CH_2]n}_{\text{polythene (polymer)}}$$

As shown above the bonds of ethene molecule undergoes addition reaction. The process is therefore called **addition polymerisation**.

Test for unsaturation

Activity 9.3

Apparatus and chemicals

- Test tubes
- Teat pipette
- Acidified potassium manganate(VII) solution
- Bromine water

Procedure

1. Bubble ethene through acidified potassium manganate(VII) solution.
   - What do you observe?
   - Record your observation in your notebook.
2. Bubble ethene through bromine water.
   - What do you observe?
   - What can you conclude from the above observations?
3. Record your observations and conclusions in your notebook.
4. What would happen if we bubbled ethane through bromine water?

Ethene as we have already mentioned, is the first alkene. Ethene has a double bond between its two carbon atoms. Hence it is said to be **unsaturated**.

I have discovered that...

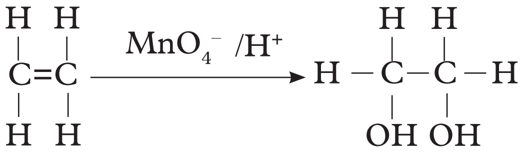
Ethene decolourises potassium manganate (VII) solution as well as bromine water.

The facts

The decolouration of potassium manganate (VII) and bromine water by ethene is the test for presence of a multiple bond i.e double or triple bond, hence the test for unsaturation.

Reaction of ethene with acidified potassium manganate(VII) solution

When alkenes are reacted with acidified potassium manganate(VII) solution, they decolourise the solution. The alkenes are oxidised to compounds with two –OH groups called **diols**.

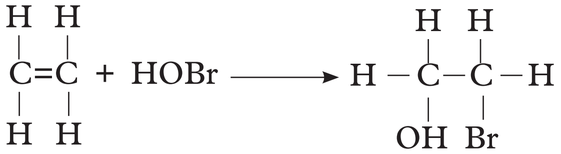


$$\underset{H}{\overset{H}{C}}{=}\underset{H}{\overset{H}{C}} \xrightarrow{MnO_4^-\ /H^+} H-\underset{OH}{\overset{H}{C}}-\underset{OH}{\overset{H}{C}}-H$$

Acidified potassium manganate (VII) is represented by MnO^-_4/H^+

Reaction of ethene with bromine water

When ethene reacts with bromine water, it decolourises bromine water and 2–bromoethanol is formed. Bromine water contains these acids, HBr and HOBr.



$$\underset{H}{\overset{H}{C}}{=}\underset{H}{\overset{H}{C}} + HOBr \longrightarrow H-\underset{OH}{\overset{H}{C}}-\underset{Br}{\overset{H}{C}}-H$$

2 – bromoethanol

Work to do: Write an equation for the reaction of hydrobromic acid with ethene.

9.4 Uses of alkenes

Discussion corner!

1. In your study groups, discuss the uses of alkenes.
2. Write down your findings and share it with other groups.

1. Extremely important in manufacture of plastics.
2. Raw material for manufacture of a wide variety of industrial chemicals such as alcohols.
3. For artificial ripening of fruits.
4. As a general anaesthetic, for making mustard gas(war gas).

Self-evaluation Test 9.2

1. An alkene has a melting point of −140°C and a boiling point of 64°C. Would you expect this alkene to be a solid, liquid or gas at room temperature (25°C)? Explain.
2. Draw and name the structural formulae of all the isomers of pentene.
3. Write an equation to show the complete combustion of ethene in oxygen.
4. Draw a structural formula to show how ethene reacts with chlorine to produce 1,2 - dichloroethane.
5. Write an equation to show how ethene reacts with hydrogen to produce ethane.

9.5 Nomenclature of alcohols and their structures

Activity 9.4

Requirements

- Molecular model kits
- Charts
- Pictures

Procedure

1. In groups model the structure of alcohols using the molecular model kits.
2. Refer to the charts provided for names of the alcohols modelled.
3. Compare your findings with those of other groups.

Study questions

1. What is the general formula of alcohols?
2. Draw the structures of alcohols up to carbon-5.
3. How are alcohols prepared?
4. What are the chemical and physical properties of alcohols?

Discussion corner!

1. In groups, discuss the structure of alcohols and the study questions above.
2. Write a report of your findings and present it in class.

I have discovered that…

- Alcohols like other organic molecules, have a consistent arrangement of atoms in their structures.
- Alcohols have the general formula $C_nH_{2n+1}OH$

The facts

Alcohols are organic compounds. They are also known as alkanols. The alkanols form a homologous series with **OH** as the functional group. They may be regarded as being derived from an alkane by replacing one of the hydrogen atoms with an **-OH** group. They are saturated organic compounds.

Nomenclature of alkanols

The general formula of alkanols is ROH, where R is an alkyl group. Alkyl group is the name given to the group of atoms remaining when one of the hydrogen atom of an alkane is removed. For example

When we remove one hydrogen atom from

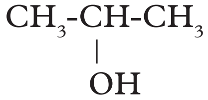
- Methane (CH_4), we get methyl (CH_3–) group.
- Ethane (CH_3 CH_3), we get ethyl (CH_3CH_2–) group.
- Propane ($CH_3CH_2CH_3$), we get propyl ($CH_3CH_2CH_2$–) group etc.

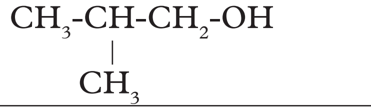
In this case methyl, ethyl and propyl are examples of alkyl groups and can be represented by letter R as described above. To get the number of carbon and hydrogen atoms, represented by R in the general formula, we use another formula C_nH_{2n+1}, where n is the number of carbon atoms which can be 1, 2, 3, 4 etc

The names of members of this homologous series (alkanols) are obtained by replacing the suffix **-e** of the corresponding alkane by **–ol**, for example, when n = 1 in alkanes, the name is methane. In this case if we replace the suffix -e in methane with -ol, the name becomes methanol. It is the first member of the family of alkanols.

We can use the general formula C_nH_{2n+1} OH to find the structural formula of each member of this family as follows. For methanol n = 1, we substitute in the formula to get $C_1H_{(2\times1)+1}$ OH which gives CH_3OH. This is the molecular formula of methanol. Methanol is almost the same as methane in structure. The difference is that we replace one of the hydrogen atoms with hydroxyl group, –OH.

Table 9.3 Names, formulae and structures of some alkanols

Molecular formula	Common name	IUPAC name
CH_3OH	Methyl alcohol	Methanol
$CH_3\text{-}CH_2\text{-}OH$	Ethyl alcohol	Ethanol
$CH_3\text{-}CH_2\text{-}CH_2\text{-}OH$	n-Propyl alcohol	Propan-1-ol
 $CH_3\text{-}\underset{\underset{OH}{\mid}}{CH}\text{-}CH_3$	Isopropyl alcohol or secondary propyl alcohol	Propan-2-ol

CH_3-CH_2-CH_2-CH_2-OH	n-Butyl alcohol	Butan-1-ol
 $\underset{\displaystyle CH_3}{CH_3\text{-}\underset{\vert}{CH}}\text{-}CH_2\text{-}OH$	Isobutyl alcohol	2-Methylpropan-1-ol

Note:

The IUPAC rules demand that the position of the –OH group be included in the name. This is why the, number 1 is included in propan-1-ol, to show that functional group is attached to first carbon of the straight-carbon chain. The same applies to all other alkanols

Characteristics of alkanols as a homologous series

- Members of this series contain,**-OH** as the functional group.
- All the names of alkanols end with **-ol**.
- They have a general formulae C_nH_{2n+1} OH or ROH
- Each member differs from the preceding one by a $–CH_2–$ group (methylene group).
- They show gradual change in physical properties such as melting points and boiling points.
- They have similar chemical properties.

9.6 Preparation of alcohols

In general, alkanols can be prepared in two ways.

- Hydrolysis of alkenes.
- Fermentation of carbohydrates e.g. starch and sugars such as glucose, etc.

1. Preparation of ethanol by hydration of ethene

Ethanol is produced in large quantities, from hydrolysis of ethene. Ethene to be used in this process is produced by cracking long chain alkanes which are broken down into an alkene and shorter alkane. The ethene produced is mixed with steam and passed over a phosphoric acid catalyst at a temperature of about 300□C and 60 atmospheres pressure to produce ethanol.

$$\text{Ethene + steam} \xrightarrow[300^0\,C/\ 60\ atm]{H_3PO_4} \text{ethanol}$$

$$CH_2 = CH_2\,(g) + H_2O\,(g) \longrightarrow CH_3CH_2OH(l)$$

Activity 9.5

2. Preparation of ethanol by fermentation of glucose

Apparatus and chemicals

- Delivery tube
- Cork

- Conical flask
- Boiling tube
- Glucose or any other sugar solution
- Yeast
- Calcium hydroxide solution

Procedure

1. Place sugar solution mixed with yeast in a conical flask.
2. Connect a delivery tube with a tight fitting cork to the conical flask and boiling tube containing calcium hydroxide solution as illustrated in Fig. 9.3.

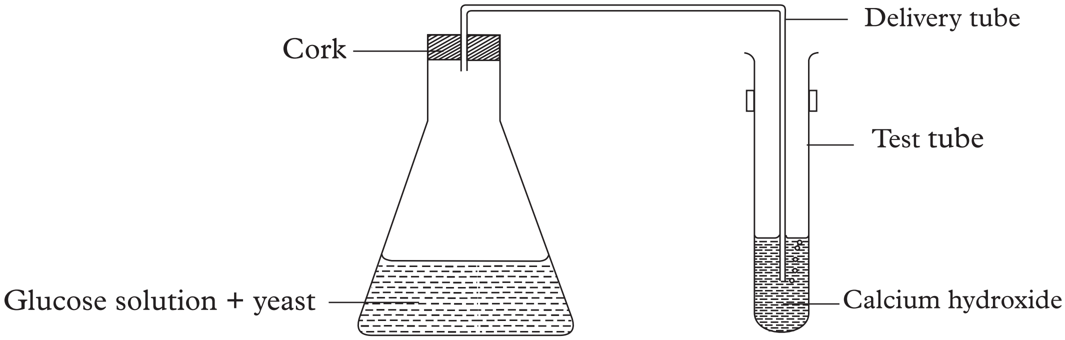


Fig. 9.3: Set-up to demonstrate the fermentation process

3. Keep the apparatus for about 3 to 4 days.
4. Record your observations and conclusions in your notebook.

Study questions

1. What do you observe in the above experiment after 4 days? Explain your observations.
2. What is the role of yeast in this process?
3. What do you conclude from the above observations?
4. What happens to the calcium hydroxide after 4 days? Write equations for the reactions that take place.
5. Calcium hydroxide acts as fermentation lock in the above experiment. Why do we need a fermentation lock?
6. What gas is produced in the above experiment? Explain with equations.

Note: Pure ethanol is obtained by fractional distillation of the mixture formed.

Discussion corner!

1. In groups, discuss the observations in the experiment and the study questions above.
2. Write a report of your findings and present it in class.

I have discovered that…

Ethanol can be prepared in the laboratory through fermentation of glucose. The process requires presence of yeast which acts as a catalyst.

The facts

The reactions whereby sugars are converted to alcohol are referred to as **fermentation** reactions. Ethanol is prepared in the laboratory by the fermentation of glucose or any other sugar solution using enzymes in yeast. Yeast is a single-celled fungus which contains a number of enzymes including **maltase** and **zymase**. The enzymes break down **carbohydrates** e.g. starch into **glucose**. The enzyme zymase, then catalyses the fermentation of glucose into ethanol and carbon dioxide. The following are the reactions that take place.

$$\text{Starch} + \text{water} \xrightarrow{\text{diastase}} \text{maltose}$$

$$2(C_6H_{10}O_5)n\ (s) + H_2O(l) \longrightarrow nC_{12}H_{22}O_{11}(aq)$$

Maltose is then hydrolysed into glucose molecules.

$$\text{Maltose} + \text{water} \xrightarrow{\text{maltase}} \text{glucose}$$

$$C_{12}H_{22}O_{11}(aq) + H_2O(l) \longrightarrow 2C_6H_{12}O_6(aq)$$

Glucose is finally converted to ethanol and carbon dioxide.

$$\text{Glucose} \xrightarrow{\text{zymase}} \text{ethanol} + \text{carbon dioxide}$$

$$C_6H_{12}O_6(aq) \longrightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

The ethanol produced during fermentation is toxic to the enzymes. It **kills** the yeast at concentrations greater than 15% by volume. To increase the alcoholic concentration from fermented liquids, they are distilled to produce spirits. For example, distillation of wine produces brandy which is about 40% ethanol. Brandy is an example of a spirit. Spirits are alcohols with high alcoholic content.

Ethanol is an intoxicating substance in alcoholic beverages. It damages the liver and causes irresponsible behavioural change at high consumption.

Health check

The first member of the alkanol series is methanol and is much more toxic than ethanol. It is added to industrial alcohol to make it undrinkable. Unfortunately, a few people drink it and the final result is blindness, death or both.

Methylated spirit is ethanol where a small amount of methanol has been added to make it undrinkable. A purple dye is often added too, to discourage people from attempting to drink it.

Note: The first fraction from the distillation of solution from fermented sorghum or maize meal which are used for making local brews contains mainly ethanol. Pure ethanol is harmful. Ethanol is a social drink. Abuse of ethanol is a major drug problem in many countries. We should avoid abuse of alcohol.

Self-evaluation Test 9.3

1. (a) Describe the general formula of alcohol.
   (b) State the rules followed when naming alcohols.
2. Give the general formula of alkenes.
3. Give the IUPAC names of the following alcohols.
   (a) CH_3OH
   (b) C_3H_7OH
   (c) $C_5H_{11}OH$
   (d) CH_3CH_2OH
   (e) C_4H_9OH
4. Write the common names of alcohols stated in question 3 above.
5. Using a chemical equation, explain how ethanol is prepared from hydration of ethene.
6. Which gas is produced in the process of fermentation of carbohydrates?
7. Write equations to explain the process of fermentation from starches.

9.7 Properties of alcohols

(a)Physical properties of ethanol

Activity 9.6

a) Solubility

Reagents/ apparatus

- Two boiling tubes
- Dropper
- Test tube rack
- Test tube holder
- Ethanol

Caution: Do not in anyway taste the alcohol.

Procedure

1. Take two boiling tubes and set them in the rack. Label them A and B.
2. Add 2cm^3 of ethanol to one of the boiling tubes(A). Heat the boiling tube. Observe the changes that occur.
3. Add 2cm^3 of ethanol to boiling tube B, add 1cm^3 of water to the same boiling tube B. Observe what happens?

Study questions

1. What happened to boiling tube A when ethanol was heated?

2. Do you think ethanol is soluble in water or not?

b) Boiling point

Reagents/ apparatus

- Source of heat
- Retort stand and clamp
- Boiling tubes
- Ethanol,propanol,butanol
- Thermometer
- Beaker (250cm^3)
- Boiling chips

Procedure

1. Pour 5cm^3 of ethanol into one of the boiling tube and label it A.
2. Fill the other boiling tube with water and label it B.
3. Put the boiling chips into the beaker then fill the beaker with water half-way full.
4. Insert the boiling tube containing ethanol into the beaker of water (used as water bath).
5. Insert the thermometer into the boiling tube containing ethanol.
6. Heat the beaker with water and note when the ethanol starts boiling.
7. Repeat the procedure for propanol and butanol.

Note: Do not allow the ethanol to boil to dryness.

Study questions

1. What is the purpose of adding chips?
2. How did you know it is the boiling point?
3. Why was the heating done in the water bath?
4. What is the boiling point of ethanol, butanol and propanol?
5. Comment on the boiling points of ethanol, butanol and propanol.

c) Action on litmus paper

Reagents/apparatus

- Red and blue litmus paper
- Ethanol
- Test tube

Procedure

1. Put two boiling tubes in a rack. Label them A and B respectively.
2. Add 5cm^3 of ethanol to each boiling tube.
3. Drop a piece of blue litmus paper in tube A and red litmus pare in tube B. What do you observe?

Study questions

1. Why is both blue and red litmus paper used?
2. Comment on the effect of litmus paper on ethanol.

d) Combustion

Reagents/ apparatus

- Ethanol
- Spatula, 10ml measuring cylinder
- Source of heat

Procedure

Take about 2cm^3 of ethanol spatula and ignite it. What do you observe?

I have discovered that…

Ethanol, like other alcohols reacts differently under different conditions. Alcohols also vary in their physical properties.

The facts

Some of the general physical properties of alcohols include the following.

a) **Physical state and odour.** Most common alcohols are colourless liquids. Alcohols containing more than 10 carbon atoms in their molecules are solids. Lower alcohols have a characteristic odour and burning taste.

b) **Solubility.** Lower alcohols such as methyl alcohol, ethyl alcohol are soluble in water, completely miscible. Solubility of alcohols in water decreases with an increase in the number of carbon atoms in the molecule.

c) **Conductivity.** Alcohols do not conduct electricity. This is because alcohols are not ionised.

d) **Action of litmus.** Alcohols have no effect on litmus. Alcohols do not change the colour of litmus. This is because alcohols are neutral compounds.

e) **Boiling points.** The boiling points of alcohols increase with an increase in their molecular masses, thus, an alcohol containing larger number of carbon atoms in its molecule has higher boiling point than alcohol containing lesser number of carbon atoms.

Trends in physical properties of alkanols

Table 9.4: Trends in physical properties of alkanols

Name	Melt point °C	Boiling point °C	Density gcm^{-3}	Soluble in water
Methanol	–98	65	0.791	Soluble
Ethanol	–117	78.5	0.789	Soluble
Propan-1-ol	–127	97	0.803	Soluble
Butan-1-ol	–89.5	117	0.810	Fairly soluble
Pentan-1-ol	–79	138	0.814	Slightly soluble

Self-evaluation Test 9.4

1. State 3 physical properties of ethanol.
2. Explain the trends in melting and boiling points of alkanols.
3. Describe a simple experiment you would carry out to test the properties stated in 1 and 2 above.

Fermentation of sugar or starch

Activity 9.7

Materials / apparatus

Distilled water, yeast, 5 conical flasks, weighing scale, source of heat

Procedure

1. Label the conical flasks as A, B, C, D and E.
2. Weigh out 30 g of sucrose and put it in A, B and C.
3. Weigh out 1 g of yeast and add it into A, B, D and E
4. Add about 10 g of cassava flour in flask E.
5. Boil the contents of B and wait for a day to make observations and conclusion.

Study questions

1. How can you tell that fermentation took place?
2. Why was the contents of flask B boiled?
3. What can you say about each flask after fermentation?

Discussion corner!

Learners use the internet to research about the uses of ethanol and present to rest of the class.

Beer and wine are produced by fermenting glucose with yeast. Yeast contains enzymes that catalyze the breakdown of glucose to ethanol and carbon dioxide. In this experiment, a glucose solution is left to ferment. Students then test for fermentation products.

Yeast has an enzyme called zymase and this catalyses the fermentation process.

Glucose + zymase $\longrightarrow$ Ethanol + carbon dioxide

$C_6H_{12}O_6\,(aq) \longrightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

Chemical properties of alcohols

Alkanols form a homologous series with similar chemical properties. Experiment 9.8 demonstrates some of the chemical properties of ethanol. This information can be applied to the other alkanols.

(b) Chemical properties of ethanol

Activity 9.8

Apparatus and chemicals

- Bottle tops/evaporating basins
- Test tubes
- Beaker
- Universal indicator
- Blue and red litmus papers
- Sodium metal
- Concentrated sulphuric acid
- Ethanoic acid
- Acidified potassium dichromate
- Acidified potassium permanganate

Procedure: As shown in Table 9.5 below

Table 9.5 Investigating chemical properties of ethanol

Procedure	Observation	Conclusion
1. Place a few drops of ethanol into a bottle top and ignite.		
2. Place about 2-3 cm^3 of ethanol in a test tube. Add about 3 drops of universal indicator. Test with blue and red litmus papers		
3. Place about 2-3 cm^3 of ethanol in a test tube. Add a very small piece of sodium metal. Test for any gas produced with a burning splint.		
4. Place about 2-3 cm^3 of ethanol in a test tube. Add an equal amount of ethanoic acid. Add two drops of concentrated sulphuric acid and gently heat. Pour the resulting mixture into about 70 cm^3 of water in a beaker. Smell the product.		
5. Place about 2-3 cm^3 of ethanol in a test tube. Add about 5 cm^3 of acidified potassium permanganate to the test tube. Warm gently.		
6. Repeat procedure 5 but with acidified potassium dichromate.		

Study questions

1. What colour changes occur on the indicators?
2. What is the pH of ethanol?

3. Which gas is produced when sodium reacts with ethanol?
4. Write the equations for reactions in 3, 4 and 5 above.

Discussion corner!

1. In groups, discuss the observations in the experiment and the study questions above.
2. Write a report of your findings and present it in class.

The facts

1. Combustion of ethanol

When ethanol is ignited in plentiful supply of air or oxygen, it burns with almost an invisible blue flame to produce carbon dioxide and water.

Ethanol + oxygen ⟶ carbon dioxide + water

$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$

2. Reaction of ethanol with sodium

When sodium metal reacts with ethanol, it forms a compound called sodium ethoxide. Effervescence occurs and a colourless gas which burns with a 'pop' sound is produced. The gas is hydrogen. The reaction is similar to the reaction of sodium with water but is much slower and less vigorous.

Ethanol + sodium ⟶ sodium ethoxide + hydrogen

$2C_2H_5OH(l) + 2Na(s) \longrightarrow 2C_2H_5ONa(aq) + H_2(g)$

Ethanol also reacts with other metals e.g. potassium to produce the corresponding ethoxide and hydrogen gas.

Note: This reaction is also possible with other alkanols. The general equation of the reaction is:

$2R{-}OH(aq) + 2M(s) \longrightarrow 2ROM(aq) + H_2(g)$

Where, R – alkyl group of the reacting alkanol.
M – metal reacting with the alkanol e.g. potassium or sodium.

3. Reaction of ethanol with concentrated sulphuric acid

Concentrated sulphuric acid reacts with ethanol at 180°C to form ethene and water. This is a dehydration reaction. This is the reverse of the reaction used to produce ethanol from ethene.

$\text{Ethanol} \xrightarrow[\text{heat}]{H^+} \text{ethene} + \text{water}$

$C_2H_5OH(l) \xrightarrow[\text{heat}]{H^+} C_2H_4(g) + H_2O(l)$

4. Ester formation

Alkanols react with alkanoic acids (to be discussed later in this topic) to form compounds called **esters**. For example, ethanol reacts with ethanoic acid in the presence of a few drops of concentrated sulphuric acid (catalyst) to form a substance with characteristic sweet and fruity smell called ethyl ethanoate. This substance is called an ester.

$$\text{Ethanol} + \text{ethanoic acid} \rightleftharpoons \text{ethyl ethanoate} + \text{water}$$

$$CH_3CH_2OH(l) + CH_3COOH(aq) \rightleftharpoons \underset{\text{Ester}}{CH_3COOCH_2CH_3} + H_2O$$

The process of ester formation is called **esterification**. Esterification process is reversible and the ester formed can be reverted back to the parent alkanol and alkanoic acid as shown in the reaction above a process is called **ester hydrolysis**.

Note: The name of the ester is derived from the reactants; in this case, the respective alkanol and alkanoic acid. In our example above; the name comes from ethanol and ethanoic acid respectively.

Oxidation of ethanol

When acidified potassium dichromate (VI) is added to ethanol, its colour changes from orange to green. When acidified potassium permanganate is added to ethanol, the purple colour changes to colourless. Potassium dichromate and potassium permanganate are oxidising agents. They oxidise ethanol to ethanoic acid. The oxidation of ethanol is a two-step process.

1. Ethanol is first converted to ethanal. This is a slow process.

$$\text{Ethanol} + \text{oxygen} \longrightarrow \text{ethanal} + \text{water}$$

$$C_2H_5OH(l) + [O] \longrightarrow CH_3CHO(l) + H_2O(g)$$

2. Then rapid conversion of ethanal to ethanoic acid.

$$\text{Ethanal} + \text{oxygen} \longrightarrow \text{ethanoic acid}$$

$$CH_3CHO(aq) + [O] \longrightarrow CH_3COOH(aq)$$

Note:

- If ethanol is allowed to come into contact with air, it can be oxidised by oxygen in the air to dilute ethanoic acid. This dilute ethanoic acid is commonly referred to as **vinegar**. The souring of ethanol was the original method of making vinegar.
- Oxygen of the air slowly oxidises ethanol in beer. If a bottle of alcohol is not tightly corked and a lot of carbon dioxide escapes, alcohol is oxidised making the beer go 'flat'.

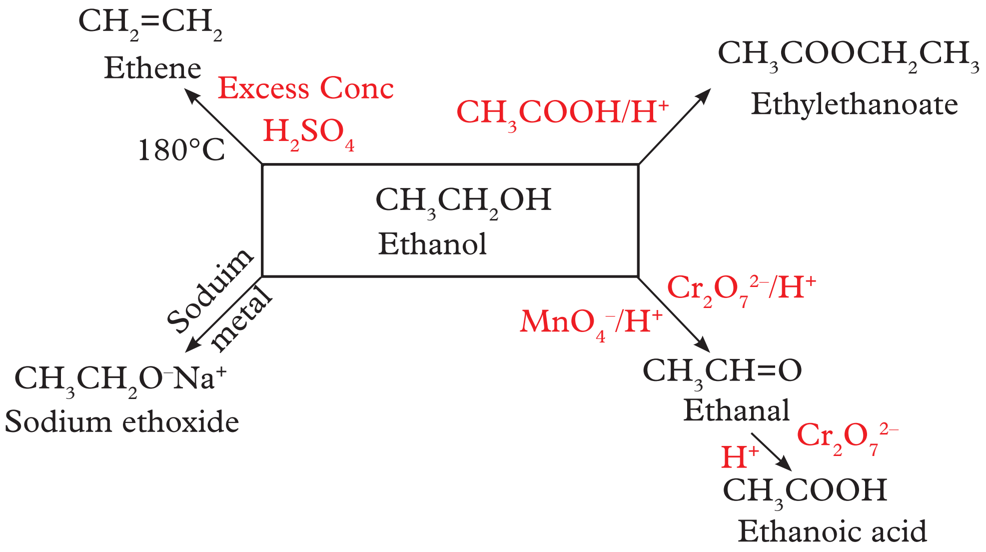


Fig. 9.4: Summary of some reactions of ethanol

9.8 Uses and misuse of alcohol

Uses of ethanol

1. It is used as a solvent to dissolve substances water cannot. It evaporates easily, so it is used a solvent in glues, printing inks, perfumes and aftershave.
2. Fuel in engines. It is usually added to petrol or used instead of it in some automobiles because it burns cleanly, it is less polluting than petrol and reduces the reliance on petrol and diesel.
3. It is used to make other chemicals such as esters which are used in food flavourings and in many cosmetics products.
4. In some cultures, ethanol produced by fermentation is used for making alcoholic drinks.
5. As an anti-freeze in automobile radiators.
6. As a preservative for biological specimens.
7. As an antiseptic to sterilise wounds in hospitals.
8. In the preparation of methylated spirit (mixture of 95% of ethanol and 5% of methanol), rectified spirit (mixture of 95.5% of ethanol and 4.5% of water), power alcohol (mixture of petrol and ethanol) and denatured spirit (ethanol mixed with pyridine).

Misuse of alcohol

Alcohol misuse means drinking excessively beyond what your body can withstand. The risk on one's health is increased by drinking large amounts of alcohol on a regular basis.

Short term health risks are:

- Accidents and bodily injuries
- violent behaviour or victim of violence

- Having unprotected sex that could lead to unplanned pregnancies and sexually transmitted diseases
- Loss of personal possessions
- Alcohol poisoning
- Reckless behaviour

Persistant alcohol misuse increases possible serious health problems.

The effect of alcohol in the human body

Discussion corner!

1. In groups discuss the sources, uses and misuse of alcohol.
2. Share experiences too of your contact with the effects of alcohol.
3. Write a report and present it in class.

Drinking too much alcohol on a single occasion or over time can have serious health effects. Here is how alcohol affects the body.

(a) Brain

Alcohol interferes with the brain's communication pathways. It affects the way the brain functions. These disruptions change the mood and behaviour of an individual making it harder to make sound decisions when one is heavily drunk. It also interferes with the coordination of movements.

(b) Heart

Alcohol can damage the heart in the following ways:

- Stretching and drooping of the heart muscle (Cardiomyopathy).
- Causing Irregular heart beat (Arrhythmias)
- Increases chances of stroke
- High blood pressure occurrence

(c) Liver

Heavy drinking affects the liver and can lead to a variety of liver problems and inflammations including:

- Fatty liver
- Alcoholic hepatitis
- Fibrosis
- Cirrhosis

(d) Pancreas

Alcohol causes the pancreas to produce toxic substances that can eventually lead to pancreatitis, a dangerous inflammation and swelling of the blood vessels in the pancrease that prevents proper digestion.

(e) Cancer

Drinking too much can increase the risk of developing certain cancers, including cancers of:

- Mouth
- Oesophagus
- Throat
- Liver
- Breast

(f) Immune system

Drinking too much alcohol can weaken the immune system making one vulnerable to a number of other diseases. Chronic drunkards are more liable to contract diseases like pneumonia and tuberculosis than people who do not drink too much.

Self-evaluation Test 9.5

1. Copy and complete the passage using the list of words in the brackets provided.

(*Bromine, decolorises, double, ethane, ethene, and orange*.)

Ethane can be distinguished from ethene by adding ________ water to each compound. Bromine water has an ________ colour. Ethene ________ bromine water but ________ does not. The bromine adds across the bond in ________ and no other compound is formed.

2. Match each word on the left with its description on the right

Ethanol	The breaking down of long-chained alkanes to alkenes and short –chained alkanes.
Saturated	Organic compounds containing only single bonds.
Unsaturated	One of the products of fermentation of glucose.
Addition	Organic compounds containing C=C double bond.
Cracking	A reaction in which two or more compounds combine to form only one compound.

3. Ethene is unsaturated hydrocarbon whereas ethane is saturated hydrocarbon.
   - (a) Describe how you can distinguish between ethene and ethane using a chemical test.
   - (b) Draw a full structural formula for ethane and ethene.
   - (c) Ethene reacts with steam to form ethanol.
     - (i) What conditions are needed to carry out this reaction?
     - (ii) Describe another method for making ethanol.
   - (d) Write the equation for the complete combustion of ethanol.

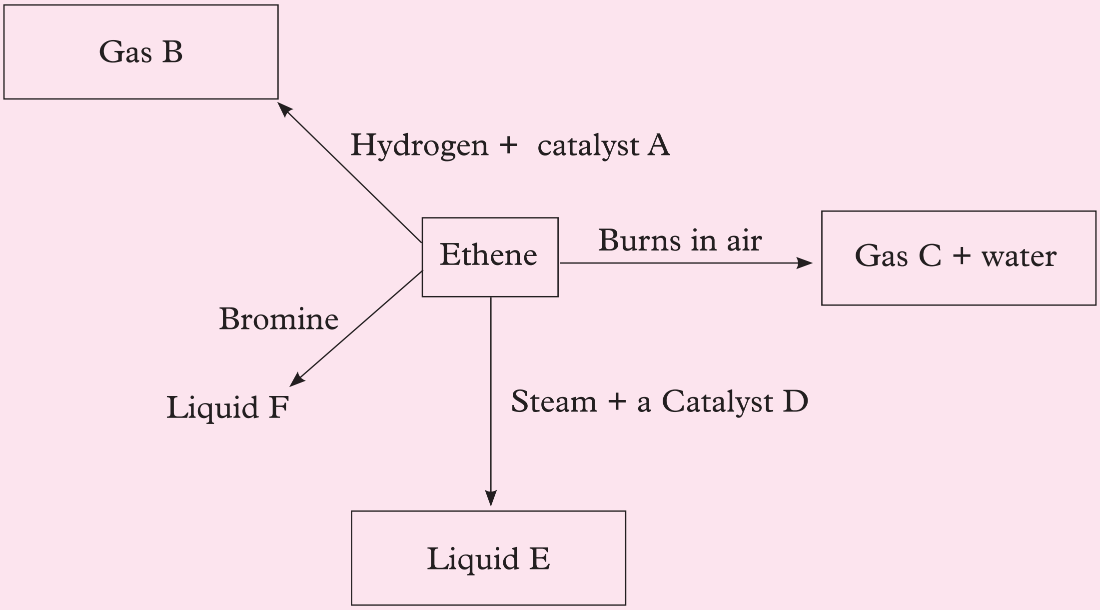
Unit summary 9

- Alkenes are hydrocarbons that contain a carbon-carbon double bond. The number of hydrogen atoms in an alkene is double the number of carbon atoms.
- The double bond in alkenes is the functional group.
- Alkenes show chain and position isomerism.
- Alkenes are prepared by dehydration of alcohols and cracking of hydrocarbons.
- Ethene is a colourless gas, has a pleasant smell, completely insoluble in water and has a low melting and boiling point.
- Alkene burns in air to form carbon dioxide and water.
- Polymerisation is a chemical reaction in which two or more molecules combine to form larger molecules that contain repeating structural units.
- Alkenes are used in the manufacture of plastics, as well as raw materials in many chemical industries.
- Alcohol also known as alkanols, are organic compounds whose molecule contains one or more hydroxyl groups attached to a carbon atom.
- Examples of alkanols include methanol, ethanol, propanol and butanol.
- Alcohols are prepared in two ways: hydration of alkenes and fermentation of carbohydrates.
- Most alcohols are colourless, their solubility decreases with increases in number of carbon atoms, they do not conduct electricity and have no effect on litmus paper.
- Ethanol burns in plenty of air to produce carbon dioxide and water.
- Alkanols react with alkanoic acids to form compounds known as esters.
- Alcohols are used as solvents in fuel engines, to make chemicals such as esters among others.
- Too much alcohol affects the brain, heart, liver, pancreas and is linked to some types of cancer.

Test your Competence 9

Short answer questions

1. Why is the conversion of ethanol to ethanoic acid an oxidation reaction?
2. Write five adverse effects of alcohol drinking.
3. What happens when ethanol reacts with
   (i) sodium
   (ii) potassium permanganate solution?
4. Which of the following hydrocarbons undergo addition reactions:
   C_2H_6, C_3H_8, C_3H_6, C_2H_2 and CH_4.
5. How can you differentiate saturated and unsaturated hydrocarbon on the basis of burning behaviour?
6. What is observed on adding 5% solution of acidified potassium permanganate solution drop by drop to some warm ethanol taken in test tube. Write the name of the compound formed during chemical reaction.
7. The following question is about some of the reaction of ethene.



   (a) Give the name and the formula for substance A to D.
   (b) Write a word equation to represent the reaction in which liquid E is formed.
   (c) Name the homologous series that gas B belongs to.
   (d) Describe chemical test which would allow you to identify gas C.

Multiple choice questions

8. Which process allows large hydrocarbons to be broken down into smaller hydrocarbons?
   A. Cracking
   B. Polymerisation
   C. Fractional distillation
   D. Saturation
9. Which of these is a formula of an alkane?
   A. C_2H_2
   B. C_2H_4
   C. C_2H_6
   D. CH_4
10. Which statement about hydrocarbons is correct?
   A. Alkanes are unsaturated because they have single bonds.
   B. Alkenes are saturated because they have double bonds.
   C. Alkenes are unsaturated because they have double bonds.
   D. All above.
11. When an alkene is mixed with bromine water
   A. the bromine water changes from colourless to orange
   B. the bromine water stays orange
   C. bromine water changes from orange to colourless
   D. no change is observed.
12. In the formula of an alkene, = represents:
   A. Two atoms that are the same
   B. A shared pair of electrons
   C. A double covalent bond
   D. Equal bonds.
13. Which of the following is an alkane?
   A. Ethanol
   B. Hexane
   C. Propene
   D. Ethyne
14. Looking at the chemical formulae of these compounds, which is an alkane?
   A. C_3H_8
   B. C_2H_4
   C. C_5H_{10}
   D. CH_4

15. What is the name of the alkane which has the formula C_5H_{12}?
   A. Ethane
   B. Octane
   C. Butane
   D. Pentane.

16. Alkanes are saturated hydrocarbons. What does this mean?
   A. There are some C=C double bonds present in the molecule
   B. All C atoms are bonded by single covalent bonds
   C. They are very reactive
   D. They contain water.

17. When hydrocarbons like the alkanes burn in plenty of air, what type of reaction takes place?
   A. Complete combustion
   B. Incomplete combustion
   C. Thermal decomposition
   D. Dissociation reaction.

18. Which is the correct statement about incomplete combustion?
   A. Poisonous carbon monoxide may be produced.
   B. It releases more energy than complete combustion.
   C. It is cleaner than complete combustion.
   D. All the above.

19. What is needed in order for chlorine to react with ethane in a substitution reaction?
   A. High temperature
   B. Catalyst
   C. Ultraviolet light
   D. Hydrogen.

20. How do you test for an alkene?
   A. They turn universal indicator purple.
   B. They turn bromine water colourless.
   C. They turn limewater cloudy.
   D. They do not react.

UNIT 10

Carboxylic acids

Key unit competency

After studying this unit, learners should be able to explain the properties of carboxylic acids.

Learning objectives

By the end of this unit, I should be able to:

- Define a carboxylic acid.
- Name and write the structures of some carboxylic acids up to C-5.
- Explain the preparation of carboxylic acids from oxidation of alcohols.
- State the physical properties of some carboxylic acids.
- Explain the chemical reactions of carboxylic acids.
- Explain the reduction of carboxylic acids as a method of preparing alcohols on small scale.
- Carry out experiments to show how carboxylic acids react with other substances.
- Develop an entrepreneurship culture and that of working in a team.
- Develop orderliness in presentations of experimental results and respect the procedures of an experiment.
- Develop observation skills, research and presentation skills as well as skills to manipulate equipment.

Mind teaser

Study the diagrams below. It shows some uses of carboxylic acids.

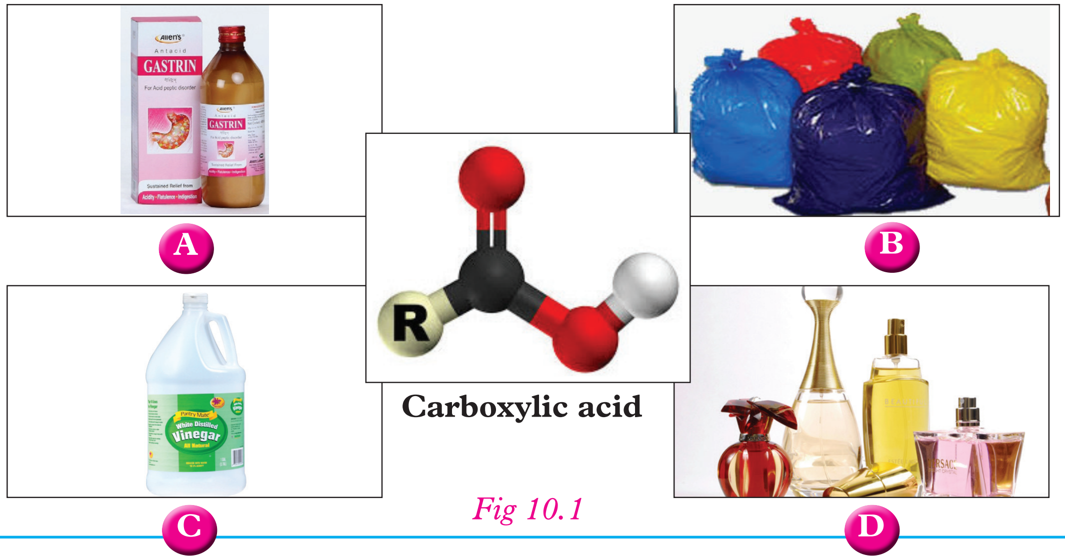


Fig 10.1

Identify each use shown. Do you think these uses are important in our lives? How?

Introduction

Carboxylic acids are organic acids which are sometimes called **alkanoic acids**. They are so called because they have a **carboxyl group** as their functional group. They may be regarded as being derived from alkanes by replacing one of the hydrogen atoms with the–COOH group. Carboxylic acids form homologous series with a general formula R–COOH where R is an alkyl group or H in case of the first one. We use the general formula $C_nH_{2n+1}COOH$ to determine their molecular mass or formula.

10.1 General formula and nomenclature of carboxylic acids

Activity 10.1

Requirements

- Molecular model kits
- Diagrams/charts
- Pictures

Procedure

1. In groups model the structure of carboxylic acids using the molecular model kits.
2. Refer to the charts provided for names of the carboxylic acids modelled.
3. Write the names against each model.

Study questions

1. What is the general formula of carboxylic acids?
2. Draw the structures of carboxylic acids with up to five carbon atoms.

Discussion corner!

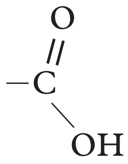
1. In groups discuss the structure of carboxylic acids and the study questions above.
2. Write a report on your findings and present to the rest of of the class.

I have discovered that…

The general formula for carboxylic acids is $C_nH_{2n+1}COOH$.

The facts

The systematic name of a carboxylic acid comes from an alkane with the corresponding number of carbon atoms. The suffix **–e** in the name of alkane is replaced with **"-oic"** acid. The first member of the series is methanoic acid with the formula HCOOH. In methanoic acid, R is represented by H while for the rest of the members, R is represented by an alkyl group. The structure of the carboxyl group is:



The second member of the alkanoic acid family is ethanoic acid, which has two carbon atoms. Since one of the carbon atoms is already attached to the functional group, the value of n in the general formula $C_nH_{2n+1}COOH$ is one. When we subtitute, we get the molecular formula of ethanoic acid as follows.

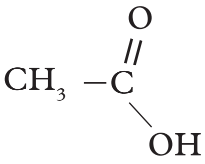
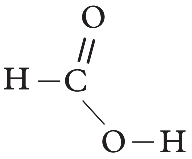
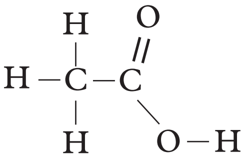
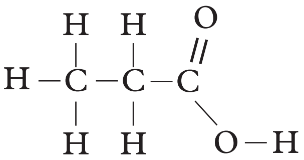
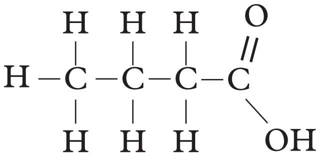
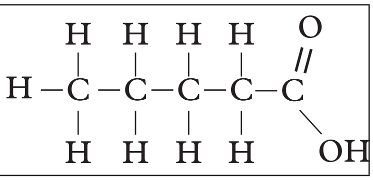


Table 10.1 The first five carboxylic acids

Carboxylic acid	Number of carbon atoms	Molecular formula	Structural formula
Methanoic acid	1	HCOOH	
Ethanoic acid	2	CH_3COOH	
Propanoic acid	3	CH_3CH_2COOH	
Butanoic acid	4	$CH_3CH_2CH_2COOH$	

Pentanoic acid	5	$CH_3CH_2CH_2CH_2COOH$	

Characteristics of carboxylic acids as a homologous series

1. They have the functional group -COOH attached to the hydrocarbon chain.
2. The members have a general formula $C_nH_{2n+1}COOH$ or RCOOH, where n = 0, 1, 2...
3. Members of the series differ from the next by a $-CH_2-$ group.
4. They have similar chemical properties.
5. They show a gradual change in physical properties such as melting points, boiling points and densities.

Self-evaluation Test 10.1

1. What is the functional group of carboxylic acids?
2. Predict the names and structures of carboxylic acids with 6, 7, 8, 9 and 10 carbon atoms.

10.2 Preparation of carboxylic acids

Discussion corner!

1. In groups discuss how carboxylic acids can be prepared. **Hint**: Refer to their structures rather than their corresponding alcohols of the same carbon.
2. Compare your findings with those of other groups.

I have discovered that…

Carboxylic acids can easily be prepared in the laboratory. The method used is the same for all the carboxylic acids.

The facts

Alkanoic acids can be prepared by oxidation of the corresponding alcohol in two steps. The oxidising agents are:

- acidified potassium dichromate
- acidified potassium manganate (VII)

The corresponding alcohol is mixed with excess acidified potassium dichromate (VI) and warmed.

The first step forms an alkanal and water as shown in the following general equation.

(i) Alkanol + oxygen $\longrightarrow$ alkanal + water

$$RCH_2OH(l) + [O] \longrightarrow RCHO(aq) + H_2O(l)$$

R represents an alkyl group.

(ii) Alkanal + oxygen $\longrightarrow$ Alkanoic acid + water

The second step involves further oxidation of alkanal to alkanoic acid i.e.

$$RCHO(aq) + 2[O] \longrightarrow RCOOH(aq) + H_2O(l)$$

Therefore, the overall reaction is:

Alkanol + oxygen $\longrightarrow$ Alkanoic acid + water

$$RCH_2OH(l) + 2[O] \longrightarrow RCOOH(aq) + H_2O(l)$$

Preparation of ethanoic acid

Ethanoic acid is prepared by warming a mixture of ethanol and excess acidified potassium dichromate(VI) or excess acidified potassium manganate(VII).

Step I

This involves the conversion of ethanol to ethanal as shown below. This is a slow process.

Ethanol + oxygen $\longrightarrow$ ethanal + water

$$CH_3CH_2OH(l) + [O] \longrightarrow CH_3CHO(aq) + H_2O(l)$$

Step II

In this step, the ethanal is oxidised to ethanoic acid. This reaction takes place at a much faster rate.

Ethanal + oxygen $\longrightarrow$ ethanoic acid

$$CH_3CHO(l) + [O] \longrightarrow CH_3COOH(aq)$$

Eventually, the orange colour of potassium dichromate (VI) turns green as the oxidation progresses. If we use potassium manganate(VII), the purple colour turns colourless. The acid is obtained from the mixture by distillation at its boiling point while continuously testing the distillate. Distillation is stopped when the distillate is no longer acidic.

Self-evaluation Test 10.2

1. Name the chemical reagents required in the preparation of butanoic acid.
2. Write down an outline of the procedure required in the above preparation.

10.3 Properties of carboxylic acids

Discussion corner!

Study the properties of carboxylic acids in table 10.2. Come up with a summary in trends of the properties of the carboxylic acids.

Table 10.2 Physical properties of some carboxylic acids.

Name	**Melting point (°C)**	**Boiling point (°C)**	**Density gcm^{-3}**	**Solubility in water**
Methanoic acid	8.4	101	1.220	Soluble
Ethanoic acid	16.6	118	1.049	Soluble
Propanoic acid	–20.8	141	0.992	Soluble
Butanoic acid	–8	164	0.964	Soluble
Pentanoic acid	–19	1870	0.939	Slightly soluble

Physical properties of alkanoic acids

1. The first ten straight chain alkanoic acids are liquids at room temperatures.
2. The first few members like methanoic and ethanoic acids are soluble in water. The rest are insoluble and actually float on water. The solubility of carboxylic acids generally decreases with increasing molecular mass.
3. The boiling and melting points of the carboxylic acids increase gradually as the length of the hydrocarbon chain increases except for a few discrepancies in the melting points. The increase is due to the following reasons.

- As the length of the hydrocarbon increases, the weak intermolecular forces (Van der Waals forces) are increased.
- There is also formation of hydrogen bonds between the alkanoic acid molecules. This is brought about by the partial positive and negative charges on the H and O atoms of the carboxyl group(-COOH). These charges enable the molecules to attract each other and form strong hydrogen bonds.

4. Ethanoic acid is a colourless liquid with a sharp smell.

Chemical properties of carboxylic acids

Alkanoic acids are another homologous series of organic compounds that have similar chemical properties. In these series, we will learn about the properties of ethanoic acid as a representative of alkanoic acids.

Apparatus and chemicals

Activity 10.2

- Test tubes
- Droppers
- Beaker
- Universal indicator or litmus papers
- Spatula
- Test tube holder
- Source of heat
- Calcium carbonate
- Magnesium ribbon
- Ethanol
- Ethanoic acid
- Concentrated sulphuric acid
- Sodium hydroxide solution
- Phenolphthalein indicator

Procedure

Copy Table 10.3 in your notebook. Follow the procedures as outlined in the table and record your observations and conclusions in the spaces provided.

Table 10.3 Experiments and results for the chemical properties of ethanoic acid

Procedure	Observation	Conclusion
1. Place about 2 cm^3 of ethanoic acid in a test tube. Add a few drops of universal indicator or test with litmus papers.		
2. Place about 2 cm^3 of ethanoic acid in a test tube. Add a small piece of cleaned magnesium ribbon and test for any gas produced.		
3. Place about 2 cm^3 of ethanoic acid in a test tube. Add a spatula endful of calcium carbonate and test for any gas produced.		
4. Place about 2 cm^3 of ethanoic acid in a test tube. Add an equal amount of ethanol followed by two drops of concentrated sulphuric acid. Gently warm the mixture and pour the resulting mixture into a beaker with about 60 cm^3 of cold water. Smell the product formed.		
5. Place about 2 cm^3 of ethanoic acid in a test tube. Add two drops of either phenolphthalein or methyl orange indicator. Then add 2 cm^3 of sodium hydroxide solution to the mixture drop by drop until there is colour change.		

Discussion corner!

1. In groups discuss the observations made in the experiment and draw proper conclusions.
2. Write a report of your findings and present it in class.

I have discovered that...

Ethanoic acid reacts with carbonates, metals, bases and with ethanol.

The facts

When we dissolve mineral acids in water, they dissociate into their respective ions. For example, HCl produces hydrogen ion H^+, as well as the chloride ion Cl^-

on dissolution in water. Similarly, when an alkanoic acid is dissolved in water, it dissociates to form hydrogen ion, H^+ and carboxylate anion, $RCOO^-$ as illustrated in the following general equation.

$R\text{-}COOH(aq) \rightleftharpoons H^+(aq) + RCOO^-(aq)$
Carboxylic acid   Carboxylate anion

Note: The functional group, –COOH donates the H^+ ion, **not** the alkyl group represented by R.
Ethanoic acid is a weak acid. It partially dissociates into ions in water.

Ethanoic acid $\rightleftharpoons$ hydrogen ion + ethanoate ion

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

An aqueous solution of ethanoic acid has a pH of about 4.7
Aqueous solution of ethanoic acid will show characteristic reactions of a dilute acid as follows:

1. Reaction of ethanoic acid with carbonates

Ethanoic acid will react with carbonates to produce a salt, carbon dioxide and water. For example, it reacts with calcium carbonate to give calcium ethanoate, carbon dioxide and water.
Calcium carbonate + ethanoic acid ⟶ Calcium ethanoate + carbon dioxide + water
$CaCO_3(s) + 2CH_3COOH(aq) \longrightarrow (CH_3COO)_2Ca(aq) + CO_2(g) + H_2O(l)$
Note that for an organic salt formula, the metal is written last. But, when naming the salt, we start with the metal.

2. Reaction of ethanoic acid with metals

Ethanoic acid reacts with reactive metals like magnesium to produce a salt and hydrogen gas. To understand this, let us first start with a reaction that we are familiar with.
$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
Magnesium displaces hydrogen from the hydrochloric acid to form magnesium chloride.
In organic acids, the metal displaces hydrogen from the carboxyl group (–COOH) of the alkanoic acid as shown in the following general equation.
$Mg(s) + 2RCOOH(aq) \longrightarrow (RCOO)_2\,Mg(aq) + H_2(g)$
Now, let us look at the reaction of magnesium with ethanoic acid.
Magnesium + ethanoic acid ⟶ magnesium ethanoate + hydrogen
$Mg(s) + 2CH_3COOH(aq) \longrightarrow (CH_3COO)_2Mg(aq) + H_2(g)$
Note that the metal displaces hydrogen in the –COOH group and **not** the other hydrogens.

3. *Reaction of ethanoic acid with bases*

Ethanoic acid reacts with bases to form salt and water in a reaction known as **saponification**.

Sodium hydroxide + ethanoic acid ⟶ sodium ethanoate + water

$NaOH(aq) + CH_3COOH(aq) \longrightarrow CH_3COONa(aq) + H_2O(l)$

4. *Reaction of ethanoic acid with ethanol*

In general, when organic acids are reacted with alkanols, they form esters in the process of **esterification** as we had seen earlier. Ethanoic acid will react with ethanol in the presence of a few drops of concentrated sulphuric acid to form ethyl ethanoate. Concentrated sulphuric(VI) acid is a catalyst for the reaction.

$$\text{Ethanoic acid} + \text{ethanol} \rightleftharpoons \underset{\text{ester}}{\text{ethyl ethanoate}} + \text{water}$$

$$CH_3CO\boxed{OH}(aq) + CH_3CH_2O\boxed{H} \overset{H^+}{\rightleftharpoons} CH_3COOCH_2CH_3\,(l) + H_2O\,(l)$$

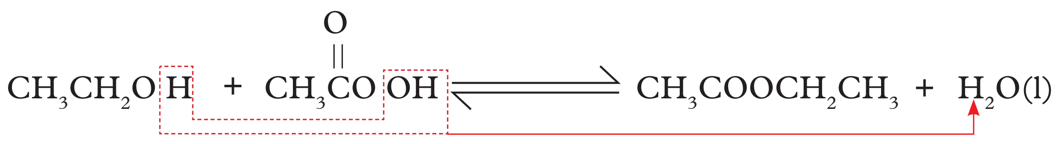
Ethyl ethanoate is one example of an ester.

Note: Always remember that the first part of the ester is derived from the acid while the second part is derived from the alcohol. In writing the name of the esters, we start with alcohol part and end with the acid part.

The process of esterification

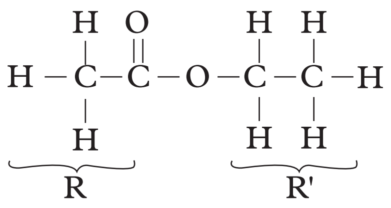
In esterification, the *hydroxyl group* from the acid combines with the hydrogen atom from the hydroxyl group of alcohol to form water and the remaining parts combine.

Ethanol + ethanoic acid ⇌ *ethyl* ethanoate + water

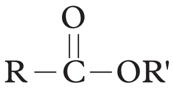


$$CH_3CH_2O\boxed{H} + CH_3\overset{\overset{\displaystyle O}{\|}}{C}O\boxed{OH} \rightleftharpoons CH_3COOCH_2CH_3 + H_2O(l)$$

Esters have the general formula RCOOR′. In the above example, R represents CH_3- from ethanoic acid and R′ represents CH_3CH_2⁻ from ethanol.



The general structure of an ester is represented as



$$R-\overset{\overset{\displaystyle O}{||}}{C}-OR' \quad \text{or} \quad RCOOR' \quad \text{or} \quad RCO_2R'$$

R and R´ represent akyl groups which are not necessarily equal in carbon chain length, but depends on the acid and alcohol used.

Esters form a homologous series of compounds that have strong pleasant smells. Esters are responsible for many of odours in flowers and flavours in fruits. Flowers and fruits have natural esters. For example, the fruity smell of a banana is because of a particular ester; the fruity smell of a pineapple is because of another ester and so on. Esters have also been used in food flavourings and perfumes.

Note: When esters are boiled with acids or alkalis, they break up. During hydrolysis, we get the acid and alcohol used to form that ester. This process is called **ester hydrolysis**.

Example

You are given the following ester to hydrolyse, $CH_3COOCH_2CH_3$.

Determine the acid and the alcohol that were used to form it.

Solution

There are 2 things to remember before you solve this problem.

(a) The **first portion** of the ester is from the acid.

(b) The **last portion** is from the alcohol.

The oxygen atom bridging the two portions is from the alcohol.

Let us start with the last part of the ester.

(a) Cut the molecule as shown below, leaving the alcohol part with an oxygen atom and the acid part also with an oxygen atom.

$CH_3COOCH_2CH_3$

We know that the alcohol ends with –OH.

Add a ‘H’ atom on the alcohol part.

$HOCH_2CH_3$

Write the formula in the usual way to get the alcohol used.

CH_3CH_2OH

Name the alcohol i.e. **ethanol**

(b) The acid must end with OOH.

In the acid part, we already have one ‘O’ atom. Add the remaining **O** and **H** atoms to get the acid.

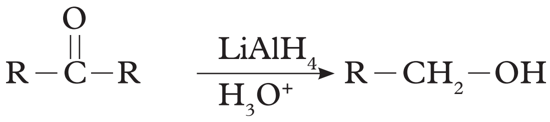
$CH_3CO\mathbf{OH}$

Name the acid i.e. **ethanoic acid**.

Reduction of carboxylic acids with alcohols

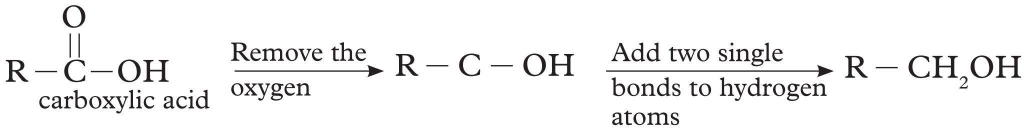
Carboxylic acids can be converted to primary alcohols using strong reducing agents such as lithium aluminium hydride ($LiAlH_4$) or diborane (B_2H_6). An aldehyde is produced as an intermediate product during the reaction. It cannot be isolated because it is more reactive than the original carboxylic acid.

The general equation for the reaction is:



$$R-\overset{\overset{\large O}{\|}}{C}-R \xrightarrow[H_3O^+]{LiAlH_4} R-CH_2-OH$$

i.e.



$$\underset{\text{carboxylic acid}}{R-\overset{\overset{\large O}{\|}}{C}-OH} \xrightarrow[\text{oxygen}]{\text{Remove the}} R-C-OH \xrightarrow[\text{bonds to hydrogen atoms}]{\text{Add two single}} R-CH_2OH$$

Example

$$\underset{\text{Butanoic acid}}{CH_3CH_2CH_2COOH} \xrightarrow[H_3O^+]{LiAlH_4} \underset{\text{Butanol}}{CH_3CH_2CH_2CH_2OH}$$

Activity 10.3

Field trip

Your teacher will organise a field trip to a soap manufacturing factory.

1. Prepare a questionnaire you will use to engage the persons in charge.
2. Write a report of your findings and present it to the teacher.

Self-evaluation Test 10.3

1. Methanoic acid is added to the same volume of ethanol in the presence of concentrated sulphuric acid and warmed.
   (a) What is the name of the ester formed?
   (b) Write an equation for the reaction in (a).
2. Write equations and name the products in the following reactions.
   (a) Propanoic acid and potassium.
   (b) Ethanoic acid and potassium hydroxide.
   (c) Methanoic acid and lead carbonate.
   (d) Butanoic acid and magnesium oxide.
   (e) Butan-1-ol and ethanoic acid heated in the presence of concentrated sulphuric acid.

3. Write the molecular formula and structural formula for butanoic and pentanoic acids.
4. Ethanoic acid is a weak acid. Explain.
5. Explain why the boiling point of ethanoic acid (118°C) is higher than that of propan-1-ol (b.p. 97°C) and yet they have the same molecular mass.
6. Which ester forms from the reaction between methanol and ethanoic acid? Write a word equation for the reaction.
7. Describe with equations, the reaction between ethanoic acid and sodium metal.
8. The following ester, $HCOOCH_3$ was hydrolysed. Write the formula and names of the organic acid and alcohol used to form the ester.

Uses of ethanoic acid

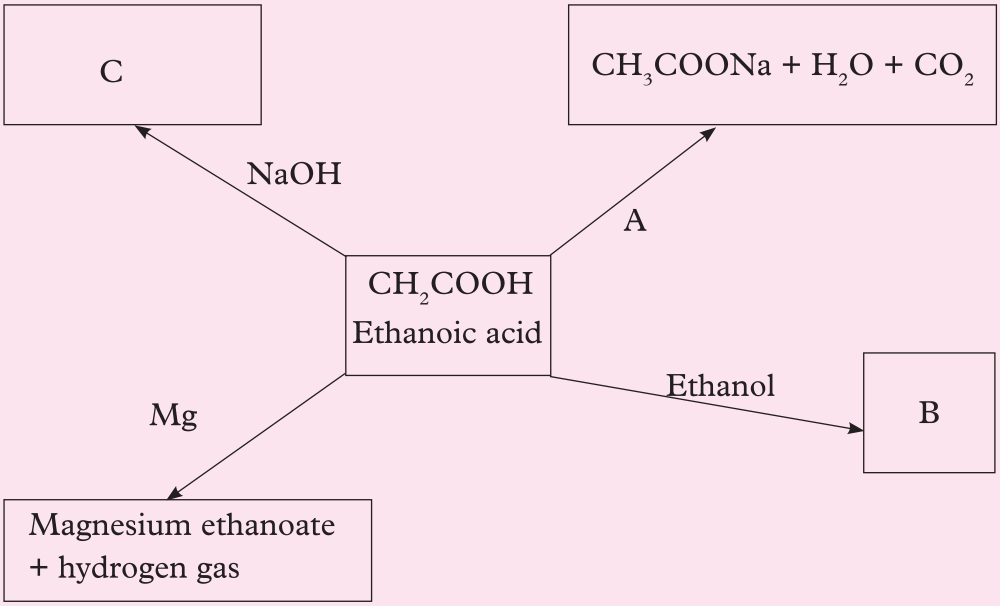
1. For making vinegar which is used as a preservative in food and fruit juices.
2. As a laboratory reagent.
3. For coagulating rubber from latex.
4. In the preparation of dyes, perfumes and medicine.

Unit summary 10

- Carboxylic acids is a homologous series in which the compounds contain a functional group called the carboxyl group (-COOH).
- Examples of carboxylic acids include methanoic acid, ethanoic acid and propanoic acids.
- Carboxylic acids are prepared by oxidation of corresponding alcohol in a two step process using acidified potassium dichromate or acidified potassium manganate (VII).
- Most carboxylic acids are liquids at room temperature, few are soluble in water, their boiling and melting points gradually increases with the length of carbon atoms.
- Ethanoic acid reacts with carbonates to produce a salt, carbon dioxide and water.
- Organic acids react with alkanols to form esters where the hydroxyl group from the acid combines with the hydrogen atom from the hydroxyl group of alcohol to form water.
- Ethanoic acid is used to make vinegar, as a laboratory reagent, for coagulating rubber from latex and for preparation of dyes, perfumes and medicine.

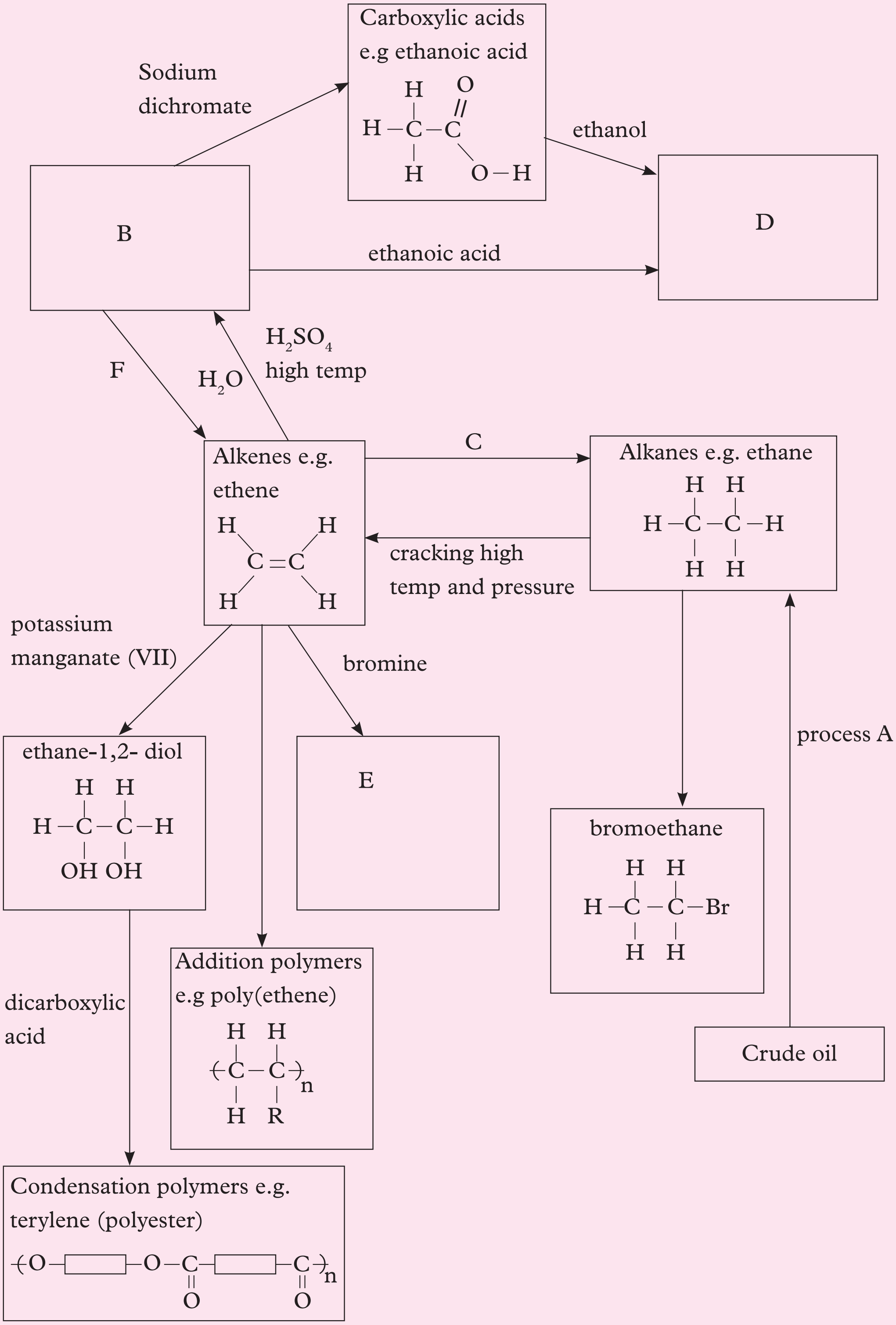
Test your Competence 10

1. Ethanoic acid CH_3COOH, is typical weak acid.
   (a) Draw the full structural formula for ethanoic acid.
   (b) Describe two ways in which ethanoic acid differs from a strong acid such as sulphuric acid.
   (c) Write balanced equations for the reaction of ethanoic acid with:
      (i) Magnesium (ii) sodium carbonate
   (d) Ethanoic acid can be made in the laboratory by the oxidation of ethanol with acidified potassium dichromate.
      (i) Name another oxidising agent that can be used to oxidize ethanol.
      (ii) What conditions are required for this oxidation?
2. Define the following terms.
   (i) Esterification (ii) Saponification
3. Ethanoic acid is a weak acid. Explain.
4. Ethanoic acid, CH_3COOH, forms a weak acid when added to water. Some reactions of ethanoic acid are shown below.



   (a) Explain what is meant by a strong acid.
   (b) Name the substance A that is added to ethanoic acid.
   (c) Substance B is formed when ethanoic acid reacts with ethanol, what type of substance is B?
   (d Draw the structural formula of salt C.
   (e) Write a balanced chemical equation for the reaction between Mg and ethanoic acid.

5. Study the chart below and answer the following questions.
   (a) Give the name of compounds B, C, D, E and F.
   (b) Write the structural formula of B, D and E.



6. When a bottle of wine is left open in the air it turns sour. This is due to the oxidation of ethanol in the wine to ethanoic acid.
   (a) Write an equation to represent this oxidation process.
   (b) What would you expect to happen to the pH of the wine over a number of days?
   (c) Why is it safe for us to use vinegar, which it would be extremely dangerous for us to use dilute nitric acid for the same purpose.
7. (a) Write the structure and names of carboxylic acids containing 2 and 3 carbon atoms?
   (b) What is the functional group of carboxylic acids?
8. (a) What type of reaction is methanoic acid and sodium hydroxide. Explain your answer?
   (b) Write equation for the reaction between methanoic acid and sodium hydroxide.
   (c) What compound is formed when a carboxylic acid reacts with alcohol?
9. Copy and complete the reactions below?

   (a) $HCOOH \xrightarrow[H_3O^+]{LiAlH_4}$

   (b) $CH_3OH + CH_3COOH \longrightarrow$

   (c) $CH_3CH_2OH + CH_3COOH \longrightarrow$

   (d) $CH_3COOH + NaOH \longrightarrow$

   (e) $CH_3COOH + Na_2CO_3 \longrightarrow$

   (f) $CH_3COOH + Na \longrightarrow$

10. Give three uses of ethanoic acid.

UNIT 11

Petroleum products and polymerisation

Key unit competency

After studying this unit, I should be able to explain the origin of petroleum products and application of polymers.

Learning objectives

By the end of this unit, I should be able to:

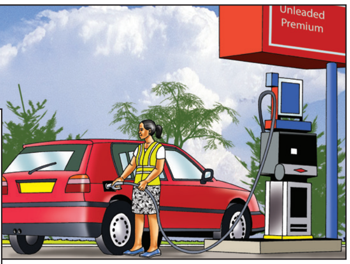
- State the origin of crude oil.
- Describe the process of fractional distillation of crude oil.
- Define polymerization and give examples of synthetic and natural polymers.
- Describe uses of polymers in daily life.
- Classify polymers as natural or synthetic polymers.
- Respect for the procedures of an experiment.
- Appreciate the uses and dangers associated with polymers in daily life.

Mind teaser

Study the pictures below. Can you name the compounds and their uses? What are they made of? What is the relationship between them?

A

C



D

B

E

Fig 11.1

Based on your analysis of the relationship between products above, what do you think this unit entails?

Introduction

There are other carbon-containing compounds apart from alkanes, alkenes and alcohols. Examples include bottled liquid petroleum gas (LPG), kerosene, petrol, diesel among others.

One of the world's most important raw material is **crude oil**. Crude oil is a mixture of many hydrocarbons. These hydrocarbons, in crude oil, have important uses such as fuels and raw materials for many important products such as fuels, plastics and fabrics. In this unit, you will learn about the various petroleum products, their uses and polymerisation which is the process by which plastics are made.

11.1 Origin of crude oil

Activity 11.1

Carry out research from the internet about formation of crude oil. Use the following questions as your guideline.

Research questions

1. Describe how crude oil is formed?
2. How is crude oil produced?

Summarise your findings then share with other class members.

Crude oil is formed from the remains of plants and animals that lived millions of years ago. These remains form sediments at the bottom of seas, and become buried under layers of sedimentary rock. They decay, without air (oxygen), under the action of heat and pressure to form crude oil over millions of years.

It is a fossil fuel because it is formed from once living organisms and the Sun is the original source of energy. It is a non-renewable and finite (limited reserves) energy source because it takes millions of years to form and we use it faster than it is generated. It is also known as a finite energy resource because it will eventually run out.

Coal and natural gas (mainly methane CH_4 are often found with crude oil). They are also non-renewable fossil fuels formed from the remains of plants or animals.

When the fossil fuels are burned the 'carbon', in form of carbon dioxide, is returned to the environment.

11.2 Fractional distillation of crude oil

Activity 11.2

You will be provided with charts, diagrams and computer animations of the fractional distillation process.

1. In your study groups discuss the process of fractional distillation of crude oil.
    - What are the uses of the products formed?
2. Write report on your findings and present it to the rest of the class.

Fractional distillation

Crude oil is a thick black liquid with a strong smell. Unless its components are separated, crude oil does not have much use. During fractional distillation at a refinery, the mixture of hydrocarbons is sorted out into groups or individual hydrocarbons called **fractions**. A fractionating column is used for separating the mixtures.

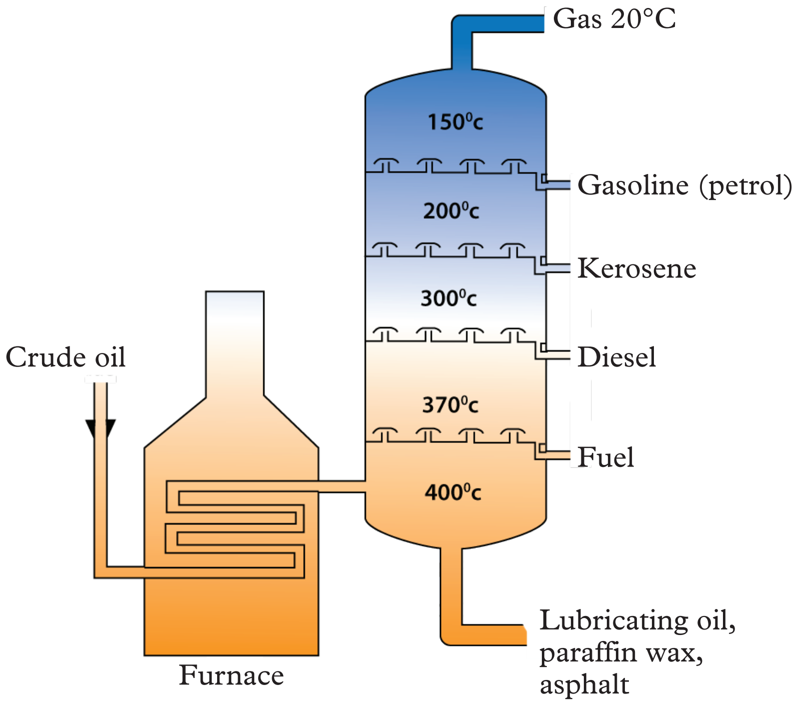


Fig. 11.2 Fractional distillation of crude oil

The crude oil is first heated up in a furnace. As it is heated, the small molecules boil off first. They enter the column as a gas. The fractionating column is hot at the bottom and cooler at the top. The gas molecules then condense. The larger hydrocarbons have higher boiling points. This means that the larger hydrocarbons, with the high boiling points, turn back to liquids easily nearer the bottom. At high temperatures, the hydrocarbons are in form of gases. They rise up the column. The different fractions condense and are collected at different levels as shown in Fig 11.2.

As we have said crude oil is a complex mixture of mainly hydrocarbon compound molecules. What are some of the properties of these compounds that make up crude oil?

Table 11.1 Some properties of components of crude oil

Names of fractions	**c atoms in the molecule**	**Boiling range in °C**	**Uses of the fraction**
Fuel Gas, LPG, Refinery Gas	1 to 4	-160 to 20°C	Methane gas fuel, C_{3-4} easily liquefied, portable energy source bottled gas for cooking (butane), higher pressure cylinders (propane).
Gasoline, Petrol	5 to 11	20 to 60°C	Easily vapourised, highly flammable, easily ignited, car fuel.
Naphtha	7 to 13	60 to 180°C	Not good as a fuel, but valuable source of organic molecules to make other things, cracked to make more petrol and alkenes.
Paraffin, Kerosene	10 to 16	120 to 240°C	Less flammable than petrol, domestic heater fuel, jet fuel.
Diesel oil, Gas oil	15 to 25	220 to 250°C	Car and larger vehicle fuel.
Fuel and lubricating Oils and Waxes	20 to 70	250 to 350°C	Not so easily evaporated, not as flammable, safe to store for central heating oil, quite viscous (sticky) and can also be used for lubricating oils, clear waxes and polishes
Bitumen	over 70	over 350°C	Forms a thick, black, tough and resistant adhesive on cooling, used as waterproofing material and to stick rock chips on roofs or road surfaces

The chemical properties of each substance in the mixture does not change. This means crude oil can be separated by physical methods, in this case fractional distillation, because they have different boiling and condensation points.

The most volatile fraction, i.e. with the lowest boiling point, boils or evaporates off first and goes to the top of the column as seen in table 11.1.

The rest separate out according to their boiling point so that the highest boiling fraction, i.e. the less volatile with higher boiling points, tend to condense more easily lower down the column.

The bigger the molecules, the greater the intermolecular forces, so the higher the boiling points. Chemical bonds are not broken in the process, only the intermolecular forces of attraction.

Uses of petroleum products

1. The refinery gas fractions, under pressure, are conveniently pumped to burner systems to be used for cooking and heating. They are easily ignited and are explosive.
2. **Vehicle fuels** are liquid for compact and convenient storage. They must be easily vaporised to mix with air in the engine prior to ignition. The ease of vaporisation does however make them flammable.
3. **Paraffin** and **kerosene** are less flammable and safer, but not as easily ignited. They are not too viscous to pump not very volatile and so not as flammable and dangerous as petrol or diesel and therefore used for **domestic use purpose**.
4. **Lubricating oil** are quite viscous to stick onto surfaces. Smaller molecules might be more runny but they would evaporate away. It is also water repellent and helps reduce corrosion on moving machine parts.
5. **Candle wax** is used to make candle. The heat from the flame is sufficient to vaporize the hydrocarbons to burn them.
6. **Bitumen** is a water repellent solid at room temperature but is readily melted (sometimes too easily in hot weather). Used as base for a road chipping top surface or sometimes directly. It is also used to waterproof roofing felt.

When crude oil has been distilled into useful fractions it is found that the quantities produced do not match the ratio required for commercial purposes. The needs for petrol and diesel for motor vehicles is rising. Larger molecules which do not make good fuels or have other uses are in plenty while fuel oil, naphtha and bitumen in crude oil exceed demand.

Also, alkenes are not found in crude oil but are one of the most valuable types of organic molecule in the chemical industry. They are used to make polymers (plastics) or ethanol (an alcohol).

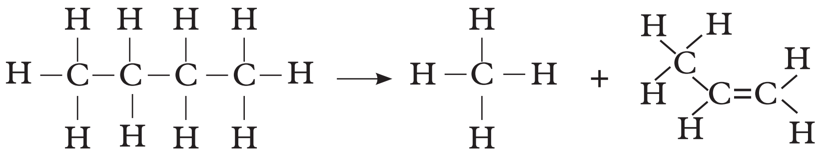
The two deficiencies are remedied by the process of cracking, which converts big molecules into useful smaller ones.

Cracking is done by heating some of the less used fractions at a high temperatures and passing over a suitable hot catalyst. The cracking reaction is an example of thermal decomposition - (a reaction that breaks down molecules into smaller ones using heat). The main products from cracking alkanes from oil are smaller alkanes (for petrol or diesel) and alkenes (for plastics).

The cracking involves breaking single carbon-carbon bonds to form the alkanes (saturated hydrocarbons) and alkenes (unsaturated hydrocarbons) products as shown below.

1. $\underset{\text{Butane}}{CH_3CH_2CH_2CH_3} \longrightarrow \underset{\text{ethane}}{CH_3CH_3} + \underset{\text{ethene}}{CH_2{=}CH_2}$

2.



butane ⟶ methane + propene

3. $C_8H_{18} \longrightarrow C_6H_{14} + C_2H_4$ (making ethene)

4. $C_{12}H_{26} \longrightarrow C_9H_{20} + C_3H_6$ (making propene)

Self-evaluation Test 11.1

1. Why is petroleum thought to have an organic origin?
2. (a) What is the purpose of the fractional distillation of crude oil?
   (b) What is the basic principle that is used in fractional distillation.
   (c) How can the process of fractional distillation be demonstrated in the laboratory. Explain using a diagram.
3. How does the number of carbon atoms in a hydrocarbon affect its boiling point.

11.3 Polymers and polymerisation

Activity 11.3

Field visit and research

Your teacher will organise a field visit to a plastic factory or to the methane gas production plant at Lake Kivu.

1. Prepare questionnaire you will use to engage the officers at the plant site.
2. Interact with the officers during the field visit.
3. Research from the internet about:
   - Definition of polymer.
   - Types of polymerization.
4. Write a report of your finding and present it in class.

Polymers and polymerisation

The word polymer comes from Greek words, **'poly'** which means **many** while **'mers'** meaning **parts**. Therefore polymers simply means many parts of an organic compound.

Polymers which do not stretch readily are suitable for use as fibres. They occur naturally or may be artificially made in industries by joining small units called **monomers**. The process by which they are formed is called **polymerisation**. Polymers and fibres can be classified into two groups.

(a) Natural polymers and fibres.

(b) Synthetic polymers and fibres.

Natural polymers and fibres

All living things contain polymers. Proteins, carbohydrates, wood and natural rubber are all polymers. These polymers occur naturally. Examples of some natural polymers and fibres include:

- Starch
- Cellulose e.g. cotton, wool, paper and silk.
- Proteins
- Glycogen
- Fats
- Rubber

Synthetic polymers and fibres

These are polymers and fibres that are man-made. They include:

- Polythene
- Polychloroethene(PVC)
- Polyphenylethene(Polystyrene)
- Terylene
- Nylon
- Perspex (artificial glass)

Polymerisation

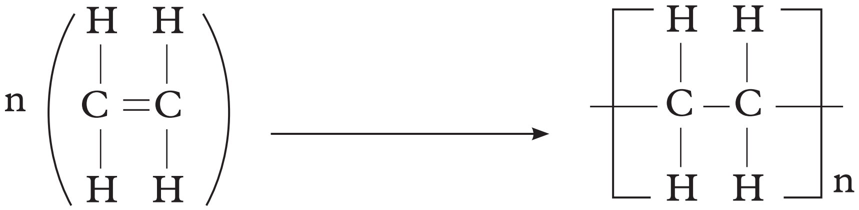
Sythetic polymers are made by polymerisation. There are two kinds of polymerisation:

- Addition polymerisation
- Condensation polymerisation

1. Addition polymerisation of some synthetic polymers

Addition polymerisation involves joining monomers, which are unsaturated molecules like ethene, chloroethene and phenylethene etc. During polymerisation of ethene for example, the double bonds between the carbon atoms are broken down enabling many molecules of ethene to join to each other to form a large molecule called poly (ethene). It is commonly called **polythene**.

(a) **Polythene** is therefore an addition polymer. This can be summarised according to the following equations.

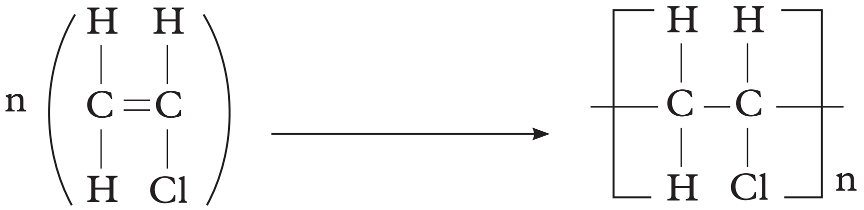


Ethene (monomer) Polythene (polymer)

Note: 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.

(b) **Polychloroethene (Polyvinylchloride) PVC**

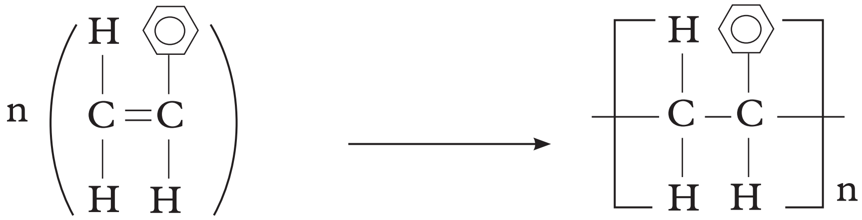
Polychloroethene is made from monomers called chloroethene. It is also formed by addition polymerisation.



Chloroethene or Vinyl

Polychloroethene or Polyvinyl Chloride (PVC)

(c) **Polyphenylethene** is also formed by addition polymerisation



Phenylethene

Polyphenylethene

Table 11.2: Summary of polymers formed by addition polymerisation

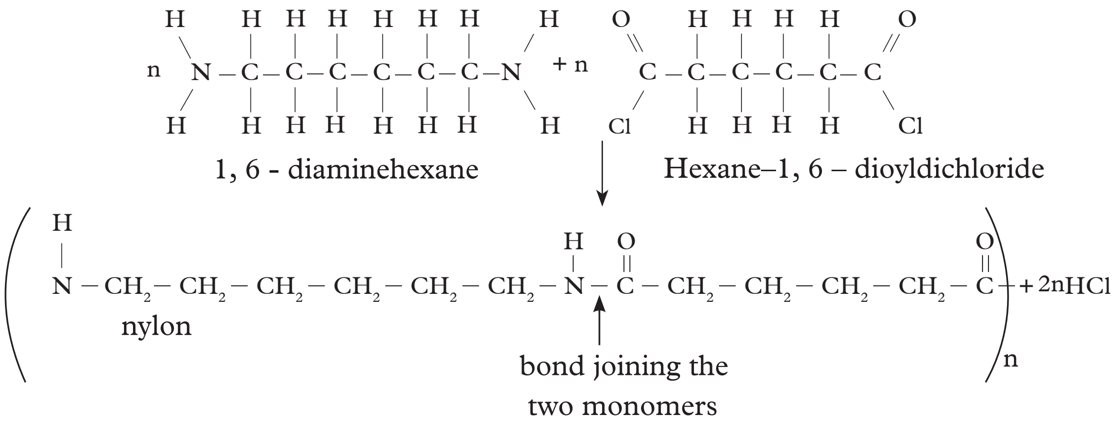
Polymer	**Name of monomer**	**Reaction condition**
Polyethene (high density form)	Ethene (gas)	Low temperature and pressure, special catalyst
Polychloroethene or Polyvinyl Chloride (PVC)	Chloroethene (Vinly Chloride)	60 □C, high pressure, H_2O_2 catalyst
Polyphenylethene (Polystyrene)	Phenylethene (styrene)	Catalyst, heat

2. Condensation polymerisation

This is the formation of a single molecule by the combination of two or more molecules. This process usually is accompanied by elimination of a simple molecule e.g. water or hydrogen chloride. Nylon and terylene are made by condensation polymerisation.

Let us consider condensation polymerisation of nylon. Nylon is a polymer of 1,6-diaminehexane and hexane-1,6-dioyl dichloride.

The structure of 1,6-diaminohexane is



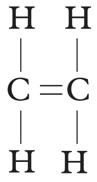
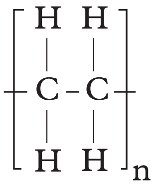
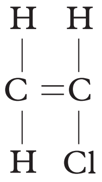
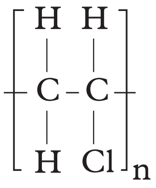
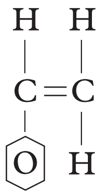
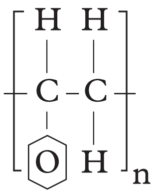
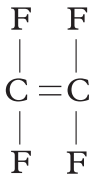
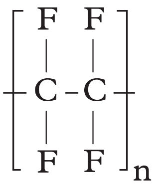
When a molecule of 1, 6-diaminehexane reacts with a molecule of hexane–1,6–dioyl dichloride, the two join by eliminating two hydrogen chloride molecules as shown above.

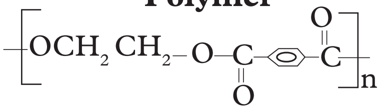
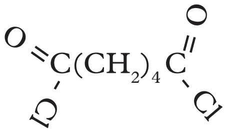
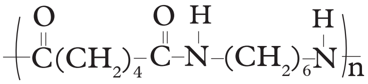
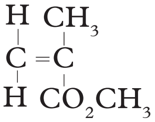
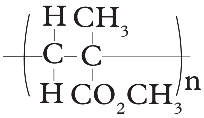
Properties of synthetic polymers

Although different polymers have unique characteristics, many synthetic polymers have the following general properties.

- They do not corrode in air or water.
- They are thermal insulators.
- They are electrical insulators.
- 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 11.3 Properties and uses of specific synthetic polymers

Name of polymer	**Monomer**	**Polymer**	Properties	**Used in making**
Polythene	 $\begin{array}{ccc} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & H \end{array}$	 $\left[\begin{array}{ccccc} & H & & H & \\ & \vert & & \vert & \\ - & C & - & C & - \\ & \vert & & \vert & \\ & H & & H & \end{array}\right]_n$	Tough Durable	• Pipes and plastic bags • Bowls and bottles • Packaging material • Insulation for electric wiring
Polychloroethene or Polyvinyl chloride(PVC)	 $\begin{array}{ccc} H & & H \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & Cl \end{array}$	 $\left[\begin{array}{ccccc} & H & & H & \\ & \vert & & \vert & \\ - & C & - & C & - \\ & \vert & & \vert & \\ & H & & Cl & \end{array}\right]_n$	Tough Durable	• Crates and boxes • Plastic rope • Artificial leather • Water pipes • Insulation for electrical wires
Polyphenylethene			• Light • Poor conductor of heat	• Insulation material • Packaging (foam) • Ceiling tiles • Clothes • Plastic model kits
Polytetrafluoroethene (PTFE) or Teflon	 $\begin{array}{ccc} F & & F \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ F & & F \end{array}$	 $\left[\begin{array}{ccccc} & F & & F & \\ & \vert & & \vert & \\ - & C & - & C & - \\ & \vert & & \vert & \\ & F & & F & \end{array}\right]_n$	• Resistant • Non stick • Strong and withstand high temperature	• Coating non-stick frying pan

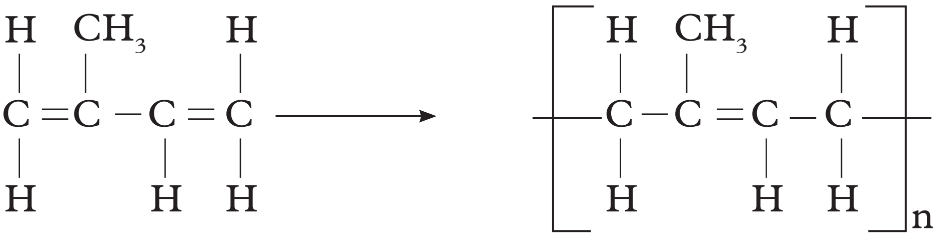
Terylene	**Monomers**  Ethylene glycole Terephthalic acid **Polymer** 	- Light - Tough - Durable	- Clothing e.g. shirts and trousers - Sails and ropes - Plastic model kits
Nylon	**Monomers** $H_2N–(CH_2)_6–NH_2$ 1,6–diamine hexane and  Hexane -1, 6 – dioyldichloride **Polymer**  $-\!\!\left(\overset{O}{\overset{\Vert}{C}}(CH_2)_4\overset{O}{\overset{\Vert}{C}}-\overset{H}{\overset{\vert}{N}}-(CH_2)_6-\overset{H}{\overset{\vert}{N}} \right)_{\!n}$	- Light - Tough - Durable	- Ropes - Clothings - Carpets
Perspex	**Monomer**  $\begin{array}{ccc} H & & CH_3 \\ \vert & & \vert \\ C & = & C \\ \vert & & \vert \\ H & & CO_2CH_3 \end{array}$ methyl -2-methylpropanoate **Polymer**  $-\!\!\left(\begin{array}{ccc} H & & CH_3 \\ \vert & & \vert \\ C & - & C \\ \vert & & \vert \\ H & & CO_2CH_3 \end{array} \right)_{\!n}$	- Strong - Transparent	- Safety glass - Reflectors - Contact lenses - False teeth

Rubber

You probably have seen rubber. An example is the one you use to erase when drawing using a pencil. Rubbers are plastic materials which are elastic. Rubber can be natural or synthetic.

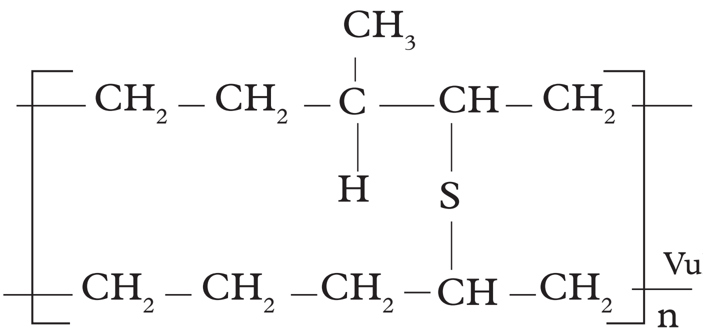
Natural rubber

Natural rubber is a polymer of 2–methylbuta–1,3– diene.



(2-methylbut-1, 3-diene) natural rubber (n =100–5000)

Natural rubber is obtained from rubber trees. The trees produce a milky substance called latex. Latex is a mixture of rubber and water. The latex is collected from incisions in the trunks of rubber trees then taken to the factory for processing to produce rubber. During processing, acid is used to coagulate the rubber. The solid rubber can then be made into different shapes for different uses. Pure rubber is soft and sticky. However, its physical properties can be improved or modified by heating with sulphur. The process by which rubber is heated with sulphur is known as **vulcanisation** of rubber. The sulphur atoms form cross-links between chains of rubber molecules decreasing the number of double bonds in the polymer. Vulcanised rubber is harder, tougher and less flexible. It also has a wide range of applications in industries.

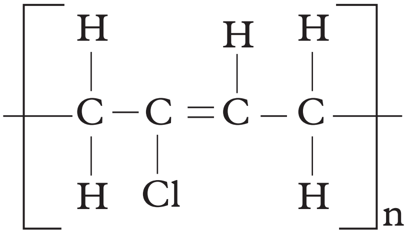


Some uses of vulcanised rubber

Vulcanised rubber is used to make tyres, valves, shoes and for flooring. Research on other uses of vulcanised rubber.

Synthetic Rubber

Because of disadvantages associated with the physical properties of natural rubber i.e. being soft and sticky; a tougher rubber was necessary. Synthetic rubber is made in such a way that it resists action of oils, solvents and abrasion. An example of synthetic rubber is shown below.

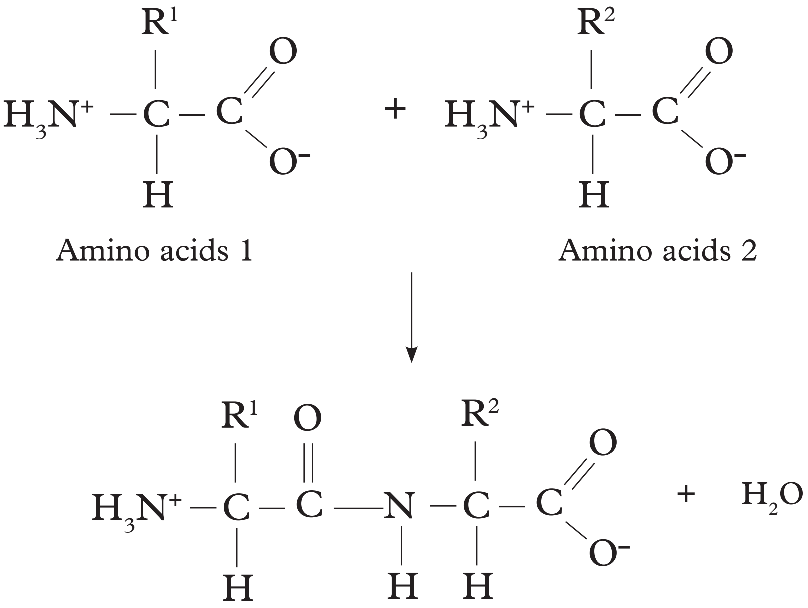


Poly-2-chlorobuta-1,3-diene

(a) Proteins

Proteins are naturally occurring materials that contain the peptide bond. Proteins in the food we ingest are broken down by stomach acids and enzymes which work at body temperature. Once hydrolyzed they form **amino acids** which can then be used by the human body to prepare vital chemicals needed to sustain life.

Proteins contain the same linkages as nylon, but with different units (monomers) called **amino acids**.



Amino acids 1

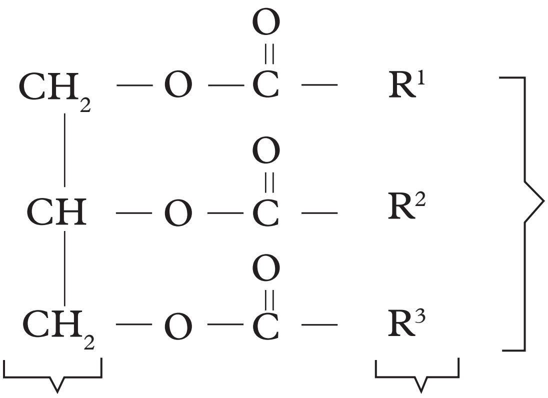
Amino acids 2

dipeptide

Several dipeptides combine through peptide bonds to form a **polypeptide** which is a **protein**.

(b)Lipids (fats or oils)

Fats contain the same linkage as terylene (ester links) but with different units. Each molecule of fat is composed of one molecule of glycerol and three molecules of fatty acids as shown below.



where R is alkyl long chain group

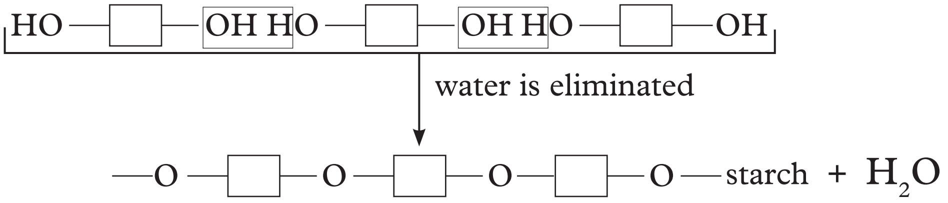
Glycerol

Fatty acids

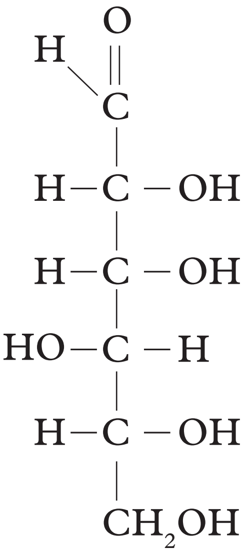
(c) Complex carbohydrates

Examples of complex carbohydrates are starch and cellulose. They have a large number of joined sugar units (monosaccharides like glucose). The sugar units are represented like this:

They join together in a condensation polymerisation reaction:



Starch and cellulose are two very similar polymers. In fact, they are both made from the same monomer i.e glucose, and have the same glucose-based repeat units.



glucose

Self-evaluation Test 11.2

1. Name two methods of preparing polymers and fibres.
2. With equations, describe the preparation of:
   (i) Polyphenylethene
   (ii) Nylon.
3. (a) What is polymerisation ?
   (b) How many products are formed from:
      (i) addition polymerisation?

(ii) condensation polymerisation?

4. The molecular mass of a polymer does not have a fixed value. Explain.
5. Ethene is used in the manufacture of the plastic poly(ethene). Ethene is heated under high pressure in the presence of a catalyst. Many ethene molecules join together to form a giant molecule poly(ethene). The diagram below shows how the polymer is formed.

$$n\ CH_2 = CH_2 \longrightarrow \left[-CH_2 - CH_2- \right]_n$$

   (i) What is the name given to this reaction?
   (ii) Describe how ethene molecules join together to form poly(ethene).

11.4 Dangers associated with polymers

Activity 11.4

1. Carry out a research from the internet about:
   - Problems associated with the use of polymers or plastics.
   - Oil products and environment problems.
2. Discuss your findings in a group.
3. Write a report and present it in class.

Plastic waste like polythene bags, plastic containers and coffee wrappers litter our environment within the cities and on the countryside. Cattle and other domestic animals have been choked after swallowing plastic materials. The problem is made worse by the fact that plastic articles are non-biodegradable i.e. cannot be broken down by micro-organisms like bacteria. They do not decompose like paper, wood among other natural polymers. So, we should avoid careless disposal of or use of these wastes.

How then do we dispose of these wastes?

1. Dumping in rubbish pits.
2. Destroy by burning.
3. Produce biodegradable plastics. Biodegradable plastic bags are already available in some parts of the world.
4. Recycling the plastics.

Note: The safest methods are the ones named in number 3 and 4 above.

• Explain why methods 1 and 2 above are not recommended for disposal of plastics?

Plastics are non-biodegradable and will stay underground for a very long time. Burning plastics can produce toxic gases into the atmosphere and also carbon dioxide which causes global warming. Polymers or plastics cannot be easily broken down by micro-organisms because they are non biodegradable. This leads to waste disposal problems and 'non-rotting' litter around the environment and land-fill sites.

When plastic materials burn they can produce highly toxic gases such as carbon monoxide, hydrogen cyanide and hydrogen chloride (particularly from PVC and other plastics containing chlorine and nitrogen).

- Many countries have become very dependent on the extraction, sale and use of oil based products. The challenges associated with this include pollution and much conflicts or strifes in countries with oil economics. Oil rig accidents, broken pipelines, oil tanker wrecks among other all have terrible effects on the plant and animal life.
- The burning of oil and other fossil fuels is a major contribution to the 'Greenhouse Effect' of global warming. The extra carbon dioxide formed absorbs and traps sunlight (or more precisely the re-radiated sunlight energy from the Earth's surface) like in a greenhouse. The cumulative greenhouse effects and global warming are predicted to be dramatic, for example, rising sea levels as polar ice melts causing flooding in low lying coastal land regions, more energy in the global weather system leads to more frequent violent weather patterns.
- Fossil fuels contain the element sulphur or compounds of sulphur. When the fuel is burned the sulphur also burns to form sulphur dioxide. This is an acidic gas and dissolves in rain water, it then reacts with water and oxygen to form a very dilute solution of sulphuric acid. The acid rain formed destroy buildings and plants.

Self-evaluation Test 11.3

1. State three advantages and disadvantages of:
    (i) synthetic polymers and fibres.
    (ii) natural polymers and fibres.
2. What is a biodegradable and non-biodegradable material?
3. List two examples of each in (2) above.
4. What problems do synthetic polymers and fibres cause in the environment?
5. Suggest four ways to deal with pollution of the environment by synthetic polymers and fibres.

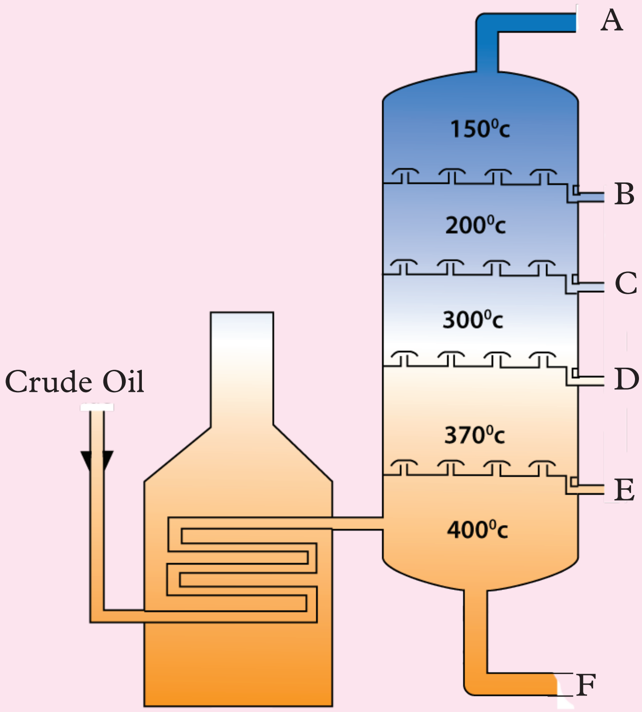
Unit summary 11

- Petroleum products are materials derived from crude oil (petroleum) during processing in oil refineries.
- Crude oil is a naturally occurring, unrefined petroleum product composed of hydrocarbon deposits and other organic materials.
- Fractional distillation is the separation of a mixture into its component parts, or fractions.
- Components of crude oil include fuel gas, petrol, naphtha, paraffin, waxes and bitumen.
- Natural polymers are found in living organisms. They include starch, cellulose, proteins, glycogen, fats and rubber.
- Synthentic polymers and fibres are man-made. They include polythene, PVC, nylon, and terylene.
- Plastic wastes litter the environment and cause pollution.

Test your Competence 11

Short answer questions

1. (a) What is the difference between addition polymers and condensation polymers?
   (b) Give two examples of the following polymers and their uses.
   - Additional polymer
   - Condensational polymer
2. Differentiate between a polymer and monomer.
3. (a) What do you understand by the term **polymerization** and what **conditions** are required for it to take place?
   (b) What is the repeat unit?
   (c) What is the importance of polymers?
4. (a)What are the solutions to the disposal of unwanted polymers?
   (b) Why should we not bury unwanted polymers in the ground?
   (c) Why are we advised not to burn unwanted polymers?
5. Crude oil is a mixture of hydrocarbons. The refining of crude oil produces fractions which are more useful to us than crude oil itself. Each fraction is composed of hydrocarbons which have boiling points within a specific range of temperatures. The separation is carried out in a fractionating column as shown below.



(a) Which separation technique is used to separate the fractions?

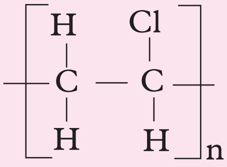
(b) Name the fractions labelled A to F and give one use for each.

(c) Why do the fractions come from the fractionating column in this order?

(d) What is the connection between your answer in C and the size of the molecules in each fraction?

(e) Which of the fractions will be the most flammable?

6. The polymer below is the product of a polymerisation reaction.



$$-\left[\begin{array}{ccc} H & & Cl \\ | & & | \\ C & - & C \\ | & & | \\ H & & H \end{array}\right]_n-$$

(a) Give the structural formula of the monomer in this polymer.

(b) What is the name of the monomer?

(c) Draw the abbreviated structural formula for the polymer.

(d) Has this polymer been formed through an addition or condensation polymerisation reaction?

7. A condensation reaction takes place between methanol and Methanoic acid.

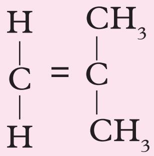
Give the structural formula for:

i. Methanol

ii. Methanoic acid

iii. The product of the reaction

8. The following monomer is a reactant in a polymerisation reaction.



$$\begin{array}{ccc} H & & CH_3 \\ | & & | \\ C & = & C \\ | & & | \\ H & & CH_3 \end{array}$$

(a) What is the IUPAC name of this monomer?

(b) Give the structural formula of the polymer that is formed in this polymerisations reaction.

(c) Is the reaction an addition or condensation reaction?

Multiple choice questions

9. Approximately how many carbon atoms might be linked together in a polymer molecule?

A. Five

B. Fifty

C. Tens thousand

D. Ten

10. What name is given to the small molecules that can be joined together to form a polymer?
   A. Monomer
   B. Monocle
   C. Monoculture
   D. Glucose
11. What monomer would be used to make the polymer poly(propene)?
   A. Propane
   B. Propene
   C. Ethene
   D. Propylene
12. What is a typical use for polythene?
   A. Plastic bags
   B. Insulating wires
   C. Non-stick surfaces on frying pans
   D. Making chairs
13. Which plastic is commonly used for window frames, gutters and drain pipes?
   A. Poly(chloroethene), also called PVC
   B. Poly(phenylethene), also called polystyrene
   C. Poly(propene), also called polypropene
   D. Teflon
14. What does the term biodegradable mean?
   A. It can be made from plants
   B. It can be dissolved into water
   C. It will rot away due to the action of living things
   D. It can be broken into smaller piece
15. Which is the most environmentally friendly way to dispose of plastics?
   A. Landfill
   B. Recycling
   C. Incineration
   D. Burying
16. What feature of alkene molecules allows them to act as monomers?
   A. The double C=C bond
   B. The presence of carbon atoms
   C. The presence of hydrogen atom
   D. They react very fast

17. Polymerisation is the process of
   A. Making large molecules by joining many small molecules together
   B. Making small molecules by breaking down large molecules
   C. Splitting up molecules into their component atoms
   D. Making new elements

18. The building blocks of polymers are
   A. unsaturated
   B. saturated
   C. hydrogenated
   D. catalysed.

19. What happens to the C to C double bond of ethene during polymerisation?
   A. It becomes a triple bond
   B. It remains a double bond
   C. It breaks to become a single bond
   D. It breaks and then reforms

20. Carbon and hydrogen are held together by which type of bonding?
   A. Metallic
   B. Covalent
   C. Hydrogen
   D. Ionic

21. Name the monomer used to produce (poly)styrene.
   A. Propene
   B. Vinyl chloride
   C. Ethene
   D. Styrene

22. Which of the following is not a disadvantage of polymers?
   A. They are not biodegradable
   B. They fill landfills
   C. They produce toxic gases when burnt
   D. Easy to shape and strong for their weight

Appendices

Appendix I

Atomic numbers and relative atomic masses of some 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	29	64
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	201
Neon	Ne	10	20
Nickel	Ni	28	59
Nitrogen	N	7	14
Oxygen	O	8	16

Phosphorus	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^+	Carbonate	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 sulphate	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

Glossary

- **Activation energy** – the minimum amount of energy that the reacting species must possess in order to undergo a specified reaction.
- **Addition reaction** – organic reaction where two or more molecules combine to form a larger one.
- **Addition polymerization** – Polymerization that occurs through the coupling of monomers using their multiple bonds.
- **Alkene** – hydrocarbons which contain one or more double bonds between carbon atoms.
- **Alcohol** – organic compounds which contain OH attached to saturated carbon atom as functional group.
- **Allotropes** – Are different forms of the same element under the same physical state.
- **Allotropy** – Is the existence of an element in more than one form under the same physical state.
- **Amorphous** – of a substance without a clearly defined shape or form.
- **Aqua regia**- is a mixture of nitric acid and hydrochloric acid, optimally in a molar ratio of 1:3.
- **Antacids** – substances that neutralise the acid made by the stomach.
- **Acid–base titration** - the determination of the concentration of an acid or base by exactly neutralising the acid or base with an acid or base of known concentration.
- **Anion** – Negative ion
- **Anode** – Positive electrode connected to positive terminal of d.c. source.
- **Bleach** – to whiten by exposure to sunlight or by a chemical process.
- **Cations** – are positively charged ions.
- **Catalyst** – substance which increase the rate of a chemical reaction but remain unchanged at the end of reaction.
- **Collision** – an instance of one moving particle striking violently against another.
- **Carbon cycle** – The series of processes by which carbon compounds circulate in the environment.
- **Carbonate** – a salt of the anion CO_3^{2-}, typically formed by reaction of carbon dioxide with bases.
- **Cracking** – The process of breaking down complex chemical compounds by heating them.
- **Claus process** – a gas desulfurising process for recovering elemental sulfur from gaseous hydrogen sulfide.

- **Chlorofluorocarbon** (CFC) – an organic compound that contains only carbon, chlorine, and fluorine.
- **Condiment** – a substance such as salt or ketchup that is used to add flavor to food.
- **Concentration** – in industry and consumer world, the most common method of expressing concentration is based on quantity of solute in a fixed quantity of solution.
- **Cathode** – Negative electrode connected to negative terminal of d.c. source.
- **Cation** – positive ion.
- **Condensation polymerization** - Monomers are connected by a reaction in which two molecules are covalently bonded to each other through loss of a water molecule.
- **Carboxyl groups** – weak acids, dissociating partially to release hydrogen ions.
- **Dehydrating agent** – a substance that dries or removes water from a material.
- **Dichlorodiphenyltrichloroethane** (DDT) – was a commonly-used pesticide for insect control. It was banned in many countries due to its side effects.
- **Dioxins** – are a group of chemically-related compounds that are persistent environmental pollutants.
- **Environment** –The immediate surroundings of an organism.
- **Esterification** – reaction between an alcohol and carboxylic acid to form an ester as the only organic compound.
- **Energy** – the strength and power required to sustained chemical reaction.
- **Equivalence point** - the point in a titration where the amount of titrant added is enough to completely neutralize the analyte solution.
- **End point** – the point in a titration at which a reaction is complete, often marked by a colour change.
- **Dilution** – the action of making a liquid more dilute.
- **Standard solution** – a solution whose concentration is known.
- **Percent composition (by mass or by weight)** – parts of solute per 100 parts of solution or the fraction of solute in a solution multiplied by 100.
- **Molarity** – the number of moles of solute in exactly one liter of solution.
- **Mole fraction** – the mole fraction of a component in a solution is the ratio of the number of moles of that component to the total number of moles of all components in the solution.
- **Volume percent** – the ratio of solute in a solution multiplied by 100. Often used when preparing solutions of liquids.
- **Discharge** – the removal of electrons from negative ions to form atoms or the

gain of electrons of positive ions to become atoms.

- **Electrolysis** – decomposition of a compound using electricity.
- **Electrolyte** – an ionic compound which conducts electric current in molten or aqueous solution, being decomposed in the process.
- **Electrode** – a rod or plate where electricity enters or leaves electrolyte during electrolysis. Reactions occur at electrodes.
- **Electroplating** – the process of plating one metal onto another by hydrolysis.
- **Esterification** – the reaction between an alcohol and carboxylic acid to form an ester as the only organic compound.
- **Fractional distillation** – separation of a liquid mixture into fractions differing in boiling point (and hence chemical composition) by means of distillation, typically using a fractionating column.
- **Frasch process** – a method to extract sulfur from underground deposits.
- **Fractional distillation** – the process of separation of mixtures which are miscible to each other basing on their differences on their boiling points.
- **Fossil fuel** – a natural fuel such as coal or gas, formed in the geological past from the remains of living organism.
- **Hard water** – water that has high mineral content.
- **Haber process** – is a specific step by step procedure used in the manufacture of ammonia.
- **Halogens** – a group of five non-metallic elements found in group 17 of the periodic table. The term “halogen” means “salt-former” and compounds containing halogens are called “salts.
- **Hygroscopic** – the property of a substance tending to absorb moisture from the air.
- **Indicator** – any substance that gives a visible sign, usually by a colour change.
- **Metahemoglobin** – a stable oxidized form of hemoglobin that is unable to release oxygen to the tissues due to poisoning.
- **Monomer** – The repeating units that serve as the building blocks of a polymer.
- **Nomenclature** – a system of naming substances
- **Natural resources** – materials or substances such as minerals, forests, water, and fertile land that occur in nature and can be used for economic gain.
- **Neutralisation** – a chemical reaction in which an acid and a base interact with the formation of a salt.
- **Ostwald process** – a chemical process for making nitric acid (HNO_3).
- **Oleum** - a dense, corrosive liquid consisting of concentrated sulfuric acid containing excess sulfur trioxide in solution.

- **Pollution** – the presence in or introduction into the environment of a substance or thing that has harmful or poisonous effects
- **Polyvinyl butyral** (PVB) – a resin mostly used for applications that require strong binding, optical clarity, adhesion to many surfaces, toughness and flexibility.
- **Polyvinyl chloride** (PVC) – the -most widely produced synthetic plastic polymer.
- **Polymer** – A long molecule consisting of many identical or similar building blocks linked by covalent bonds.
- **Polymerisation** – the process by which small molecules called monomers combine to form large molecules called polymers.
- **Product** – a substance that is manufactured during a chemical reaction.
- **Oxidation** – the loss of electrons or an increase in oxidation state by a molecule, atom, or ion.
- **Rate of reaction** – the change of an amount or concentration of a particular reactant or product per unit of time.
- **Reactant** – a substance that takes part in and undergoes change during a reaction.
- **Reduction** – the gain of electrons or a decrease in oxidation state by a molecule, atom, or ion.
- **Refrigerant** – a substance used for cooling (refrigeration).
- **Soft water** – water which has relatively low concentration of calcium carbonate and other ions.
- **Solvay process** – industrial process for the production of sodium carbonate.
- **Saturation** – the state of a substance in which the atoms are linked by single bonds. A fully saturated compound contains no double or triple bonds.

Periodic Table of elements

x Z y Name

X is the atomic Mass (RAM)
Z is the symbol of the element
y is the Atomic number of the element

Groups

Periods	**I** 1	**II** 2											**III** 3	**IV** 4	**V** 5	**VI** 6	**VII** 7	**VIII** 8
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	Co Cobalt	59 Ni 28 Nickel	64 Cu 29 Copper	65 Zn 30 Zinc	70 Ga 31 Gallium	73 Ge 32 Germenium	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	271 Uun 110 Ununniliam	272 Uuu 111 Ununnium	285 Uub 112 Ununbium	Uut 113	289 Uuq 114 unnuqnabium	Uup 115	292 Uuh 116 ununhexium	Uns 117	Uuo 118 ununoctium

Lanthanides	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
Actinides	227 Ac 89 Actium	232 Th 91 Thorium	231 Pa 91 Protactinum	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 Carlifonium	254 Es 99 Einsteinium	253 Fm 100 Fermium	256 Md 101 Mendelevium	254 No 102 Nobelium