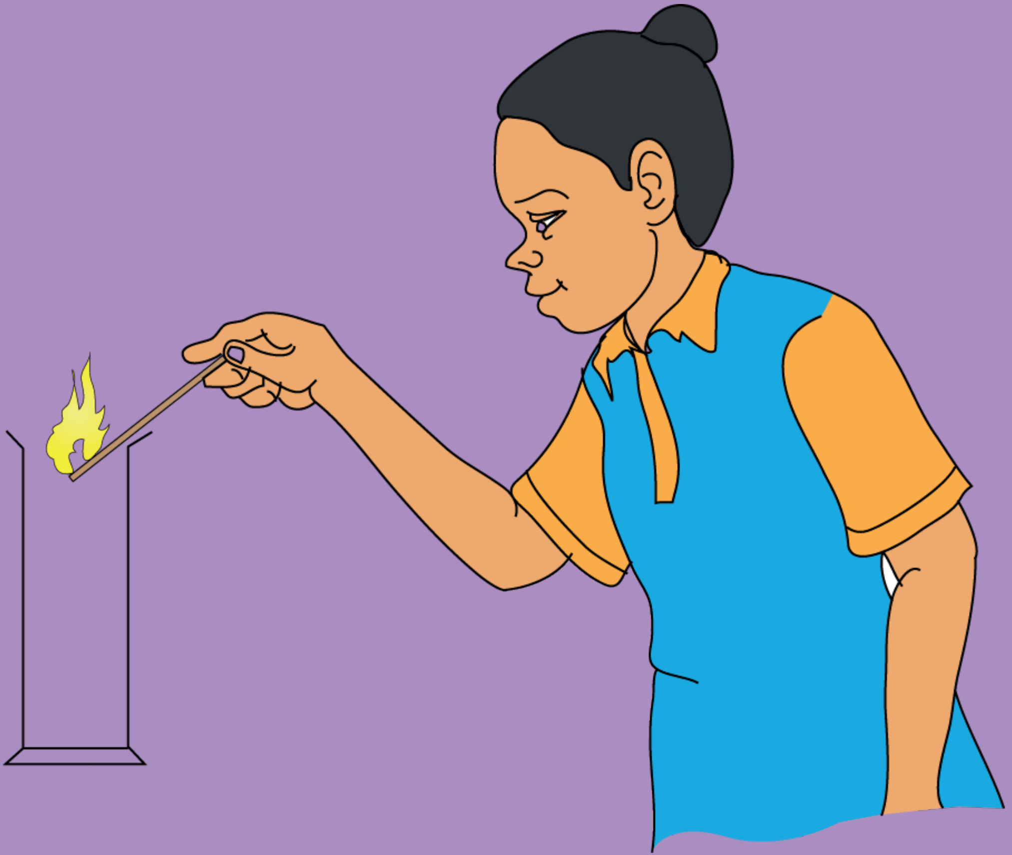
South Sudan

Secondary Chemistry 2

Student's Book



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Chemistry

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Table of Content

UNIT 1

The atmosphere and properties of common gases

Learning Outcome

Knowledge and understanding	Skills	Attitudes
• Describe the composition of atmosphere.	• Carry out practical investigations to determine the percentage of gases by volume. • Investigate the properties of common gases. • Design and perform experiments to investigate the properties of common gases Select, organise and present information from a variety of sources. • Use kinetic molecular theory to explain the properties of common gases.	• Appreciate the importance of the atmosphere.

Introduction

Air is the most common substance known to man. It is all around us. We breathe air to live. We need air to burn fuel and keep warm. We cannot see it, but we can feel it when it moves as wind. Air is a mixture of gases, but the most important component that we need for burning and breathing is oxygen.

1.1 Air is a mixture

Discussion corner

1. Classify the following substances as either compounds, elements or mixtures: Salt water, sugar, tea, iron nails, and tap water, a dozen of books.
2. Is air a compound or a mixture?

The facts

Air is a mixture of gases. The composition of air does not vary from one place to another with the exception of moisture content and dust. These two varies depending on human activities in a given area as well as the weather conditions during the time of analysis.

An air sample from a stone mining area on analysis will have more dust particles than a sample from a region where there are no mining activities.

The table below shows the percentage composition by volume of the components of air.

Table 1.1 Percentage composition of air by volume

Component	Percentage composition
Oxygen	20.9%
Carbon (IV) oxide	0.03%
Nitrogen	78%
Noble gases	1% (0.9%Argon)
Water vapour	Variable
Dust	Variable

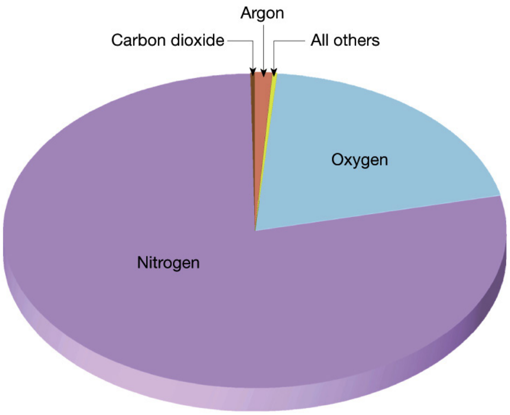


Fig 1.1: Pie chart showing compostion of air by volume

What percentage of air is used up when a candle burns?

Activity 1.1

In group,

Apparatus and reagents

- Water trough
- Gas jar
- Petri dish
- Candle
- Match box
- Ruler
- Sodium hydroxide solution

Procedure

1. Light a candle and allow some melted wax to drop on a petri dish. Place the bottom of the candle on this molten wax. Put off the candle after this.
2. Make the petri dish together with the candle to float on sodium hydroxide solution in the trough.
3. Cover the candle with the gas jar and measure the level of sodium hydroxide solution in the gas jar using a ruler.
4. Uncover the candle and light it once more. Allow it to burn for a few seconds and then cover it with the gas jar for two minutes.
5. Predict what will happen to the level of sodium hydroxide when the candle burns.
6. Measure the level of sodium hydroxide solution in the gas jar at the end of the experiment.
7. Compare your prediction to your results.
8. What can you conclude about your results?

Table 1.2 Measuring the level of sodium hydroxide in a gas jar

	At the beginning	At the end
Level of sodium hydroxide solution in the gas jar (cm).		

Study questions

1. What happens to the burning candle when it was covered with the gas jar for two minutes?
2. What happens to the level of sodium hydroxide solution in the gas jar?
3. What happens to the level of sodium hydroxide solution in the trough?
4. What is the use of sodium hydroxide solution in the experiment?

The facts

Oxygen forms the active part of air that supports combustion of substances. In the experiment below the candle immediately goes off when it was covered with the gas jar. Air initially in the gas jar contained oxygen which supports the burning of candle for a while before it goes off. The residual gases does not support combustion from a different experiment, the following results were obtained.

Carbon + Oxygen ⟶ Carbon (IV) oxide (In candle)

	At the beginning	At the end
Height of sodium hydroxide solution (cm).	8.0 cm	9.4 cm

Using the above data, the active part of air used when candle burns in a fixed volume of air can be worked out as follows.

Initial height of sodium hydroxide solution = 8.0 cm

Final height of sodium hydroxide solution = 9.4 cm

Height occupied by oxygen = (9.4 -8.0) cm = 1.4 cm

Percentage of oxygen used = 1.4/8.0 x 100 % = 17.5 %

Observation made during the experiment

(a) The candle goes off.

(b) The level of sodium hydroxide solution in the gas jar rises.

(c) The level of sodium hydroxide in the trough drops.

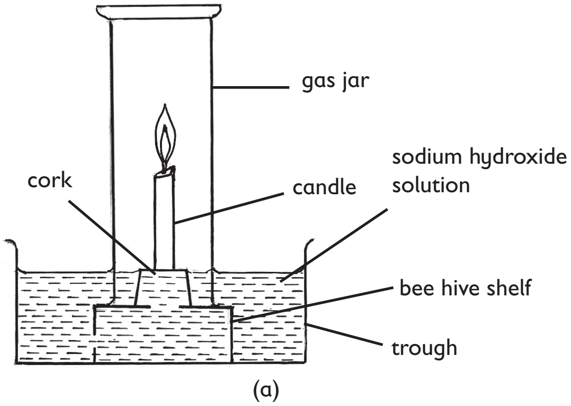


Fig 1.2: Diagram showing the experiment set up

The facts

Candle burns in air to form carbon (IV) oxide and water. Sodium hydroxide solution in the experiment is used to absorb carbon (IV) oxide initially present in the gas jar and the one produced by the burning candle.

Oxygen forms about 20.9% of air by volume. However, experiment results shows a slightly lower value. The error in the experimental value may be due to:

(i) Candle may have gone off before all the oxygen has been used up.

(ii) Accumulation of carbon (IV) oxide in the gas jar making the candle to go off earlier than usual.

(iii) Error in reading the values from the apparatus.

The apparatus are allowed to cool as air expands when heated.

What percentage of air is used when copper is burned in a fixed volume of air?

Activity 1.2

Work as a group

Apparatus

- Copper turnings
- Glass tube
- Glass wool
- Syringes
- Source of heat.

Procedure

1. Place a small portion of copper turnings in a dry glass tube. Put glass wool on both ends of the tube.
2. Label two syringes A and B. Remove air from syringe A and fix it on one end of the glass tube. Pull in air in syringe B to 100 cm^3 mark and fix it on the other end of the glass tube.

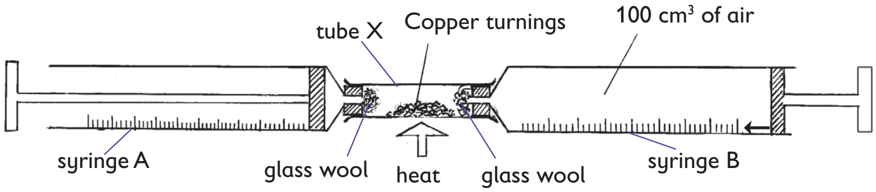


Fig 1.3: Amount of air used by heated copper turnings.

3. Heat the glass tube containing copper turnings strongly.
4. Pass air slowly over heated copper turnings by slowly pushing the plunger of syringe B to and fro several times until there is not further change in volume.

5. Allow the apparatus to cool.
6. Record the volume of the residual air in syringe B at the end of the experiment.
7. Note the color of the solid residue if any. Use the data obtained to work out the percentage by volume of air that was used up during the experiment

Initial volume of air in syringe B = 100cm^3
Final volume of air in syringe B = _________ cm^3
Volume of air used = _________ cm^3
Percentage of air used = _________ cm^3

The facts

Copper burns in air to form copper (II) oxide. During the experiment, air is passed slowly and repeatedly over heated copper turnings to allow more time of contact with copper as well as to ensure that all the oxygen has been used up. The brown copper metal will change colour to black copper (II) oxide at the end of the reaction.

$$\text{Copper} + \text{Oxygen} \longrightarrow \text{Copper (II) oxide}$$

$$\underset{\text{(Brown)}}{2Cu} + O_2 \longrightarrow \underset{\text{(Black)}}{2CuO}$$

The apparatus are allowed to cool since air expands when heated which may lead to erroneous readings. Oxygen forms about 20.9% of air by volume. However, experimental results show a lower value than the theoretical value.

This may be due to;

(i) Error in reading of values in the apparatus.

(ii) Air initially present in the glass tube may not have been accounted for.

(iii) Possible leakage of air from the syringe.

Glass wool prevents copper turnings from being pulled into the syringe when air is pulled in syringe B.

1.2 Determining the percentage of air used when iron rust

Activity 1.3:

In groups

Apparaatus

- Boiling tubes
- Boiled water
- 6 clean iron nails
- Cotton wool
- Oil
- Calcium chloride

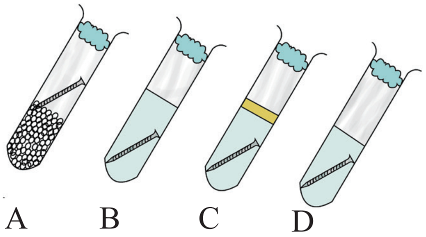


Fig 1.4: Diagram showing experiment set up

Procedure

1. Arrange the apparatus as show along side.
2. Put iron nails in 4 boiling tubes.
3. To one boiling tube add tap water, add boiled water and oil to the second boiling tube, to the third boiling tube add calcium chloride.Put ice cubes the fourth boiling tube.
4. Cover all the boiling tubes with cotton.

Study questions

1. What happened to the iron nails in the first boiling tube?
2. Did the iron nails in the second boiling tube undergo rust?
3. What is the purpose of calcium choride in the third boiling tube.

Activity 1.4

Choose a suitable heading for this activity

In groups,

Apparatus

- Gas jar
- Iron fillings
- Trough
- Water

Procedure

1. Wet the bottom of a gas jar using a little amount of water.
2. Spread some clean iron fillings on the wet surface of the gas jar and make it stick.

3. Invert the gas jar containing the iron fillings into a trough containing water. Allow some water to enter the gas jar to be in the same level with the water in the trough.
4. Measure the level of the air column in the gas jar and note it somewhere.
5. Leave the set up undisturbed for five days.
6. Measure the level of the air column at the end of the experiment.
7. What do you think is being determined in this activity ?

The figure below shows the expected results at the end of the experiment.

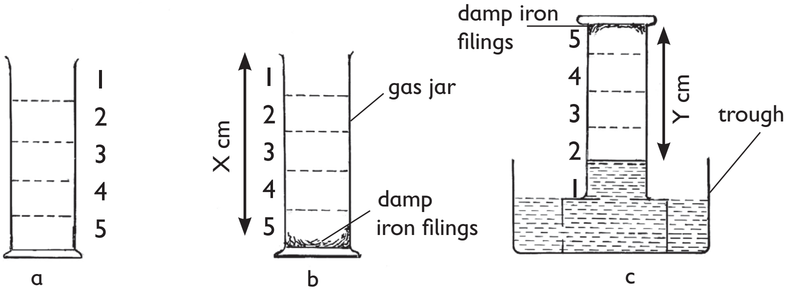


Fig 1.5: To determine percentage volume of air when iron rusts

Initial height of air column (cm) = X cm

Final height of air column (cm) = Y cm

Volume of air used = (X-Y) cm

% volume of air used {(X-Y) /X} x 100%

Worked example

Initial height of air column (cm) = 10 cm

Final height of air column (cm) = 12 cm

Volume of air used (12-10) = 2 cm

% volume of air used = (2/10 x 100) % = 20%

The facts

In the presence of moisture, iron undergoes oxidation to form rust. At the end of the experiment, iron fillings will be covered by a brown substance. During the rusting process, oxygen initially present in the gas jar is used up causing the level of water to rise in the gas jar to occupy the vacuum space left by the oxygen that has been used up. The residual gases present in the gas jar at the end of the experiment are mainly nitrogen, carbon (IV) oxide and the noble gases. Oxygen is therefore a factor necessary for rusting to occur.

What percentage of air is used when phosphorus smoulders in air?

Health check

Phosphorus is highly toxic. Handle it with extreme care.

Activity 1.5: To investigate the percentage of air used when red phosphorus smoulders in a fixed amount of air

In groups

Apparatus and reagents

- Graduated tube
- Red phosphorus
- Water
- Thin copper wire

Procedure

1. Invert a graduated tube into a trough containing water. Note the level of water in the tube.
2. Mount a small piece of red phosphorus on the end of the thin copper wire.
3. Set up your apparatus as shown below.

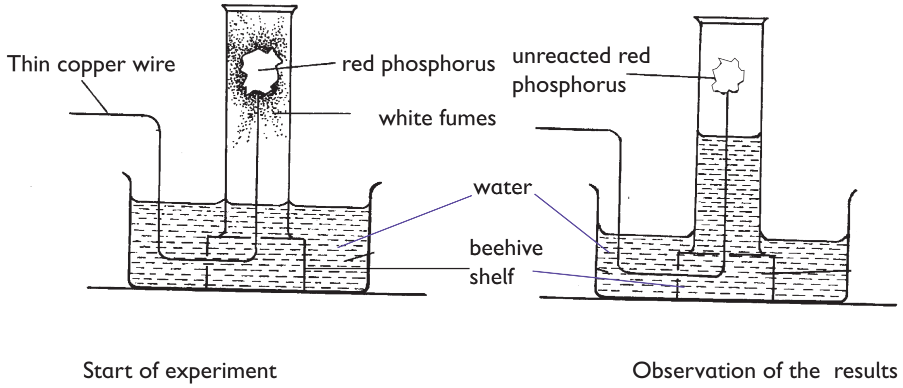


Fig 1.6: To detemine the percentage volume of air when smouldering red phosphorous

Study questions

1. What happened when red phoshours was mounted? Explain.
2. What did you observe at the end of the experiment?
3. Measure the pH of the water at the end the experment.Explain your results.

The facts

Phosphorus smoulders in air to form an acidic oxide phosphorus (V) oxide. Experimental data can be used to work out the percentage volume of air used when phosphorus burns slowly (smoulders) in air.

1. Level of water in the graduated tube rises to occupy the vacuum space left by the oxygen.
2. The level of water in the trough drops.
3. The pH value of the resulting solution drops. Phosphorus (V) oxide dissolves in the water to form a weak acidic solution phosphoric acid.
4. The amount of phosphorus on the thin copper wire reduces.

1.3 Determining the presence of carbon (IV) oxide in air

Preparation of carbon (IV) oxide and its properties will be discussed further in unit 4 of this book.

Activity 1.6: To investigate whether air contains carbon (IV) oxide

In groups

Requirements

- Boiling tube
- Delivery tube
- Freshly prepared lime water

Set up your apparatus as shown below.

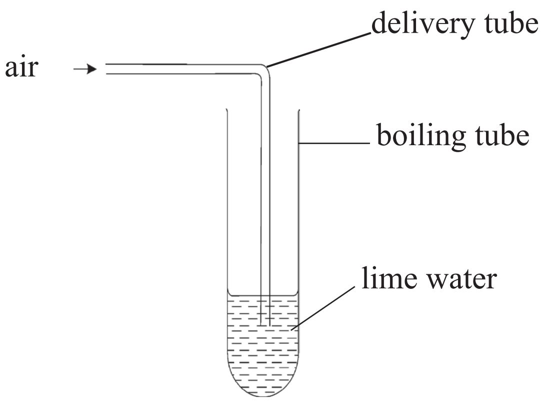


Fig 1.7: Diagram showing air bubbled through lime water

Procedure

1. Bubble air through freshly prepared calcium hydroxide (Lime water) for 30 minutes.
2. Record your observations.

Study questions

1. What was the color of freshly prepared lime water?
2. What was observed when air was bubbled through lime water for 30 minutes?

The facts

When air is bubbled through calcium hydroxide solution for a very long time a white suspension forms on the surface of the solution. This indicates that air contains carbon (IV) oxide.

Carbon (IV) oxide forms a white precipitate with calcium hydroxide solution due to the formation of an insoluble compound calcium carbonate. This is the confirmatory test for carbon (IV) oxide.

1.4: Determining the presence of water vapour in the air

Activity 1.7: Investigating the presence of water vapour in the atmosphere

In groups

Apparatus and reagents

- Anhydrous calcium chloride granules
- Watch glasses
- Anhydrous copper (II) sulphate

Procedure

1. Label two watch glasses A and B.
2. Place a small amount of anhydrous calcium chloride and anhydrous copper (II) sulphate separately on watch glass A and B respectively.
3. Leave the set ups undisturbed and observe after one day.
4. Record your observations.

Study questions

(a) What was the color of anhydrous copper (II) sulphate?

(b) What was the color of copper (II) sulphate after one day?

(c) Describe the nature of calcium chloride after one day?

The facts

Anhydrous calcium chloride is a deliquescent substance. It absorbs water vapour from the atmosphere to form a solution. Copper (II) sulphate absorbs some water vapour from the atmosphere and becomes damp. This explains the change in colour from white to blue.

The two experiments help us to understand that air contains water vapour as one of its components.

1.5 Oxygen

Oxygen occurs naturally as a free molecule in the atmosphere forming about 21% of air by volume. Some processes such as photosynthesis adds oxygen to the atmosphere. Oxygen is an essential element in all living organisms. Without oxygen, there will be no life.

Activity 1.8: Laboratory preparation of oxygen gas

Work as a class

Apparatus and reagents

- Flat bottomed flask
- Clamp and stand
- Hydrogen peroxide
- Trough
- Gas jar
- Water
- Dropping funnel
- Manganese (IV) oxide

Procedure

1. Using a dropping funnel, add hydrogen peroxide carefully into a flat bottomed flask containing manganese (IV) oxide powder.
2. Let the first few bubbles produced escape to the atmosphere then collect several gas jars of oxygen to be used in the following experiments.

Set up your apparatus as shown in the figure below.

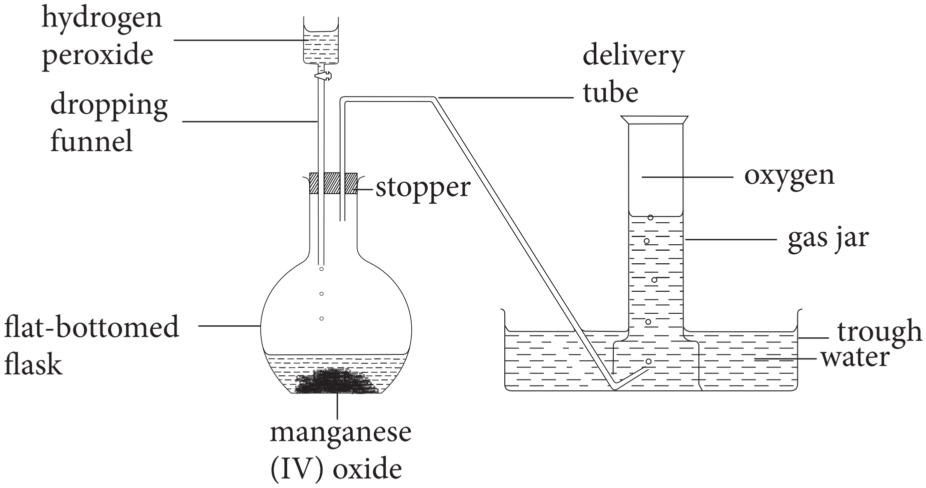


Fig 1.8 : Diagram showing lab preparation of oxygen gas using hydrogen peroxide and MnO_2.

Activity	**Observations**
Lower a glowing splint into a gas jar containing oxygen gas	
Lower a burning piece of magnesium into a gas jar containing oxygen	
Lower a burning piece of Sulphur on a deflagrating spoon into a gas jar containing oxygen	

Study questions

1. Why was it necessary to allow the first few bubbles of oxygen to escape?
2. What was the colour of oxygen gas?
3. Does oxygen gas have any smell?

The facts

Hydrogen peroxide decomposes naturally into oxygen and water. However the rate of decomposition can be speeded up by use of a catalyst such as manganese (IV) oxide. A **catalyst** is a chemical substance that alters or speeds up the rate of a chemical reaction but remains chemically unchanged at the end of the reaction.

Hydrogen peroxide decomposes in the presence of manganese (IV) oxide catalyst as follows.

$$\text{Hydrogen peroxide} \xrightarrow{\text{Manganese (IV) oxide}} \text{oxygen} + \text{water}$$

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

The first few bubbles of oxygen are allowed to escape since it will be a mixture of oxygen and air, hence not a pure sample.

Oxygen gas relights or rekindles a glowing splint. This is the **confirmatory test** for oxygen gas. When a burning piece of oxygen is lowered to a jar containing magnesium, it burns more brightly to form a white ash Magnesium oxide. Burning Sulphur continues to burn in oxygen to form white fumes of Sulphur (IV) with a chocking smell as shown in the equations below.

$$\text{Magnesium} + \text{oxygen} \longrightarrow \text{Magnesium oxide}$$

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

$$\text{Sulphur} + \text{Oxygen} \longrightarrow \text{Sulphur (IV) oxide}$$

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

Health check

Magnesium burns with a blinding white flame which can cause blindness. Do not look directly at the flame of burning magnesium

Oxygen can be prepared in the laboratory using the same set up by the action of water on sodium peroxide.

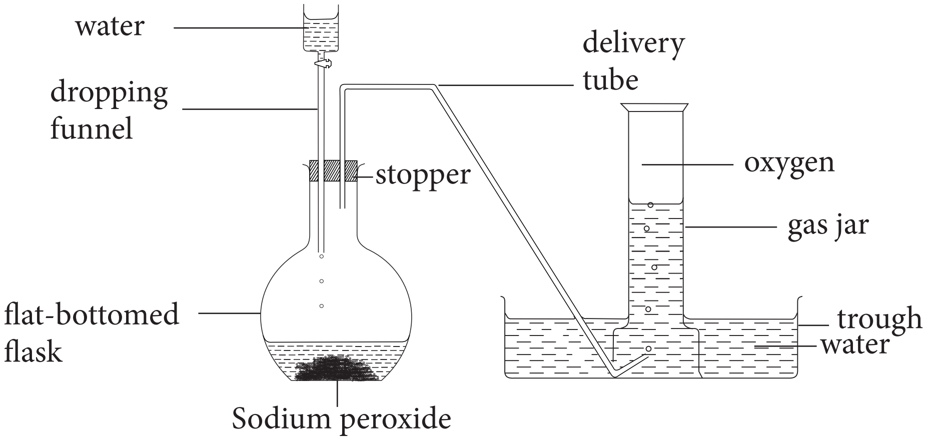


Fig 1.9: Diagram showing laboratory preparation of oxygen gas from sodium peroxide and water

Other ways in which oxygen can be prepared in the laboratory includes:

(i) Heating potassium permanganate.

(ii) Heating potassium dichromate (VI).

When required dry, oxygen is passed through concentrated anhydrous calcium chloride.

Physical properties of oxygen

1. It is as colourless gas.
2. It is oduorless.
3. It is a neutral gas. Has no effect on litmus paper.
4. It is slightly soluble in water. Hence collected by over water method. ($100cm^3$ of water dissolves about $4cm^3$ oxygen).
5. It is 1.1 times denser than air.

Chemicals properties of oxygen

1. Oxygen gas relights a glowing splint. To test a gas produced in a reaction if it is oxygen gas, lower a glowing splint in a containing the gas.If the glowing splint burst into flames, the gas is oxygen. This is the confirmatory test for oxygen gas.

2. Burning of substances in oxygen.

Substances burn in pure oxygen to form their respective **oxides** and **peroxides**.

For example magnesium burns brightly in oxygen with a white flame to form a white solid magnesium oxide.

$$\text{Magnesium + oxygen} \longrightarrow \text{Magnesium oxide}$$

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

Sodium burns in oxygen with a bright yellow flame to form a yellow solid sodium peroxide.

$$\text{Sodium + oxygen} \longrightarrow \text{Sodium peroxide.}$$

$$2Na\ (s) + O_2\ (g) \longrightarrow Na_2O_2\ (s)$$

3. Burning of non metals in oxygen sulphur burn in oxygen vigorously with a blue flame to form white irritating and chocking fumes of sulpur (IV) oxide.

Sulphur + Oxygen ⟶ Sulphur (IV) oxide

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

4. Oxygen has no action on calcium hydroxide.

1.6 Large scale production of oxygen by fractional distillation of liquid air

The temperature of a gas falls as it expands when passing from a region of high pressure to a region of low pressure. Its molecules move further apart to work against their attraction forces. During the process they lose some energy causing the temperature to fall.

Fractional distillation of liquefied air

Air is taken through various stages where some of its components are removed.

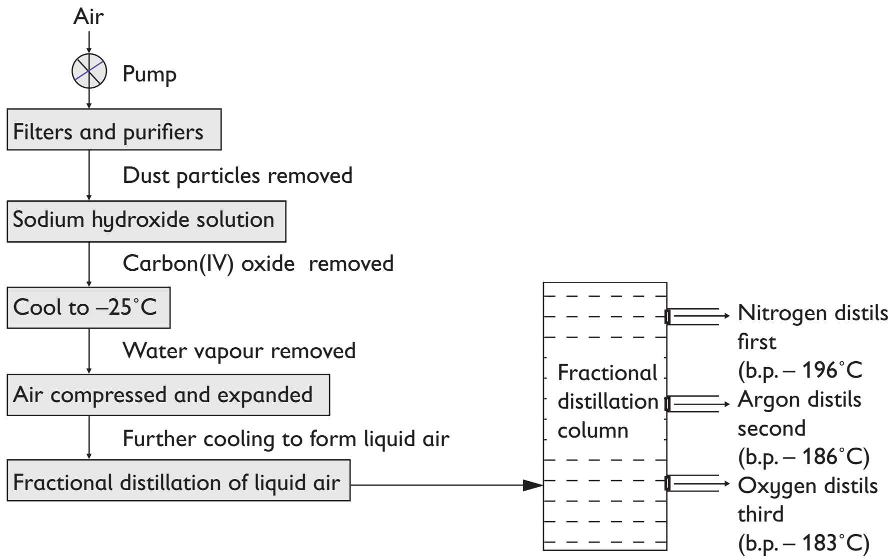


Fig 1.10: Diagram fractional distillation of liquid

The facts

Air is first passed through electrically charged chambers to remove dust particles through electrostatic precipitation. Dust free air is then passed through concentrated sodium hydroxide solution or Potassium hydroxide solution to remove carbon (IV) oxide. Air is then cooled to -25^{0}C where water vapour is condensed and crystalizes out as ice where it is removed. The remaining air that is now free from dust, carbon (IV) oxide and water vapour is then taken through repeated expansion and then compressed to about 200 atmospheres.

At this stage, air will exist as a pale blue liquid. It is then finally taken through fractional distillation where nitrogen distills out in the first fraction at -196^{0}C then argon at -186^{0}C as the second fraction and finally oxygen at -183^{0}C.

Fractional distillation is suitable in separating liquid mixture of different boiling points.

Uses of oxygen

Discussion corner

Study questions

(a) Discuss what is happening in each photograph.

(b) What is the significance of each activity on the photographs?

The facts

1. It is used in hospitals by patients having breathing difficulties or those under anaesthetic drugs.
2. It is used in welding to cut metals. A mixture of oxygen and acetylene produces oxyacetylene flame which is used to cut metals.
3. A mixture of hydrogen flame which is used to melt iron when joining metal pieces together.
4. A mixture of oxygen and helium is used by mountain climbers and deep sea divers.
5. It is used in steel production. A blast of oxygen is usually blown through the molten pig- iron to oxidize the impurities.
6. It is used as rocket fuel. A mixture of oxygen and hydrogen is used as a propellant in rocket engines.

Note: Nitrogen and its properties will be discussed later in the chemistry course

The kinetic molecular theory

Discussion corner

1. How does the smell of a very strong perfume reach every corner of the house?
2. Why do inflated tyres seem deflated after sitting in a cold garage?
3. A few students were on a quest for adventure,so they decided to hiking and camping. On that day they decided to camp at the peak of the mountain, 2000 feet above sea level. They noticed that when they boiled some water, the water boiled faster than usual. Explain why water boils faster in high altitude areas than in low altitude areas?

The facts

Gases are made up of tiny particles that are in continuous random motion. The rapid motion of these particles makes them to collide with one another and the walls of the container.

The kinetic theory postulates that:

i. Gas particles are in constant random motion.

ii. Gas particles have no volume since there is no definite space between them.

iii. Gas particles have no force of attraction holding them.

iv. The pressure in a gas is due to collisions of gas particles with the walls of the container.

v. There is no energy loss or gain when particles collide with one another and the walls of the container.

vi. All gases have the same average kinetic energy at a specified temperature.

vii. The lighter the particles making the gas, the faster they move and vice versa.

viii. An increase in temperature causes a proportional increase in kinetic energy of gas particles which causes a corresponding increase in the speed they move in.

When the temperature of a gas is decreased, the kinetic energy of its particles decreases too. They tend to move relatively slower leading to a fewer collisions. This will cause a proportional decrease in pressure. This explains why inflated tyres appear flat when left to sit on a cold garage.

The activated particles can move randomly to fill every place of the container holding the gas.

When a liquid is heated to boiling, it changes its physical state to gaseous state. For a liquid to boil, the vapour pressure on the surface of the liquid must equal the atmospheric pressure. The atmospheric pressure changes with the altitude.

At low altitudes, atmospheric pressure is high and vice versa. Water boiled at a high altitude will boil faster than water boiled at low altitude area.

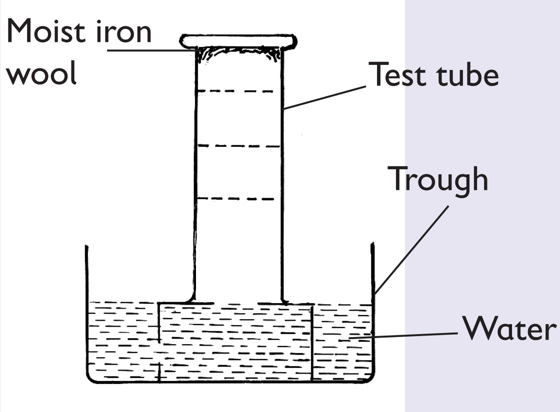
Pressure of a gas results from the collisions between the gas particles and the walls of the container. The more the collisions, the higher the pressure of the gas. When the temperature of a given gas is increased the average kinetic energy of its particles increases which increase their velocity. This particle tends to move faster resulting in more collisions with the walls of the container. As the force from these collisions increases, the pressure increases in the same proportion.

SUMMARY

1. Air is a mixture of gases.
2. Oxygen forms the active part of air.
3. Oxygen does not burn. It only supports combustion of substances.
4. All substances burns in oxygen.
5. Oxygen forms about 20.9% of air by volume.
6. Carbon (IV) oxide forms about 0.03% of air by volume.
7. Nitrogen forms about 78% of air by volume.
8. Oxygen is obtained in large scale from fractional distillation of liquid air.
9. Fractional distillation is used to separate miscible liquid with different boiling points.

Check your progress 1.1

1. The set up below was used to investigate some properties of air. Study it and answer the questions that follow.



   (a) What happens after sometime?
   (b) Draw a diagram to show the expected results of the experiments after two days.
2. In an experiment, 3.0g of copper turnings was completely burnt in an open crucible after cooling the product weighed 3.6g.
   (a) Why was there an increase in mass?
   (b) Give an equation for the reaction that took place.
   (c) State the observations made during the experiment.
3. Copper turnings were heated in a fixed volume of air. The following results were obtained. Volume of air at the beginning = 100 cm^3

Volume of air the end = 79 cm^3

Volume of air used = ________ cm^3

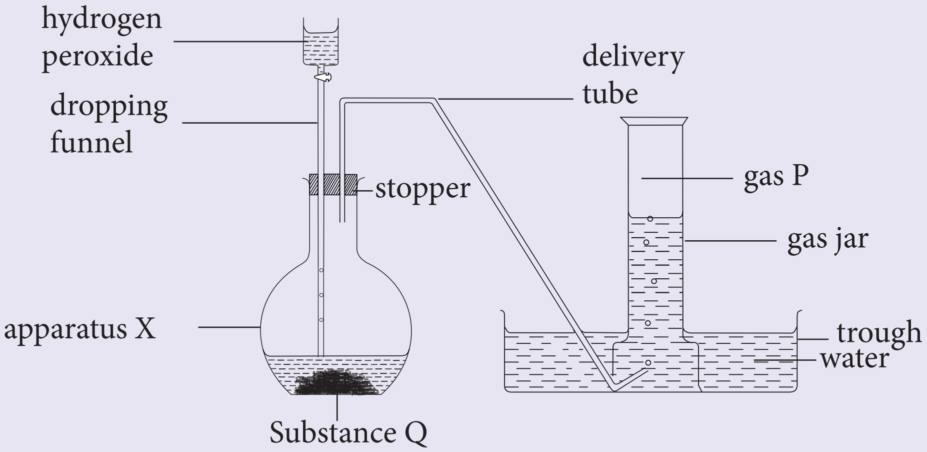
Use the above information to work out the percentage of air that was used during the experiment.

4. Explain why old roofing sheets made of iron appear brown in colour?
5. A certain colourless liquid was suspected to be water. Suggest two chemical tests than can confirm its identity.
6. Complete the following chemical equations

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

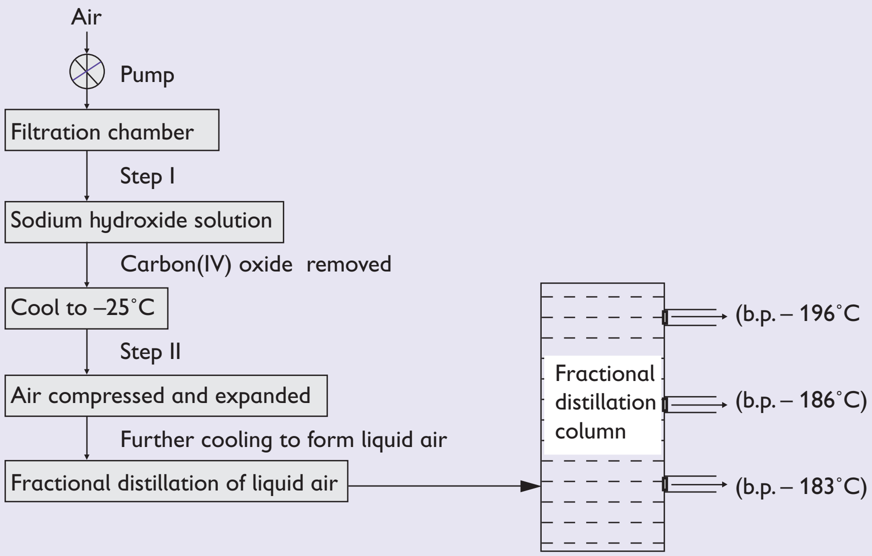
$$Cu\ (s) + O_2\ (g) \longrightarrow$$

7. The set- up below was used by a student to prepare a certain gas in the laboratory.



(a) Identify the mistake in the set up.
(b) Name gas P, apparatus X and substance Q.
(c) What is the nature of the resulting solution at the end of the experiments?
(d) Give one property of gas P that makes it to be collected by the method shown in the set up.
(e) Give two commercial uses of gas F.
(f) Substance Q is used in the preparation of hydrogen peroxide. What role does it play in the reaction?

8. Oxygen is obtained in large scale by fractional distillation of liquid air. The scheme below shows some of the processes involved during the fractional distillation process.Study it and answer the questions that follow.



(a) Identify

(i) Step I

(ii) Step II

(b) Name the substances removed in the filtration and purification chamber.

(c) Identify the component of air collected at -186°C.

9. Moving particles possesses energy in form of ________________

10. State two assumptions of Kinetic theory.

11. State the effect of raising the temperature of a gas in a light bulb.

12. Fish and other aquatic organism breathe in oxygen which is

A. Combined with water.
B. Present in water molecule.
C. So plentiful in the air.
D. Dissolved in water.

UNIT
2

Atomic structure, The Periodic Table and Bonding

Learning Outcome

Knowledge and understanding	Skills	Attitudes
• Understand simple atomic structure, the periodic table and bonding. • Know the numbers of protons, neutrons and electrons in atoms and ions given an atomic mass • Understand the formation of ions	• Calculate relative atomic masses from isotopic compositions. • interpret and use symbols such as $^{23}_{11}\mathrm{Na}$. • Design experiments to investigate properties of ionic and covalent compounds drawing on a range of sources. • Carry out experiments to investigate properties of ionic compounds and covalent compounds.	• Appreciate how knowledge of atoms contributes to our Understanding about electricity.

Introduction

Discussion corner

1. How did you define an atom from your chemistry secondary 1?
2. (a) What is the smallest thing you know that occupies space, has mass and cannot be divided any further?
   (b) Why do you think it is the smallest?

2.1 Structure of the atom

Activity 2.1 Research Activity

Individually:

Carry out a research using reference materials in the library. Come up with the appropriate definition on what atoms are.Is the theoretical definition same as the one you gave in the introduction section?

The facts

In secondary 1, you learnt that matter is made of tiny particles.Different substances are made when particles combined in different forms to form various substances around us.Each substance consists of smaller subtances called elements.Elements are made of smaller particales called atoms.

Fig 2.1 shows the structure of the atom. it is made up of sub atomic particles namely protons,electrons and neutrons. **Protons** and **neutrons** are found a at the centre atom. This location is called the **nucleus**. **Electrons** are found outside the nucleus but surrounding the atom.They occur in energy levels or orbits.

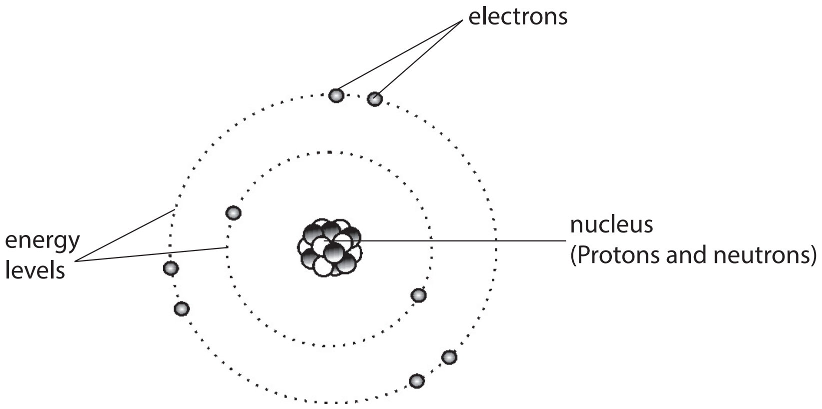


Fig 2.1: Structure of the atom

The nucleus of the atom is positively charged due to the presence of protons which are positively charged. Neutrons carry no charge.

Electrons are negatively charged and keep on moving round the nucleus experiencing a force of attraction from the nucleus.

Table 2.1 summarizes the characteristics of the sub-atomic particles.

Table 2.1 Properties of sub-atomic particles

Particle and its symbol	Relative Mass	Relative charge
Protons (p)	1	+1
Electrons (e)	1/1840	-1
Neutrons (n)	1	0

Note: The number of protons is equal to the number of electrons. For this reason, the atom is considered **electrically neutral**.

In a neutral atom:

Number of protons = Number of electrons

This means that each proton(+ve) has an electron (-ve) as a partner.

Atomic characteristics

The facts

The number of protons in the nucleus of the atom is always referred to as the **atomic number**.

The symbol for the atomic number is (**Z**).

When the total number of protons and neutrons are added we get the **mass number**.

The symbol for mass number is (**A**).

We can use the numbers on any atom to work out the number of protons, neutrons and electrons.

Using the relationship:

$\mathbf{A = Z + N}$

Where **N** is the number of neutrons in an atom, we can work out the number of neutrons as follows.

Example

The mass number of sodium is 23 and the number of protons in its atom is 11. Calculate the number of neutrons in an atom of sodium.

If $\mathbf{A = Z + N}$ then $\mathbf{N = A - Z}$

Therefore, **N**=23-11=12, the number of neutrons is hence 12.

The difference in number of neutrons leads to difference in mass number.

One element may have more than one type of atom based on differences in number of neutrons. Such atoms are known as **isotopes**.

Isotopes have the same chemical behaviour but slightly different physical properties.

Example:

There are two isotopes of chlorine.

- Chlorine – 37
- Chlorine – 35

With the knowledge of mass number, atomic number and the chemical symbol, an atom of an element can be represented using the symbol of the element. On the symbol, mass number is written as a superscript while the atomic number is written as a subscript. Both are written before the symbol.

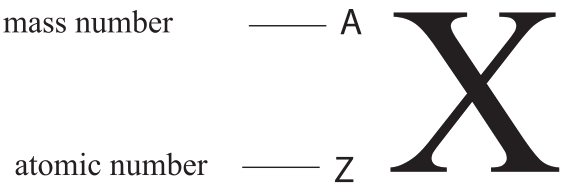


Fig. 2.2: Atomic characteristic

Can you think of how the symbol of sodium atom would be if represented this way?

Relationship between the number of protons, neutrons and mass number of some elements

Work to do

Individually:

Study the table on page 27, what, is the relationship between the number of protons, neutrons and mass numbers?

Atom	Symbol	Number of protons: Atomic number, Z	Neutrons	Mass number, A
Hydrogen	H	1	0	1
Carbon	C	6	6	12
Nitrogen	N	7	7	14
Sodium	Na	11	12	23
Chlorine	Cl	17	18	35

The facts

From the Table, we notice that mass number, A, is a sum of protons and neutrons i.e

Mass number A = number of protons Z + number of neutrons N. Therefore, to get the number of neutrons N, in the nucleus of an atom, use Equations:

$N = A - Z$

Worked Examples

Example 1

Calculate the number of neutrons in a chlorine atom given that the atomic number, Z = 17 and mass number, A = 35

The number of neutrons $N = A - Z$

$35 - 17 = 18$

Example 2

A certain element X has atomic number 6 and mass number 14.

Draw the atomic structure of element X showing all the sub-atomic particles.

The sub-atomic particles to show are protons, neutrons and electrons.

Atomic number = number of protons = number of electrons = 6

Number of neutrons = mass number – atomic number or A - Z

= 14 – 6

= 8 neutrons

This means that we will show 6 protons and 8 neutrons in the nucleus. The 6 electrons will be indicated in the energy levels as shown in Fig. 2.3.

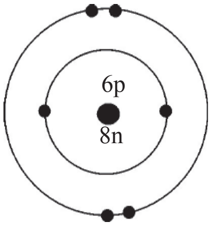


Fig. 2.3: Atomic structure of element X

Isotopes

Discussion corner

In pairs

1. Study the information in the table.

Table 2.2 Examples of isotopes

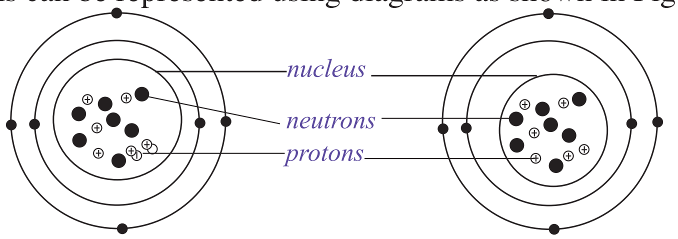
Hydrogen	Hydrogen (99.99%)	Deuterium (0.01%) Heavy hydrogen	Tritium (Trace)
Symbols of isotopes of hydrogen	$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
No. of protons	1	1	1
No. of neutrons	0	1	2
No. of electrons	1	1	1
Mass number (p + n)	1	2	3

2. Given above are the isotopes of hydrogen, can you come up with explanations of what you think isotopes are?

The facts

Isotopes are atoms of the same element with the same number of protons but different number of neutrons. This phenomenon is called **Isotopy** i.e the existence of atoms of the same element with the same number of protons but different numbers of neutrons.

For example Carbon atoms exist in two forms represented as $^{12}_{6}C$ and $^{14}_{6}C$. The two atoms can be represented using diagrams as shown in Fig. 2.4.



Carbon-12 *Carbon-14*

Fig. 2.4: Carbon isotopes

What difference do you notice between the two carbon atoms? The bottom numbers are the same but the top numbers are different. This means that they have the **same atomic number** but **different mass numbers**. What causes the difference in mass number? Calculate the sub-atomic particles for each of the carbon atoms.

You will notice that the different masses are caused by 2 **extra neutrons** in the nucleus of carbon-14. Such atoms are called **isotopes**.

Some elements are made up of just one type of atom, while others exist as mixtures of isotopes. Table 2.3,2.4 and 2.5 below show isotopes of carbon, chlorine and oxygen.

Table 2.3 Isotopes of carbon

Carbon	**Carbon-12 (98.9%)**	**Carbon -13 (1.1 %)**	**Carbon -14 (trace)**
Symbols of carbon isotopes	${}^{12}_{6}C$	${}^{13}_{6}C$	${}^{14}_{6}C$
Protons	6	6	6
Neutrons	6	7	8
Electrons	6	6	6
Mass number (p + n)	12	13	14

Table 2.4 Isotopes of chlorine

Chlorine	**Chlorine -35 (75%)**	**Chlorine -37 (25%)**
Symbols of chlorine isotopes	${}^{35}_{17}Cl$	${}^{37}_{17}Cl$
Protons	17	17
Neutrons	18	20
Electrons	17	17
Mass number (p + n)	35	37

Table 2.5 Isotopes of oxygen

Oxygen	Oxygen -16	Oxygen -17	Oxygen -18
Symbols of oxygen isotopes	${}^{16}_{8}O$	${}^{17}_{8}O$	${}^{18}_{8}O$
Protons	8	8	8
Neutrons	8	9	10
Electrons	8	8	8
Mass number	16	17	18

Relative Atomic Mass

Atoms are so tiny that their masses cannot be measured on a balance or scale directly. Hydrogen is the lightest known element. Hydrogen atom is therefore the lightest.

It was initially used as the standard element and its mass was arbitrarily fixed as one atomic mass unit (a.m.u). As a result the mass of any other element was found by comparing its mass with that of hydrogen. The idea was to find out how many times the atom of another element is as heavy as one atom of hydrogen hence relative atomic mass (R.A.M) with a symbol (Ar). This can be expressed mathematically as follows.

$$\text{Relative Atomic Mass} = \frac{\text{mass of 1 atom of the element}}{\text{mass of 1 atom of hydrogen}}$$

For instance an oxygen atom (O), has a mass of 16. This means that one oxygen atom is 16 times heavier than one atom of hydrogen.

Sometimes the symbol A_r is used for R.A.M with the symbol of the atom in parenthesis after the symbol. For example, A_r (O) means relative atomic mass of oxygen.

Many changes have occurred since the *hydrogen scale* was introduced. In the twentieth century, the hydrogen atom was replaced by oxygen as the standard element since oxygen combines with most elements. Later the oxygen scale was found to be unsatisfactory because oxygen has several isotopes which would have different masses depending on which oxygen isotope was being considered (See Table 2.5).

In 1961, the International Union of Pure and Applied Chemistry (IUPAC) recommended carbon isotope carbon - 12 (${}^{12}_{6}C$ which is 98.9%) as the standard element.

Nowadays, the atomic mass of any element is measured by comparing it to the mass of one atom of a carbon-12 (${}^{12}_{6}C$).

Therefore, the **relative atomic mass** (A_r) of an element, is defined as *the average mass of one atom of the element compared with* $\frac{1}{12}$ *of the mass of one atom of carbon - 12.*

One atom of $^{12}_{6}C$ isotope is taken to have a mass of 12.00 a.m.u.

of the mass of one atom of carbon – 12 = 1.00 a.m.u

Hence

$$A_r \text{ of an element} = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12} \times \text{mass of one atom of carbon-12}}$$

Note: Relative atomic mass has no units. It is a ratio of two masses. The relative atomic masses are not whole numbers like mass numbers. This is because the abundance of isotopes of an element is different.

For example a sample of chlorine is a mixture of two isotopes, ^{35}Cl and ^{37}Cl in the ratio of 3 : 1 respectively, i.e 75% is ^{35}Cl and only 25% is ^{37}Cl.

Calculation of the Relative Atomic Mass

As mentioned earlier, many elements naturally consist of a mixture of isotopes. The abundance in which the isotopes occur in an element differs in different elements. This is why the term 'average' mass of one atom is used in the definition above. The proportions (abundance) in which the isotopes occur in an element may be stated as:

a) a ratio b) percentage of the total c) a fraction of the total

Worked Examples

Example 1

Chlorine consists of two isotopes, i.e chlorine – 35 and chlorine – 37 in the ratio 3 : 1 respectively.i.e. 75% is ^{35}Cl and 25% is ^{37}Cl. Calculate the relative atomic mass (R.A.M) of chlorine.

Solution

Suppose the sample contains 4 atoms of chlorine, in the ratio 3 : 1, respectively, then 3 atoms will each have a mass of 35 and 1 atom will have a mass of 37.

The total mass of $^{35}Cl = 35 \times 3$

While the total mass of $^{37}Cl = 37 \times 1$

Therefore, the average mass of chlorine atoms will be

$$\frac{\text{Total mass of all atoms}}{\text{Total } \textbf{number} \text{ of atoms}} = \frac{(35 \times 3) + (37 \times 1)}{4}$$

The R.A.M = 35.5

Note: R.A.M has no units. It is not the mass number since chlorine has 2 mass numbers; 35 and 37.

The R.A.M is also nearer to the mass number of the more abundant isotope, i.e, ^{35}Cl.

Example 2

Silicon (Z = 14) consists of three isotopes: silicon – 28, 92.2%, silicon – 29, 4.7% and silicon - 30, 3.1%.

Find the relative atomic mass of silicon.

Solution

Percentage abundance simply means that if we have 100 atoms of an element called silicon, 92.2 atoms will each have a mass of 28.

Therefore, the total mass of these $= \dfrac{28 \times 92.2}{100}$

4.7 atoms will be silicon – 29

Therefore, the total mass of these $= \dfrac{29 \times 4.7}{100}$

3.1 atoms will be silicon – 30

Therefore, the total mass of these $= \dfrac{3.1 \times 30}{100}$

The average mass of a silicon atom is

$$\text{total mass of all atoms} = \frac{28 \times 92.2}{100} + \frac{4.7 \times 29}{100} + \frac{3.1 \times 30}{100}$$

Therefore, R.A.M of silicon = 28.1

Example 3

A sample of an element X consists of $\frac{9}{10}$ of ^{16}X and $\frac{1}{10}$ of ^{18}X

Show that the relative atomic mass of X would be 16.2

Solution

$$\text{R.A.M} = (16 \times \frac{9}{10}) + (18 \times \frac{1}{10}) = 16.2$$

Check your progress 2.1

1. Name the particles that contribute to the mass of an atom.
2. An element X consists of isotopes of masses 10 and 11 with a percentage of 18.7% and 81.3% respectively. Calculate the R.A.M of X.
3. An atom Q has atomic number of 11 and a relative atomic mass of 23. A teacher asked her students to models atom Q and place it in its position in Periodic Table How many electrons, protons and neutrons did the students model for atom Q?

Arrangement of electrons in an atom

Activity 2.2:

In pairs

1. Complete the table below.

Element and symbols	**Atomic number (z)**	**Mass number (A)**	**Sub-atomic particles**			**Structure of the atom**
			p	**n**	**e^-**	
Hydrogen (H)	1	1	1	0	1	
Helium (He)	2	4	2	2	2	
Lithium (Li)	3	7	3	4	3	
Beryllium (Be)	4	9	4	5	4	
Boron (B)	5	11	5	6	5	
Carbon (C)	6	12	6	6	6	
Nitrogen (N)	7	14	7	7	7	
Oxygen (O)	8	16	8	8	8	

2. Using the information you learnt in chemistry under atomic characteristics and eletronic configuration;
3. Make a sketch of each atom in the last column of the table.
6. Compare your work with other class members.

Energy levels

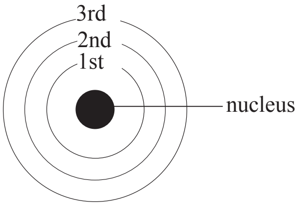


Fig 2.5. Energy Levels

The first energy level (**labelled 1**) can hold up to only two electrons. The second energy level (**labelled 2**) can hold a maximum of eight electrons. This energy level is filled after the first energy level and before the third level.

The third energy level (**labelled 3**) can hold a maximum of 18 electrons, however when 8 electrons are filled there is a degree of stability; and other remaining electrons are added to the fourth energy level.

Fig 2.6 and 2.7 show electron arrangment of sodium and calcium calcium atoms respectively.

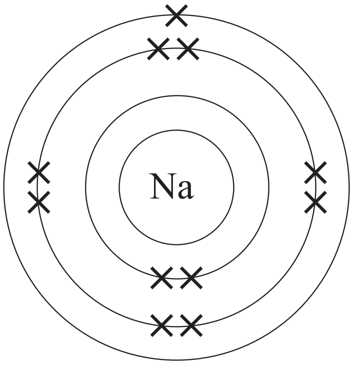


Fig 2.6: sodium (Na) structure

Number of protons=11, number of electrons=11

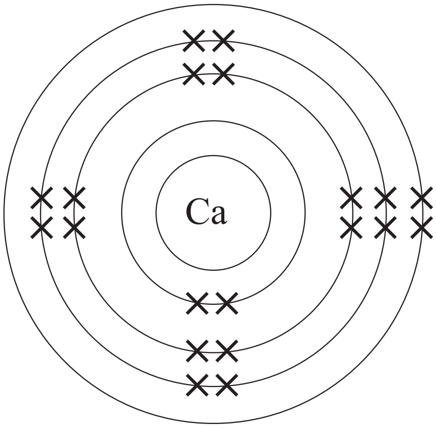


Fig 2.7: Shell model for calcium (Ca) structure

Number of protons = 20, number of electrons = 20

Note: Electrons must first fill the first energy level before they start occupying the second, third and fourth energy level.

Work to do

In Pairs

1. Draw the structure of these atoms on a manila paper.

a) Fluorine b) magnesium c) Aluminium d) phosphrous. Colour them and then hang on your classroom wall.

Check your progress 2.2

1. The symbol of an element is represented by $^{27}_{13}X$

   (a) Calculate the number of neutrons in the atom of X.

   (b) State the number of electrons in the atom of X.

   (c) Write the electronic configuration of X.

2. Study and complete the table below.

Sub atomic particle	Symbol	Charge	Location in the atom
		-1	
Neutron			
			Nucleus

3). Explain the following terms:

a) Atomic number.
b) Mass number.
c) Relative atomic mass.
d) Isotopes.

2.2 The Periodic table

Activity 2.3

In pairs

Materials required

- 2 graph papers
- Pair of scissors or razor blades
- Pens or pencils
- Glue stick
- Rulers

Procedure

1. Draw twenty squares on the graph papers (4 cm × 4 cm).
2. Cut out the squares.
3. On the top left hand corner of each square, label A on one side and B on the other side as illustrated in Tables 2.6 and 2.7.
4. On side A, write the electron arrangement of the element only. For example if it is magnesium, just write 2.8.2 as illustrated in Table 2.6.
5. On side B, write the actual symbol of the element at the centre of the paper. Include the atomic and mass numbers of the elements e.g. as in Table 2.7. Write the name of each element below the symbol.
7. With side A facing up, shuffle or mix the squares or rectangles randomly.
8. Take a piece of manila paper and draw a table of eight columns (up-down) and four rows (left to right) as shown in Table 2.8. The squares should be 5 cm × 5 cm.
9. Place the pieces of paper, with side A facing up on the squares on the manila paper as follows.
    (i) Those with one energy level only, place them in the first row. If it has 1 electron in the outermost energy level place it in column I. If it has 2 electrons in the outermost energy level place it in column II.

Table 2.6 Side A

A 1	A 2	A 2 .1	A 2 .2	A 2 .3
A 2 .4	A 2 .5	A 2 .6	A 2 .7	A 2 .8
A 2 .8. 1	A 2 .8. 2	A 2 .8. 3	A 2 .8. 4	A 2 .8. 5
A 2 .8. 6	A 2 .8. 7	A 2 .8. 8	A 2 .8. 8. 1	A 2 .8. 8. 2

Table 2.7 Side B

B $^{1}_{1}H$ Hydrogen	B $^{4}_{2}He$ Helium	B $^{7}_{3}Li$ Lithium	B $^{9}_{4}Be$ Beryllium	B $^{11}_{3}B$ Boron
B $^{12}_{6}C$ Carbon	B $^{14}_{7}N$ Nitrogen	B $^{16}_{8}O$ Oxygen	B $^{19}_{9}F$ Fluorine	B $^{20}_{10}Ne$ Neon
B $^{23}_{11}Na$ Sodium	B $^{24}_{12}Mg$ Magnessium	B $^{27}_{13}Al$ Aluminium	B $^{28}_{14}Si$ Silicon	B $^{31}_{15}P$ Phosphorus
B $^{32}_{16}S$ Sulphur	B $^{35}_{17}Cl$ Chlorine	B $^{40}_{18}Ar$ Argon	B $^{39}_{19}K$ Potassium	B $^{40}_{20}Ca$ Calcium

(ii) If it has two energy levels, place it in row two. If there is only 1 electron in the outermost energy level, place it in column I; if it has 2 electrons in the outermost energy level, place it in column II and so on.

(iii) Place those with three energy levels in row three, and those with four energy levels in row number four.

Always place them under the correct columns according to the number of electrons in the outermost energy level.

(iv) Once you have arranged your papers neatly do not disturb them.

(v) Turn each piece without changing its position to side B. Arrange them neatly.

(vi) Now compare the pattern with an actual periodic table of elements. see appendix 1.

(vii) Paste your pieces of papers on the manila paper.

Table 2.8

	Column I	Column II	Column III	Column IV	Column V	Column VI	Column VII	Column VIII
Row one								
Row two								
Row three								
Row four								

The facts

The Periodic Table is organised as a big grid. In a grid, there are **rows** and **columns**. The periodic table has rows and columns too and they each mean something different. The elements are placed in specific columns because of the way they behave. They have **similar chemical properties** because they have the same number of electrons in their outermost energy level.

In a periodic table, each of the rows is called **Period.**

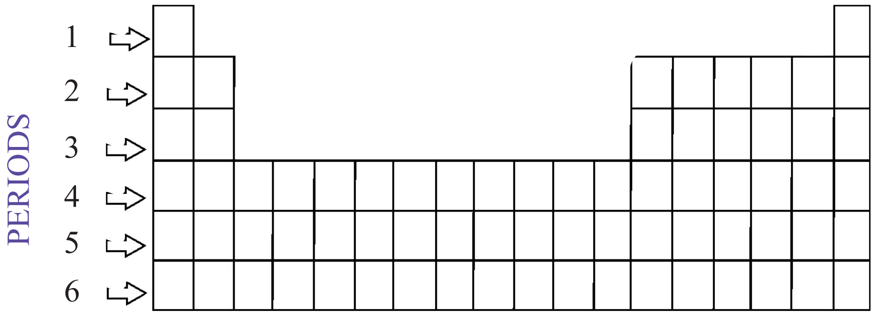


Fig. 2.8: Periods in the periodic table

Elements which are located in the same row or period have something in common. All the elements in a period have the **same number of energy levels**. If we want to know in which period to find an element we count the number of energy levels. If it has 4 energy levels; it is in the 4th period. Every element in the top row (the first period) has one energy level. All the elements in the second row (the second period) have two energy levels. The maximum number of energy levels is seven representing period seven.

Therefore, the table has a maximum of seven periods.

The periodic table has a special name for its columns too. A column in the periodic table is called a group. The elements in a group have the same number of electrons in their outermost energy level. Therefore periodic table has a maximum of eight groups.

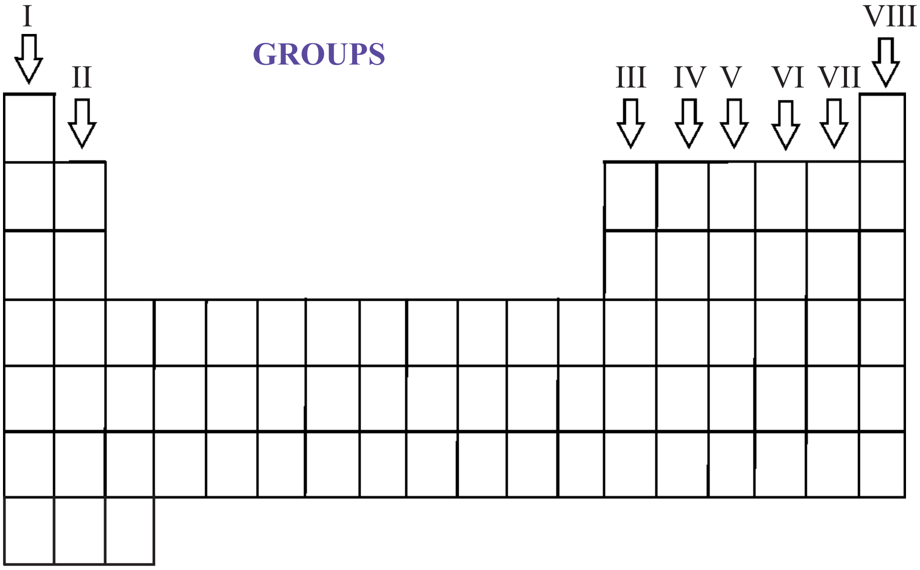


Fig. 2.9: Groups in the periodic table

An element in the first column belongs to group I. It has *one* electron in its outermost energy level. An element in the second column is in group II and has two electrons in the outermost energy level. As we keep counting the groups, we know how many electrons are in the outermost energy level. Therefore, the group number also is the number of electrons in the outermost energy level.

Check your progress 2.3

1. X,Y and Z are three unkown elements with atomic number 12,13 and 17 respectively.
   (i) Write the electronic arrangment of each element.
   (ii) Identify the group that each element belongs to.
   (iii) Which ones are metal?
   (iv) which one is a non-metal?
   (v) What is the valency of each of elements?
2. Put the following letters A to H in the correct places in the periodic table above to fit these descriptions.

Groups

Periods	I	II		III	IV	V	VI	VII	VIII
1									
2			▨						
3			▨						
4									

A. An element with 7 protons.

B. It belongs to period three and has 4 electrons in the outermost energy level.

C. An element with oxidation number +3 and with 13 protons.

D. An element that forms its ion by loss of one electron and belongs to period four.

E. An element that can be placed in either group I or group VII.

F. An element with the highest number of energy levels.

G. An element with electronic arrangement of 2.8.8.

H. An element with valency of 3. It gains electrons to form ions and belongs to period three.

2.3 Ion formation

Activity 2.4

Individually:

1. Study the elements in the table below.

Element	Symbol	Atomic Number	Electron arrangement
Helium	He	2	2
Neon	Ne	10	2.8
Argon	Ar	18	2.8.8

2. Draw the structure of each element in your notebook.What do you discover?

The facts

Group VIII elements are described as *stable*. This means that their outermost energy levels have the maximum number of electrons that they can possibly accomodate.

Octet duplet rule

The process of achieving a *stable* electron arrangement of 8 electrons in the outermost energy level is known as *octet rule*. Note that in helium, stability is achieved by acquiring two electrons in the outermost energy level.*This is known as duplet rule* respectively.

In order to achieve a more stable outermost electron arrangement of 2 or 8 electrons, electrons must be gained or lost in the outermost energy level. When an atom gains or loses an electron or electrons, the particle formed is charged either negatively (–ve) or positively (+ve). Such a particle is called an **ion**.

Formation of lithium and magnesium ions

Activity 2.5

In pairs

1. Study the diagram below.

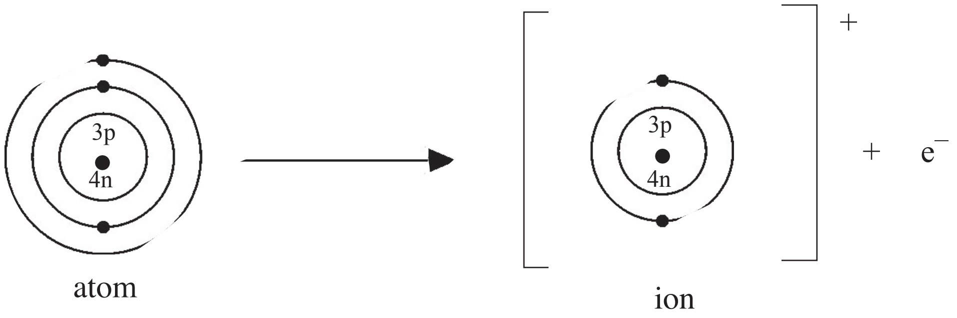


Fig 2.10: Formation of lithuim ion

2. What has happened to the lithium atom?

The facts

A lithium atom has an electron arrangement of 2.1. It has only one electron in the outermost energy level. For it to be stable, it can either lose one electron to have an electron arrangement of 2 or gain seven electrons to attain an arrangement of 2.8.

More energy will be required for the atom to gain 7 electrons than to lose 1 electron see Fig. 2.11 Therefore, it is easier for a lithium atom to lose 1 electron to form an iron.

How about magnesium?

For magnesium, it is easier to lose two electrons than gain six electrons.

Therefore a magnesium ion is formed by loss of two electrons as follows.

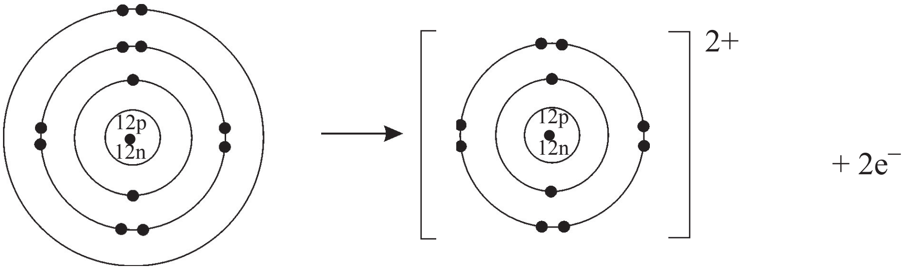


Fig. 2.11: Formation of magnesium Ion

Li^+ and Mg^{2+} are called **cations**. This is because they are positively charged ions. Cations are formed when atoms lose one or more electrons.

Work to do

How are sodium and aluminium ions formed?

Follow the same procedure for lithium and magnesium to work this out.

Formation of fluoride and sulphide ions

Activity 2.6

In pairs

1.Study the diagram below.

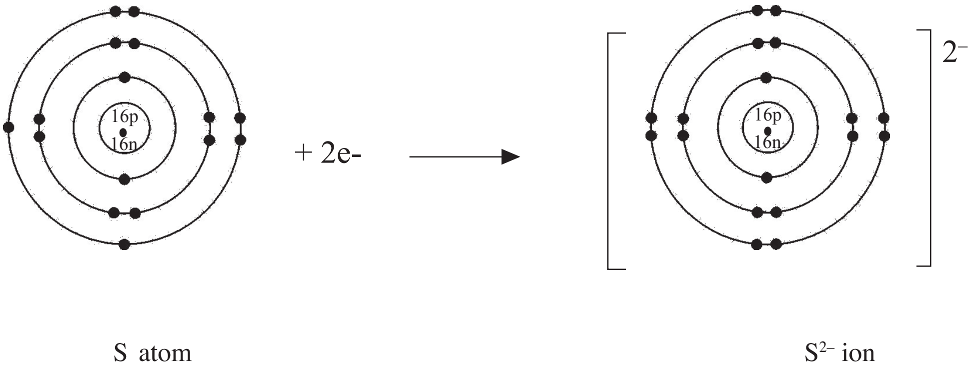


Fig. 2.12: Formation of sulphur ion

2.What has happened to the sulphur atom.

The facts

Sulphur atom has the electron arrangment of 2:8:6. For sulphur it is easier to gain two electrons than to lose electrons.

Therefore, a sulphur ion is formed by gaining electrons

A fluorine atom has the electron arrangement of 2.7. For fluorine atom to attain a stable electron arrangement of 2.8, it will either lose seven electrons or gain one electron respectively. Much more energy will be required for fluorine to lose seven electrons than to gain one electron. For this reason it is easier for fluorine to gain one electron than to lose the seven electrons. Therefore, a fluorine atom forms a **fluoride** ion by gaining one electron.

F^- and S^{2-} are anions. They are negatively charged. Negatively charged ions are called ***anions***. They are formed when a neutral atom gains one or more electrons.

Work to do

- How is a chloride ion formed? Follow the same procedure as for fluorine and sulphur to show the formation of a chloride ion.

Ionisation energy and electron affinity

We have seen from the structure of atoms that protons in the nucleus which are positively charged (+ve) attract electrons which are negatively charged (–ve) and located in the energy levels. Therefore, in order to remove electron, we must overcome this force of attraction. In other words we must supply energy to pull off the electron(s). The energy supplied is called *ionisation energy*.

The equation for loss of electrons is represented as:

$M(g) \rightarrow M^+(g) + e^-$ M-represents a metal atom

If an atom loses 2 electrons the equation is as follows:

$$M(g) \rightarrow M^{2+}(g) + 2e^-$$

Note the charge on the ion is the same as the number of electrons lost. If an atom loses three electrons, then the charge will be 3^+, etc

$$M(g) \rightarrow M^{3+}(g) + 3e^-$$

Therefore, **Ionisation energy** is *defined as the minimum amount of energy required to remove electron(s) from an atom in gaseous state*.

Electron affinity

We have seen that non-metals gain electron(s) to attain stability, remember that electrons are negatively charged. Therefore, when an electron tries to enter the outermost energy level, it will be repelled by the electrons which are already there. So some "force" is required to put the electrons into the energy level. This "force" is a form of energy. This energy is known as *electron affinity*. The ions formed are always negatively charged.

For example:

Chlorine - $Cl(g) + e^- \rightarrow Cl^-(aq)$

Sulphur - $S(g) + 2e^- \rightarrow S^{2-}(aq)$

2.4 Bond formation

The electronic arrangement of noble gases, as discussed earlier, is very stable. For example, helium 2, neon 2.8 and argon 2.8.8. They have the stable duplet (2) for helium and octet (8) for the others, in the outermost energy levels. Consequently, noble gases are in most cases chemically unreactive and do not form compounds with other elements.As they do this they end up forming chemical bonds. The tendency among other elements is to strive to attain the stable noble gas electron arrangement. As they do this they end up forming chemical bonds. Some will do this by donating electrons; others by accepting them; and still others will share electrons.**A chemical bond is therefore a lasting attraction between ions, atoms or molecules that enables formation of chemical compounds**.

Ionic bonding

Activity 2.7

In pairs

1. Study what is happening in the diagram below.

(a) Before an electron is transferred.

(b) After electron transfer.

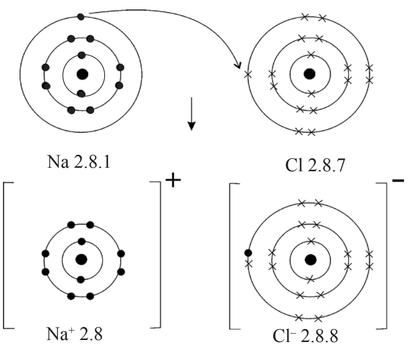


Fig. 2.13: Formation of NaCl

2. State what holds the natural CI^- ion together in the final molecule.

The facts

Ionic bonding is also known as **electronic bond** and it involves complete transfer of electrons from one atom to another. This results in formation of opposite charges which attract each other forming a bond. This kind of bond is referred to as **electrostatic attraction** since each atom carries a fixed (static) charge. The force holding the two ions together is known as **electrovalent bond** or **ionic bond**. An example of a compound whose particles are held together by ionic bond is sodium chloride.

Examples

(i) Ionic bond in sodium chloride

A sodium atom contains 11 protons and has an electronic arrangement of 2.8.1. The arrangement differs from the nearest noble gas electronic structure, that of neon, 2.8 by the presence of one extra electron in the third energy level. On the other hand, a chlorine atom contains 17 protons and has the electronic arrangement of 2.8.7. It differs from the nearest noble gas electronic arrangement, that of argon, 2.8.8, by missing one electron in the third energy level. In order to attain the stable noble gas electron arrangement, a sodium atom would have to lose the electron in the outermost energy level. The chlorine atom would need to take one electron into its outer energy level to gain the noble gas structure. During ionic chemical bonding of sodium and chlorine atoms, the single electron from the outermost energy level of sodium atom is transferred to the outermost energy level of the chlorine atom.

A sodium ion is positively charged because of the one excess positive charge. Similarly chloride ion is negatively charged because of the one excess negative charge. These ions have opposite charges and therefore attract.

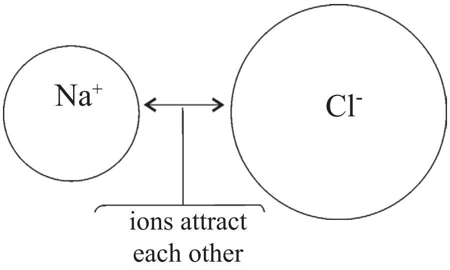


Fig 2.14: Sodium ion attracts chloride ion

Structure of sodium chloride

The structure of sodium chloride contains numerous sodium and chloride ions in equal proportions. The electrical attraction (electrostatic attraction) resulting from their opposite charges constitutes the ionic bond and therefore attract.

The ions arrange themselves into a rigid solid shape called a **crystal**. Each sodium ion is surrounded by six (6) equidistant chloride ions and vice versa.

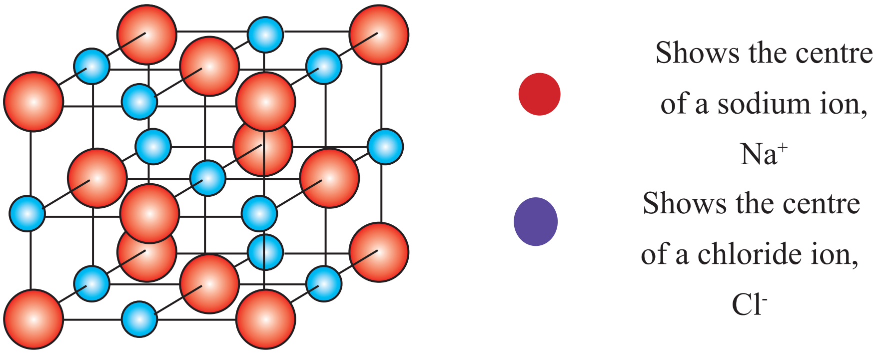


Fig 2.15: Arrangement of sodium and chloride ions in sodium chloride crystal

The ions form a giant ionic structure.

The attraction forces between the ions are strong and therefore the ions are not free to move but they vibrate within a given space. Consequently the melting point of sodium chloride is high. In solid form, it is a non-conductor of electricity. We will see later that when sodium chloride is melted sodium ions and chloride ions separate and thus their forces of attraction are greatly reduced. When an electric current is applied, the ions in molten sodium chloride are free to move thereby conducting electricity.

The positive ions formed as a result of loss of one or more electrons are called **cations**, and their positive charges are equal to the number of electrons lost. Likewise, the negative ions formed as a result of gain of one or more electrons are called **anions** and their negative charges are equal to the number of electrons gained.

Also note that the number of electrons lost from, or added to, the outermost energy level of the atom of an element during ionic bonding is equal to the combining power (valency) of that element. Only the **outermost** energy level electrons are involved in ionic bonding. The number of ions involved must balance the valency requirements of elements.

Properties of ionic compounds

Activity 2.8

To investigate the properties of ionic compounds.

In groups

Materials and apparatus

Distilled water, beakers, sodium chloride salt, stirring rod, bulb, Copper wires or graphite rods, crucible, paraffin.

Procedure 1

1. Take a crystal of sodium chloride and feel it between your fingers.
2. Put a half-full spatula of sodium chloride salt into a beaker that is half-filled with water.

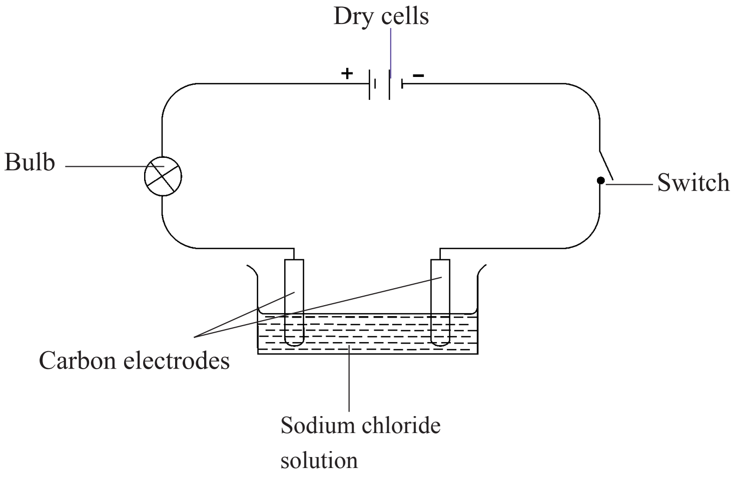


Fig 2.16: Experkmental set up

3. In this activity, you are investigating wheather NaCl will be able to conduct electricty. Using your knowledge from chemistry book 1 under Acids, bases and salts; predict on what you think will happen. Will NaCl be able to conduct electricity? why do you think this is so?
4. Stir the mixture using a stirring rod.
5. Repeat the procedure using paraffin instead of distilled water.
6. Record your observations.
7. Compare your results to your predictions

Procedure 2

1. Place some crystals of sodium chloride into a crucible.
2. Heat until there is no further change.

3. Set up apparatus as in figure 2.16 but using the sodium chloride crytals instead of sodium chloride solution.
4. Predict what will happen when your use crystals of NaCl instead of NaCl solution. Explain your answer.
5. Compare your results to your prediction. Were your prediction correct? Record your observations.

Study questions

1. How does a crystal of sodium chloride feel between the fingers?
2. What happened to sodium chloride when it was put in distilled water and stirred?
3. Did the salt dissolve in paraffin?
4. What observations did you make in procedure 2?
5. What observations did you make step 3 of procedure 2 above?

The facts

Sodium chloride crystal has fine crystalline feel. It dissolves in water but not parrafin. In procedure 2, the bulb produced light. Sodium chloride melts when heated. Molten sodium chloride conducts electricity.

1. Ionic compounds are **electrolytes** (solutions that conduct electric current) when they are in solution or molten form.

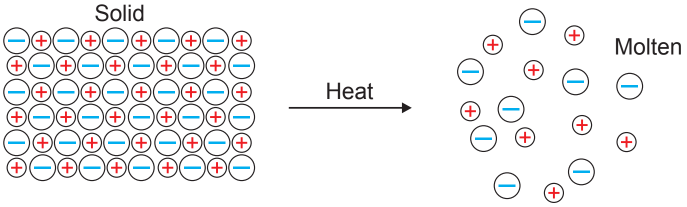


Fig 2.17: Melting of an ionic compound

2. They have high melting points (non-volatile).

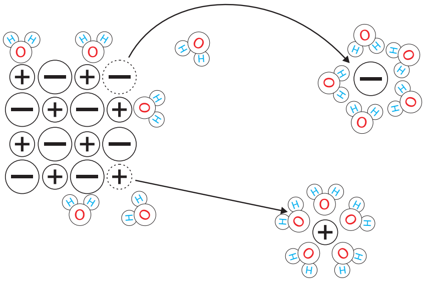


Fig 2.18: Dissolving of an ionic compound

3. They are generally soluble in water.
4. Ionic compounds are insoluble in organic liquids like benzene or propanone.
5. They are usually crystalline solids.

Check your progress 2.4

1. Distinguish between ionisation energy and electron affinity
2. What is octet rule?
3. Describe how a bond is formed.
4. Explain using electron arrangement diagrams how the following ionic compounds are formed:
    - (i) Calcium oxide
    - (ii) Lithium oxide
    - (iii) Calcium chloride

Covalent bonding

A covalent bond is a bond formed when two or more atoms share a pair of electrons.Each of the atoms contribute an electron to be shared. It can be between atoms of the same element or different elements. The electrons to be shared can be contributed by only one of the combining atoms as well. When the electrons to be shared are contributed by one species, a special type of covalent bond called a **dative or coordinative bond** is formed.

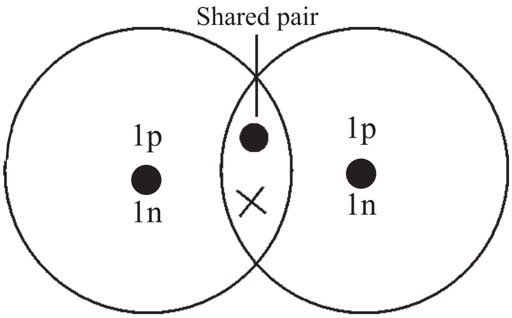
An example is the bond between hydrogen molecule, (H_2) It has two hydrogen atoms linked by a covalent bond.

Two hydrogen atoms combine by each contributing an electron. Then they share the electron pair equally. They have the stable electron duplet arrangement i.e. first energy level with 2 electrons.

Activity 2.9

In pairs

a) The figure below shows how hydrogen (H_2) molecules is formed. Study it carefully and answer the question below.



b) How does bonding take place in hydrogen molecules?

The facts

The shared pair is the covalent bond. Sometime the shared pair is represented by a short line (-) H (-)H or the pair can be shown as H_xH. Another example is bond formation in a chlorine molecule(Cl_2).These two atoms have each an electronic arrangement of 2.8.7 with 17 protons. If no other element is available from which electrons may be obtained to make these atoms have a noble gas electronic structure such as argon (2.8.8), a shared pair of electrons is formed.Each chlorine atom contributes one electron to the shared pair as follows:

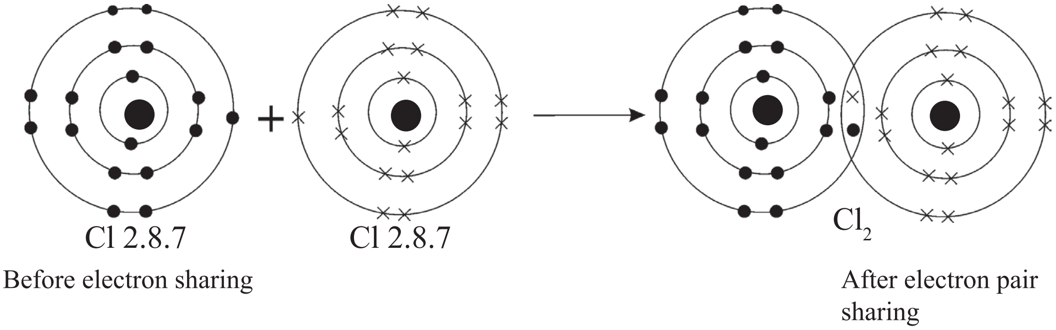


Fig 2.19: Formation of chlorine molecule

In the chlorine molecule the stable electron octet arrangement for each chlorine atom is achieved.

Some atoms can share more than one pair of electrons. The *covalency* of an atom is the number of electron pairs which it shares.

Notice also that atoms of different elements can also form covalent bonds like in carbon (IV) oxide as shown below.

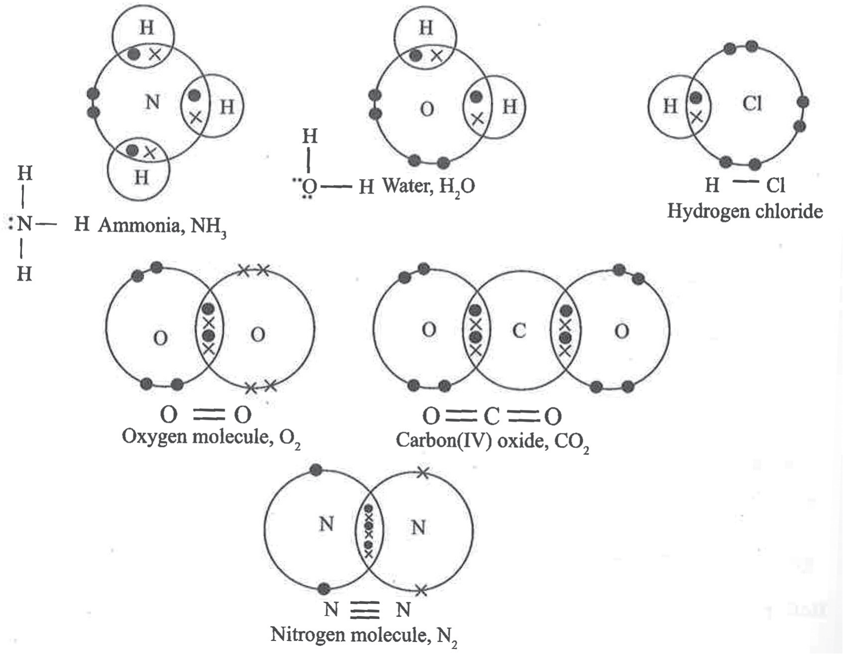


Fig 2.20: Formation of covalent bond

Properties of covalent bonds

Activity 2.10

To investigate the properties of covalent compounds.

In groups

Apparatus and reagents

Paraffin, distilled water, ethanol, potassium iodide solution, crucible, candle wax, cooking oil.

Procedure

1. Heat the candle wax in a crucible. Record your observation.
2. Mix cooking oil with water and stir. Record your observation.
3. Repeat procedure (2) using ethanol instead of water.
   - What did you observe.

Study question

1. What observation did you make when candle wax was heated?
2. Did cooking oil dissolve in water?Why
3. What did you observe when cooking oil was mixed with ethanol?

The facts

When candle wax is heated, it melts at a temperature between 46℃ and 68℃. Cooking oil does not dissolve in water but it dissolves in ethanol.These are some of the properties of covalent compounds. Others are:

1. They are non-electrolytes (don't conduct electric current) when in molten or solution form.
2. They have usually low melting and boiling points.
3. They are often insoluble in water.
4. They are usually soluble in organic solvents like benzene and propanone.
5. Most of them are gaseous at room temperature or volatile liquids.

Co-ordinate (Dative) bonding

A co -ordinate bond is a covalent bond in which both electrons shared come from only one atom. One of the atoms posses a lone pair of electrons.

This lone pair is shared by an atom or ion that needs them to attain stability. For example, ammonia molecule has a lone pair of electrons.

This lone pair can be donated to hydrogen ion (H^+) from an acid to form the ammonium ion, NH_4^+. The nitrogen atom is said to be a "donor" and the hydrogen ion is said to be an " acceptor".

$H^+ + NH_3 \rightarrow NH_4^+$

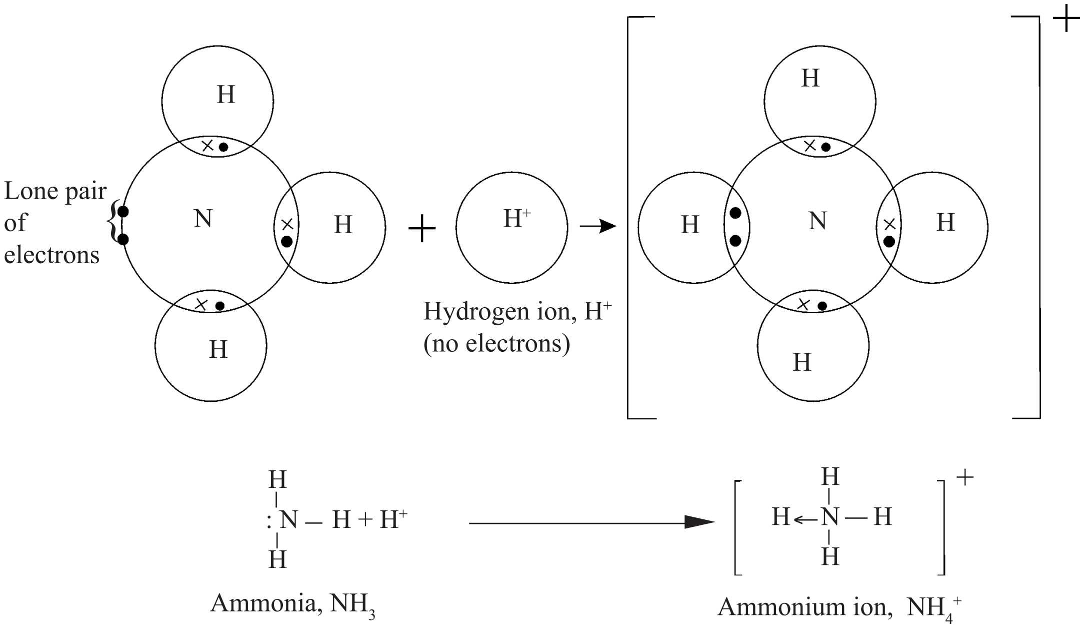


Fig 2.21: Co-ordinate bond of Ammonium ion

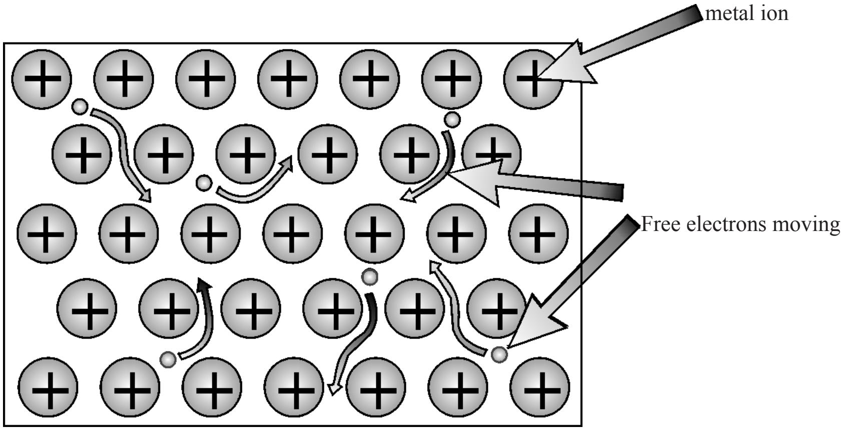
The representation of a co-ordinate bond is →. The arrow points from the donor atom to the acceptor atom as shown above. The hydrogen ion contributes the charge on the ion formed.

Metallic bonding

Activity 2.11

In pairs

1. Study the structure below.



2. Find out hows metallic bond takes place and compare with the structure above.

The facts

The outermost energy level electrons in metals are relatively few. When the atoms of metals are closely packed, each metal atom loses its outer electron(s) which form a "sea" of free electrons (**delocalised mobile electrons**). The resulting positive ions are embedded in the "sea" of electrons. There is attraction between the ions and electrons. This kind of electrostatic attraction between the positive ions and the delocalised electrons form the **metallic bond**. The ions arrange themselves into a giant metallic structure.The electron sea and the positively charged ions are responsible for the different properties of metals.

Properties of metals

Activity 2.12

In pairs

1. Study the table below.

Element	**Na**	**Mg**	**Al**
Melting point ˚C	98	650	660
Boiling point ˚C	890	1120	2450
Electrical conductivity x10^3 Ohm^{-1} at 250 ˚C	10	16	38
Thermal conductivity $kJmol^{-1}$ S^{-1} at 250˚C	1.34	1.6	2.1
Density g/cm^3	1.34	1.6	2.1

2. Make conclusion about the melting and boiling point, density and conductivity of metals.

The facts

Some properties of metals include:

1. They are good conductors of heat and electricitys

These delocalised electrons can move on application of an electric current or heat. This explains why metals are good conductors of electricity and heat. The thermal and electrical conductivity also increases across the period because of the increase in the number of valency electrons (delocalised electrons). It might be expected that increase in temperature would speed the movement of free electrons, with a consequent increase in electrical conductivity. In general, however, the electrical conductivity of metals decreases with increase in temperature. This is because increasing temperature produces increased thermal vibration within the metal structure. This upsets the regularity within the crystal and interferes with the ease of movements of electrons within the crystal. It is just like comparing soldiers matching on parade and others matching through a city crowd.

2. Metals have high boiling and melting points

Metallic bonding is very strong in some metals like copper, iron, but weak in others like sodium and potassium, which can be cut with a knife. Moving across a period (e.g. period 3) of the periodic table, the number of valency delocalised electrons increases and therefore the strength of the metallic bonds increases. The melting point increases from sodium to aluminium. The change in melting point from magnesium to aluminium is not very big. This is probably because not all the three electrons in aluminium are involved in metallic bonding.

3. They are hard with high density

As the strength of metallic bond increases across the period, the pull of positive ions towards each other increases thus also increasing the density of the metals.

My environment my life

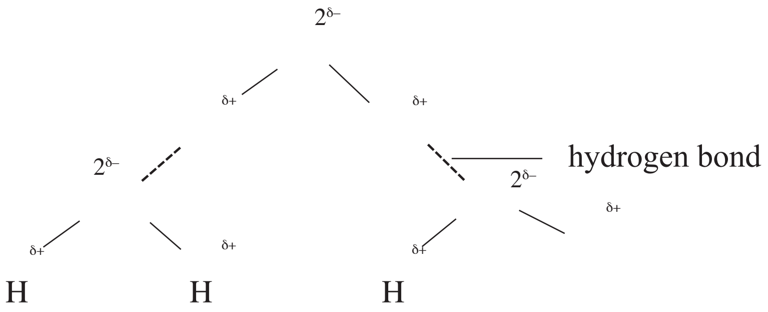
During metal extraction, we should avoid learning behind big holes after quarrying.

Hydrogen bonding

Activity 2.13

In pairs

1. Study the diagram below.



2. Identify and circle the water molecules.Can you see the hydrogen bonds?

The facts

A hydrogen bond is the electrostatic attraction between two polar groups that occur when a hydrogen (H) atom covalently bond to highly electronegative atom such as Nitrogen (N), Oxygen (O) or Flourine (F) which is attracted to another highly electronegative atom nearby.

The bonded electrons spend more time around the highly electronegative element.This makes it develop a partial negative charge while hydrogen develops a partial positive charge. The molecule is said to be **polar**.

For example in a water molecule(H_2O), oxygen atom attracts electrons more than hydrogen. Therefore, the water molecule is represented like this:

$$H^{\delta+} - O^{\delta-} - H^{\delta+}$$

The water molecules can combine together using the polar ends. The attraction between the polar end of the hydrogen ($\delta+$) and the polar end of the oxygen (δ -) is called **hydrogen bond.**

Therefore water has hydrogen bonds between molecules. Had it not been the presence of hydrogen bonds, water would be in gaseous state at room temperature and pressure.

Check your progress 2.5

1. Explain the following:
    (i) Although potassium is a very reactive metal and chlorine is a reactive non-metal, potassium chloride is unreactive.
    (ii) Electrical conductivity of metals decreases as the temperature rises.
2. Draw electronic diagrams for the following species (use dot (•) and cross (x)) and show only the outermost energy level electrons): hydroxium ion (H_3O^+), ammonium ion (NH_4^+), water molecule (H_2O and ammonia molecule (NH_3).
3. Study the table below showing properties of substances A, B, C and D. Answer the questions that follow.

Substances	Reaction with oxygen at 25°C	Melting point	Electrical conductivity	
			Solid	Molten
A	Unreactive	High	Poor	Good
B	Unreactive	High	Poor	Poor
C	Unreactive	High	Good	Good
D	Reactive	Low	Good	Good

Select the substance which is likely to be.

(a) Copper

(b) Magnesium chloride

(c) Sodium

(d) Diamond

4. Two elements A and B whose atomic numbers are 6 and 8 respectively react to form a compound C.

(i) What is the formula of C in terms of A and B.

(ii) Use dots (•) and crosses (x) to show the bonding in compound C.

5. Study the table below which gives some properties of compounds P,Q,R and S and answer the questions that follow.

	Melting point	Boiling point °C	Electrical conductivity in water
P	- 23	- 77	Does not conduct
Q	- 19	- 74	Does not conduct
R	- 85	- 61	Does not conduct
S	407	714	Conducts

(i) Which one of the compounds in the table is ionic? Explain.

(ii) Select the compound(s) that is/are liquid(s) at room temperature.

(iii) Select the compound that is gaseous at room temperature.

6. (a) Why is aluminium and copper used in making cooking pans?

   (b) The melting point of silicon is 1410°C while that of chlorine is -
   –101°C. Explain the difference.

7. With reference to chlorine:

   (a) Explain why most covalent compounds are gaseous at room temperature.

   (b) Why do silicon (IV) oxide, diamond and graphite exist as solids at room temperature while chlorine exists as a gas?

8. Explain how an eletrovalent is formed.

9. An element P^{2+} has electron arrangement of 2.8.

   (a) What is the atomic number of the element?

   (b) To which groups and period in the periodic table does the element belong?

   (c) What type of bond between the atoms of elements P.

   (d) How does the element conduct electricity when it is in molten form?

UNIT 3

Formation of salts and Electrolysis

Learning Outcome

Knowledge and understanding	**Skills**	**Attitudes**
• Understand how salts are formed. • Explain how and why electrolysis can be used to separate some salts.	• Carry out practical investigations and suggest a method of preparing a given salt from suitable starting materials. • Predict what will happen based on theory.	• Appreciate the applications of electrolysis.

Introduction

You are already familiar with the 'common salt' or sodium chloride which we add to our food every time we eat. Sea water also contains a lot of sodium chloride and for many years sodium chloride has been used in food preservation. However , in Chemistry, the term salt is widely used; many of the chemicals we study in elementary courses in chemistry are salts. We shall see in this unit that the term salt has a more general meaning.

We learnt in Secondary 1 that acids are compounds which contain hydrogen which may be replaced by a metal to give a salt and hydrogen. This reaction may be represented as follows:

Metal + acid → salt + hydrogen

Activity 3.1

In pairs

1. Study the table below.

Name	**Formula**
a) Sodium chloride	$NaCl$
b) Ammonium chloride	NH_4Cl

c) Potassium sulphate	K_2SO_4
d) Zinc nitrate	$Zn(NO_3)_2$

2. Point out the acid from which the salt was made. Explain your answer.

The facts

From the salt you can tell the acid that was used:

- Hydrochloric acid always gives **chlorides**
- Nitric acid always gives **nitrates**
- Sulphuric acid always gives **sulphates**
- Carbonic acid always gives **carbonates** among others.

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

Meaning of salt

A *salt* is a compound formed when the hydrogen ions of an acid are fully or partially replaced by a metal ion (or ammonium radical). The process involves a neutralisation reaction in which a base neutralises an acid. Names of salts are derived from the metal or ammonium radical and the acid radical from which they are formed.

3.1 Methods of preparing salts

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

- Action of an acid on a metal.
- Action of an acid on an insoluble base (metal oxide or hydroxide).
- Action of an acid on an insoluble carbonate.
- Action of an acid on an alkali.

On the other hands insolube salts are prepared by direct combination of elements or precipitation methods.

a) Preparation of soluble salts

i) Reaction of a dilute acid with a metal

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

Activity 3.2

To prepare zinc (II) sulphate.

In groups

Apparatus and chemicals

- beaker (100 cm^3)
- filter funnel
- filter papers
- evaporating dish
- conical flask
- bunsen burner
- dilute sulphuric acid
- zinc

Procedure

1. Pour 50 cm^3 dilute sulphuric acid (VI) into a beaker.
2. Add zinc granules. What do you observe?
3. If the reaction is too slow, add a little copper (II) sulphate solution and warm gently, but do not boil. What do you observe? Can you give a reason why copper (II) sulphate was added?
4. When all zinc has reacted, add more and wait until it has all reacted.
   Add more again until some remains unreacted. Give a reason why we must keep on adding zinc until no more can react.

5. Filter off the excess zinc and solid impurities and collect the filtrate in a conical flask. What is the colour of the filtrate?
6. Heat the filtrate for sometimes. Allow it to cool overnight. What do you observe in the evaporating dish?
7. Record your observations in your notebook
8. What do you conclude from the experiment? Is it possible to prepare Zinc(II) sulphate crystals using Zinc metal and sulphuric acid (VI)? Explain.

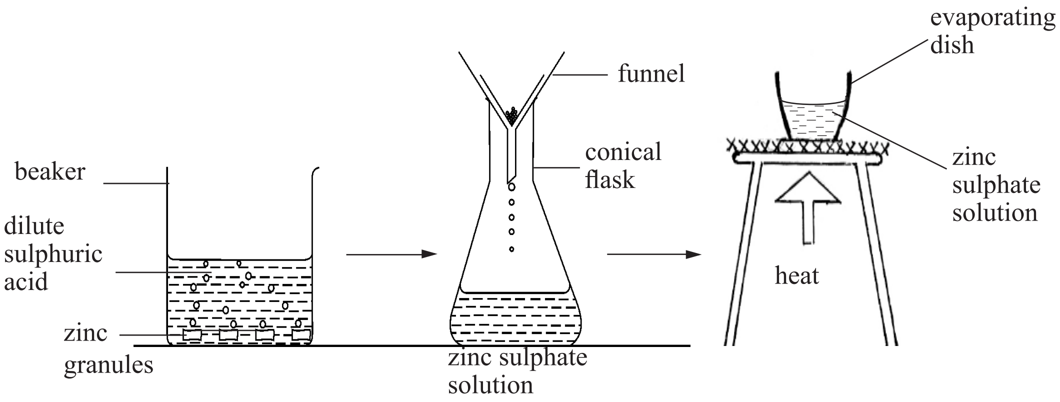


Fig. 3.1: Preparation of zinc sulphate crystal

Obtaining crystals

Study questions

As a class

Remind yourselves of what crystals are as learnt from chemisrty secondary 1.

- How do we get crystals from filtrate/
- Why is the filtrate boiled?
- How can you test the boiling filtrate for saturation?
- How are crystals dried?

The facts

When an acid comes into contact with zinc, we observe bubbles. When copper (II) sulphate solution is added to the reacting mixture, it speeds up the reaction rate. Whenever we add zinc and bubbles are observed, it means that the acid is not used up. If more zinc is added and some remain unreacted, it means all the acid has reacted and a colourless solution of zinc sulphate formed.

When excess water is evaporated and we let the saturated solution cool slowly, big crystals form. Crystals incorporate water as they form. This water is known as **water of crystallisation** and gives shape to the crystals. This explains why the salt formed when we evaporate the filtrate to dryness is a white powder. All hydrated salts contain water of crystallisation. But the number of water molecules differ in different crystals.

The reactions that occur are as follows:

(i) Formation of the salt.

Zinc + sulphuric acid (VI) ⟶ zinc sulphate + hydrogen

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

(ii) Formation of crystals from the salt

Zinc sulphate + water ⟶ zinc sulphate crystals

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

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

Iron + sulphuric acid (VI) ⟶ iron(II) sulphate + hydrogen

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

Iron(II) sulphate + water ⟶ iron(II) sulphate crystals

$$FeSO_4(aq) + 7H_2O(l) \rightarrow FeSO_4\cdot 7H_2O(s)$$

Zinc + Nitric acid (V) ⟶ zinc nitrate + water

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

Note: Metals, such as sodium, potassium and calcium react **vigorously** with acids. They should not be used to prepare salts by the method of crystallisation. Other metals such as copper, silver and mercury **do not react** with acids and hence cannot be used to prepare salts by this method. Lead metal when reacted with dilute hydrochloric acid or dilute sulphuric acid forms an **insoluble salts** (lead (II) chloride or lead (II) sulphate) which coats the metal preventing further reaction. Therefore, lead metal is not suitable for preparing salts by crystrallsation method.

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

Activity 3.3

To prepare copper (II) sulphate from copper (II) oxide.

In groups

Apparatus and chemicals

Recommend the specific apparatus chemicals are needed to perform this activity.

Procedure

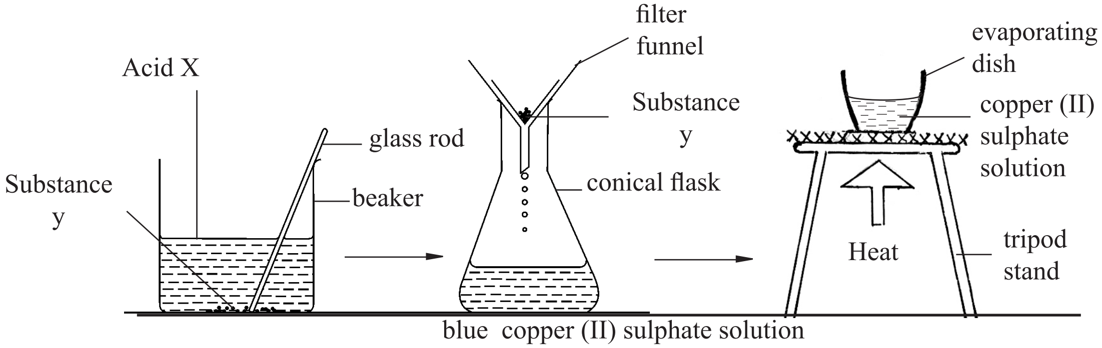


Fig 3.2: Preparation of copper (II) sulphate crystals

1. Pour 50 cm^3 of dilute acid X into a beaker.
2. Place the beaker on the tripod stand and warm gently but do not boil.
3. Add substance Y to the hot acid a little at a time. Stir and observe what happens. Test the solution with blue litmus paper ensure you note colour change.Can you be able to predict the compound formed from the change in colour.

4. Continue to add substance Y a little at a time while stirring until no more substance Y reacts with the acid. Give a reason why excess substance Y is added. Test the solution with litmus paper.
5. Filter off the excess substance Y and collect the filtrate in a conical flask.
6. Heat a little of the blue filtrate in an evaporating dish to dryness and note what happens.
7. Heat the remaining filtrate to evaporate excess water and test to see whether the solution is saturated. Leave it until next lesson.
8. When good crystals form, pour off the mother liquor.
9. Dry the crystals between two filter papers or in sunshine.
10. Observe the crystals and note the difference with the salt obtained.
11. Put a few crystals in a test tube and heat strongly.
12. Observe what happens. Place the residue on watch glass or tile and add a few drops of water.
13. Which chemical did you choose for the preparation of copper (II) sulphate?
14. What do you conclude? Why is it not possible to prepare copper (II) sulphate by reacting copper metal with dilute sulphuric acid? Explain.

The facts

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

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

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

The reactions that occurred are as follows:

(i) When the oxide reacted with acid

Copper (II) oxide + sulphuric acid (VI) ⟶ copper (II) sulphate + water

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

(ii) During crystal formation

Copper (II) sulphate + Water ⟶ hydrated copper (II) sulphate

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

(iii) On heating the blue hydrated copper (II) sulphate

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

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

(iv) On adding water to anhydrous copper (II) sulphate

Anhydrous copper (II) sulphate + water ⟶ hydrated copper (II) sulphate

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

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

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

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

(a) Magnesium sulphate from magnesium hydroxide

Magnesium hydroxide + sulphuric acid ⟶ magnesium sulphate + water

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

(b) Lead (II) nitrate from lead (II) oxide

Lead (II) oxide + nitric acid → lead (II) nitrate + water

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

Table 3.1: Examples of some other hydrated salts

Name	Formula
Magnesium sulphate -7-water	$MgSO_4.7H_2O$
Sodium carbonate -10-water	$Na_2CO_3.10H_2O$
Calcium sulphate - 2- water	$CaSO_4.2H_2O$
Cobalt (II) chloride -6-water	$CoCl_2.6H_2O$

(iii) Reaction of a dilute acid with metal carbonate or metal hydrogencarbonate

Activity 3.4: To prepare lead (II) nitrate from lead (II)carbonate

In groups

Apparatus and chemicals

- conical flask
- tripod stand
- glass rod
- lead carbonate
- beaker
- wire gauze
- filter paper
- 2M dilute nitric acid
- spatula
- bunsen burner
- funnel

Procedure

1. Pour 50 cm^3 of dilute nitric acid into a beaker.
2. Add lead (II) carbonate, a little at a time and stir. What do you observe? Why is the lead (II) carbonate added alittle at a time?
3. Test the solution with blue litmus paper to find out whether all the acid has reacted. Filter off excess carbonate. Proceed to crystallize.
4. Record observations in your notebook.

- Why is it not appropriate to use sulphuric acid or hydrochloric acid when preparing lead (II) sulphate or chloride from lead (II) carbonate?
- Which other alternative you can use to prepare lead (II) sulphate?
- What is the colour of lead (II) nitrate solution?

Carry out an experiment to shows how zinc sulphate can be prepared using the method from zinc carbonate.

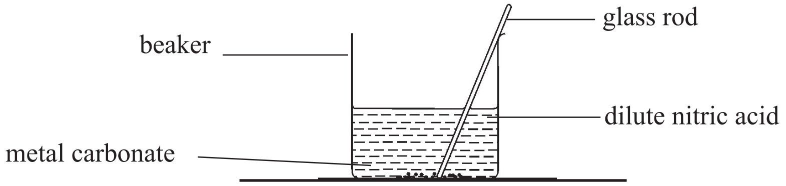


Fig. 3.3: Reacting a metal carbonate with an acid

The facts

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

Carbonate + acid ⟶ salt + carbon(IV) oxide + water.

Example

Lead (II) carbonate + nitric acid → lead (II) nitrate + carbon (IV) oxide + water

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

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

Other soluble salts that can be prepared in a similar way from the carbonates or hydrogencarbonates include; Copper (II) sulphate, copper nitrate, magnesium sulphate and zinc (II) sulphate.

(a)

(i) Copper carbonate + sulphuric acid → copper (II) sulphate + water + carbon oxide

$$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO4(aq) + H_2O(l) + CO_2(g)$$

(ii) Copper carbonate + nitric (V) acid → copper nitrate + water + carbon(IV)oxide

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

(iii) Zinc carbonate + sulphuric (VI) acid → zinc sulphate + water + carbon(IV) oxide

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

(iv) Ammonium carbonate + sulphuric (VI) acid → ammonium sulphate + water + carbon (IV)oxide

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

(b) Hydrogencarbonates can also be used as shown:

(i) Potassium hydrogencarbonate + nitric (V) acid → potassium nitrate + water + carbon (IV) oxide

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

(ii) Sodium hydrogencarbonate + hydrochloric acid → sodium chloride + water + carbon (IV) oxide

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

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

(i) $CaCO_3(s)$ and $H_2SO_4(aq)$

(ii) $PbCO_3(s)$ and $H_2SO_4(aq)$

(iii) $PbCO_3(s)$ and $HCl(aq)$

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

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

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

(ii) Lead carbonate + sulphuric (VI) acid → lead (II)sulphate + water + carbon (IV) oxide

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

(iii) Lead carbonate + hydrochloric acid → lead (II) chloride + water + carbon (IV) oxide

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

Reaction of a dilute acid with an alkali

Activity 3.5:To prepare sodium chloride from sodium hydroxide (a soluble base).

In groups

Apparatus and chemicals

- burette
- pipette
- conical flask

- beaker
- funnel
- evaporating dish
- burner
- dilute sodium hydroxide
- dilute hydrochloric acid
- phenolphthalein indicator

Procedure

1. Using a pipette transfer 25cm^3 of the alkali into a conical flask. Then add 2-3 drops of phenolphthalein indicator as shown in Fig. 3.4a. The indicator turns pink. Why is it advisable to use a pipette filler to fill the pipette rather than the mouth?
2. Add the acid from a burette, a little at a time. Then swirl the flask in a controlled way, to allow the acid and alkali to mix see Fig. 3.4b.
3. When all the alkali has been used up, the indicator suddenly turns **colourless,** showing that the solution is neutral. This is the end point. There is no need of adding more acid. Look at Fig. 3.4c.
4. You can tell how much acid was added, using the scale on the burette. Look at Fig. 3.4d. This is the amount of acid that is needed to neutralize 25cm^3 of the alkali.
5. Repeat steps 1 to 4, but this time there is no need for an indicator. 25cm^3 of alkali is put in the flask, and the correct amount of acid added. See Fig. 3.4e.
6. The solution from the flask is heated, to let the water evaporate. See Fig.3.4f. You will find that dry crystals of sodium chloride are left behind.
7. Record observations in your note book. What do you conclude.
    - What do you think would happen if we evaporated the solution obtained in step 3 to dryness?
    - Why was the whole process repeated i.e. steps 1 to 4?
    - **Using the same procedure, describe how you can prepare potassium sulphate. Choose the appropriate reagents.**

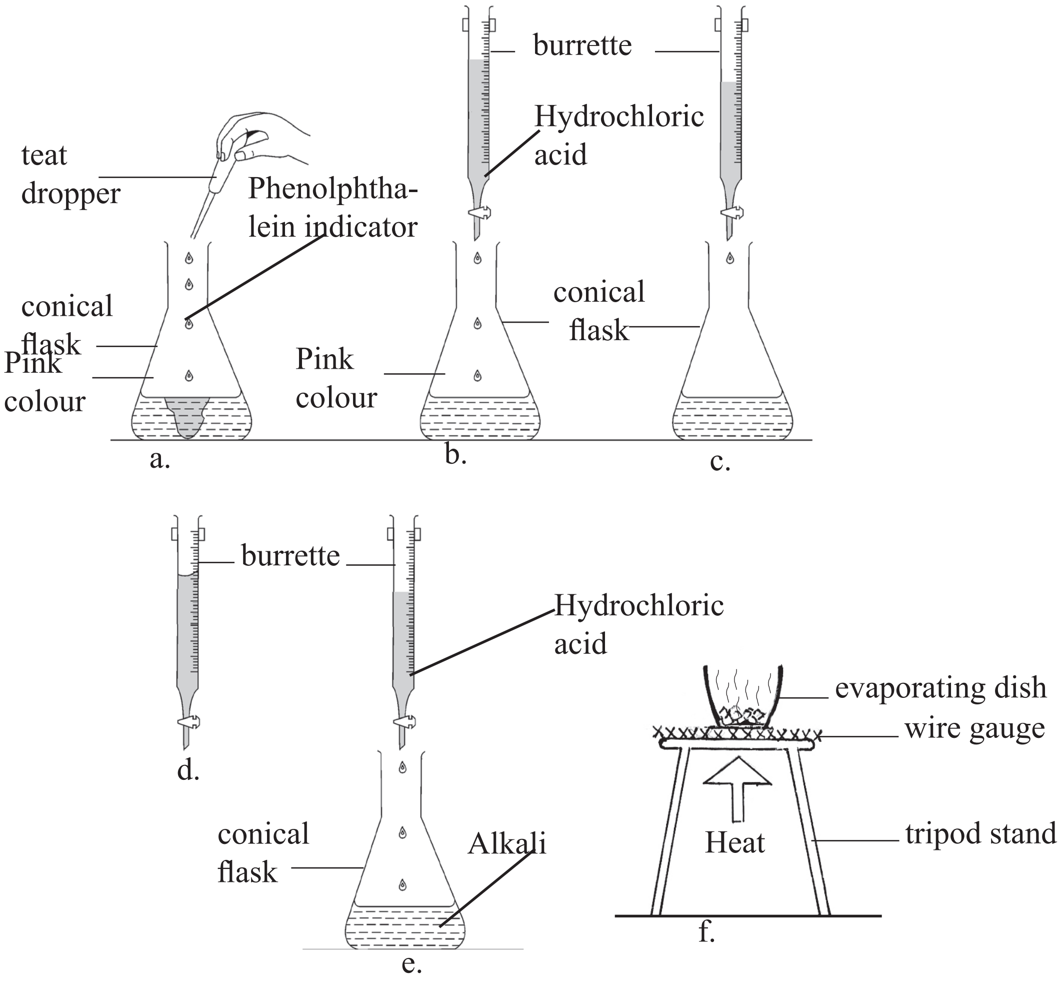


Fig 3.4: Preparation of sodium chloride

The facts

When we want to fill a pipette, it is advisable to use a pipette filler to avoid sucking the liquid into the mouth accidentally. This method of preparing a salt, involves the **neutralisation** of an acid with an alkali, for example sodium hydroxide or a soluble carbonate such as sodium carbonate. It is necessary to use an indicator because both the reactants and products are colourless and it would be difficult to know the neutralisation point or end point. If we evaporate the solution with the indicator, our salt would be coloured.

If we do not want to run the burette twice, we can boil the solution with the coloured indicator for 5 minutes with animal charcoal and then filter off the charcoal. The charcoal absorbs the colour of the indicator and the filtrate is colourless. Then we evaporate to get a colourless salt. This method of preparing a soluble salt using a soluble base is known as **neutralisation**.

Neutralisation is a chemical reaction between a base and an acid to produce a **salt** and **water** only. The reaction that occurred is as follows:

Sodium hydroxide + hydrochloric acid → sodium chloride + water

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

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

Hydrochloric acid + potassium hydroxide → potassium chloride + water

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

Sulphuric acid (VI) + sodium hydroxide → sodium sulphate + water

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

Nitric acid (V) + ammonium hydroxide → ammonium nitrate + water

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

Ionic equation:

Hydrogen ion + hydroxide ion → water

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

Note: (i) Potassium and sodium hydroxide exist in solid form but 'ammonium hydroxide'exists in solution form.

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

Sodium carbonate + hydrochloric acid → sodium chloride + water + carbon(IV) oxide

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

Ammonium carbonate + nitric acid (V) → Ammonium nitrate + water + carbon(IV) oxide

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

(b) Preparation of insoluble salts

(i) Preparation of insoluble salts by precipitation

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

Activity 3.6: To prepare lead (II) sulphate

In groups

Apparatus and chemicals

- Beaker or test tube
- Lead nitrate solution
- Sulphuric acid (or any soluble sulphate e.g. sodium sulphate)

Procedure

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

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

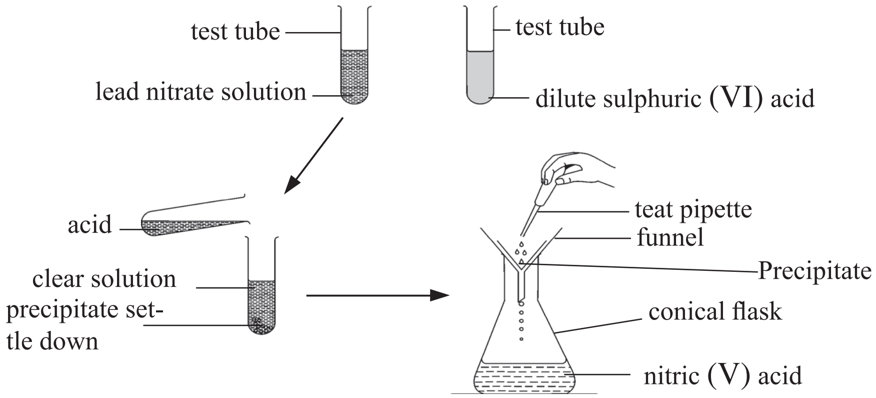


Fig. 3.5: Preparation of lead (II) *sulphate*

The facts

When dilute sulphuric (VI) acid or any other soluble sulphate is added to the lead nitrate solution, a white insoluble salt is formed. This insoluble product is also called a **precipitate**. This method of preparing salts is called **precipitation** or **double decomposition**.

Note: a precipitate is a solid that forms out of a solution. In this method, we start with solutions of substances; therefore both must be soluble in water. When we mix the two solutions, they exchange ions as shown in the following equation.

Lead (II) nitrate + dilute sulpuric (VI) acid → lead (II)sulphate + hydrogen nitrate

(solution) (solution) (precipitate) (remains in solution)

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

Note: Sulphuric acid is called hydrogen sulphate and nitric acid is called hydrogen nitrate.

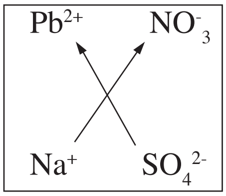
This is to show how hydrogen ions are exchanged. A solution of any sulphate could have been used instead of sulphuric acid for this preparation; for example sodium sulphate.

Lead (II)nitrate + sodium sulphate → lead (II) sulphate + sodium nitrate

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

The ions present in the solutions are

From $Pb(NO_3)_2(aq) \rightarrow$



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

From $Na_2SO_4(aq) \rightarrow$

Ionic equation:

When we write an ionic equation, we select the ions that form what we see; in this case the **precipitate**. We leave out the ions that remain in solution. These are known as **spectator ions**. Just as in a match, a spectator does not participate in the game.

For instance, in the reaction between lead (II) nitrate and sodium sulphate, the precipitate is lead (II) sulphate. The ionic equation for this reaction would be written as:

Lead (II) ions + sulphate ions → lead (II) sulphate

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

Other insoluble salts which can be prepared by double decomposition are:barium sulphate and lead (II) chloride. Lead (II) chloride is a precipitate that is soluble in hot water.

Barium chloride + sodium sulphate → barium sulphate + sodium chloride

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

Barium ion + sulphate ion → barium sulphate

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

Lead (II) nitrate + sodium chloride → lead (II) chloride + sodium nitrate

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

Lead (II) ion + chloride ion → lead (II) chloride

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

(soluble in hot water)

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

Hydrogen ion + hydroxide ion → water

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

Activity 3.7

1. Repeat Experiment 3.5 with the following pairs of aqueous solutions in Table 3.2.

Table 3.2 Formation of some coloured precipitates

Solution of salt X	Solution of substance Y	Record colour of precipitate if any
Lead (II) nitrate	Potassium iodide	
Copper (II) sulphate	Sodium hydroxide	
Iron(II) sulphate	Sodium hydroxide	
Iron(II) chloride	Sodium hydroxide	
Sodium chloride	Silver nitrate	
Lead nitrate	Ammonium chloride	
Ammonia solution	Zinc nitrate	

2. Note the colour of the precipitates and write balanced equations in each case followed by ionic equations.
3. Record your observations in the third column in the table. Write equations of the reactions and account for the colours of the precipitates.

3.2 Writing ionic equations involving salts

Reactions that take place between ionic compounds dissolved in water can be recorded in a much simpler manner. The ions are free to move and react. When a chemical reaction takes place between ionic compounds, it may happen that only certain ions undergo changes while others do not. You will remember from Unit 2 that ionic compounds are made of a lattice of ions that are rigidly fixed in solid form. When ionic compounds are dissolved in water, the ions are set free.

Knowledge of solubility of salts in water is very crucial when writing ionic equations. This assists very much so that correct symbols can be identified.

The following is a summary of solubility rules involving salts and bases.

- All potassium (K), sodium (Na) and ammonium (NH_4) salts are soluble.
- All nitrates and hydrogencarbonates (HCO_3) are soluble.

- All sulphates are soluble except lead (II) sulphate ($PbSO_4$) and barium sulphate ($BaSO_4$). Calcium sulphate ($CaSO_4$) is slightly soluble.
- All chlorides are soluble except silver chloride (AgCl). Lead (II) chloride ($PbCl_2$) is only soluble in hot water.
- All carbonate salts are insoluble except potassium, sodium and ammonium carbonates.

When sodium carbonate solution is mixed with barium chloride solution, a reaction takes place and a white precipitate is formed. Fig 3.6

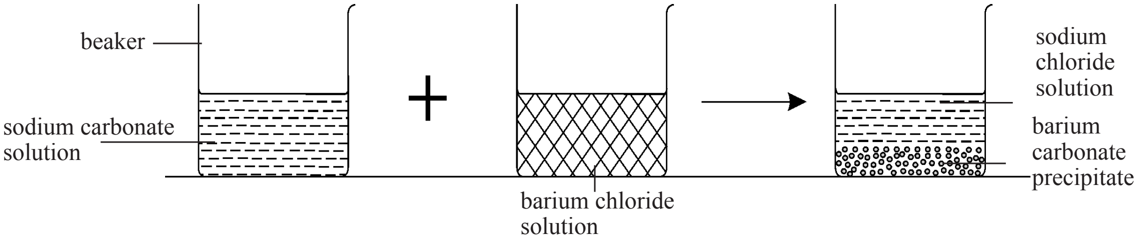


Fig. 3.6: Precipitation reaction

The balanced chemical equation for this reaction is as follows.

Sodium carbonate + barium chloride → sodium chloride + barium carbonate

$$Na_2CO_3(aq) \quad + \quad BaCl_2\,(aq) \quad \rightarrow \quad 2NaCl(aq) \quad + \quad BaCO_3(s)\text{ (white)}$$

Before the solutions were mixed, the ions present were:

From sodium carbonate solution ($Na_2CO_3(aq)$); $2Na^+$ and CO_3^{2-}

From barium chloride solution ($BaCl_2(aq)$); Ba^{2+} and $2Cl^-$

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

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

There are three possibilities during ionic changes:

- Formation of a **precipitate.**
- Evolution of a **gas.**
- Formation of **water.**

Table 3.3 Shows some substances that contain free ions and those that do not.

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

Note: Solids, liquids, gases do not have free ions.

An ionic equation can be written following certain rules.

Worked Examples

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

Step 1

Write down well balanced chemical equation for the reaction

Sodium carbonate + barium chloride → sodium chloride + barium carbonate

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

Step 2

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

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

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

Step 3

Rearrange the ions and rewrite the equation in ionic form.

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

Work to do

Write the ionic equation of iron(III) chloride with sodium hydroxide.

Solution

(a) Write down a well balanced stoichiometric or the normal chemical equation for the reaction

(b) Iron(III) chloride + sodium hydroxide → sodium chloride + iron(III) hydroxide

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

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

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

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

$$Fe^{3+}(aq) + \cancel{3Cl^-(aq)} + \cancel{3Na^+(aq)} + 3OH^-(aq)^{2-} \rightarrow \cancel{3Na+}(aq) + \cancel{3Cl^-}(aq) + Fe(OH)_3(s)$$

(c) Rearrange the ions and rewrite the equation in ionic form.

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

ii) Preparation of insoluble salts by direct combination of elements

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

Activity 3.8

Preparation of sodium chloride by direct combination of sodium and chlorine

Work in groups

Apparatus and chemicals

- deflagrating spoon
- knife/scalpel
- bunsen burner
- gas jar full of dry chlorine
- tongs
- sodium metal

Procedure

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

The facts

The reaction is violent. A white powder of sodium chloride is formed.

Sodium + chlorine → sodium chloride

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

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

Iron + chlorine → iron(III) chloride

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

Double salts

Some salts exist as double salts, for example **trona** which occurs naturally in lake Magadi in Kenya. Trona contains both sodium carbonate, (Na_2CO_3) and sodium hydrogencarbonate, ($NaHCO_3$). It is a hydrated salt with the following formula.

$$Na_2CO_3.NaHCO_3.2H_2O$$

Activity 3.9

To find out what happens when certain salts are left exposed in the air.

Apparatus and chemicals

- watch glasses/white tile
- anhydrous copper (II) sulphate
- anhydrous calcium chloride
- iron(III) chloride
- sodium nitrate
- hydrated sodium carbonate
- hydrated sodium sulphate

Procedure

1. Place a small amount of each of the above salts on separate watch glasses.
2. Observe and record the appearance of each salt in table 3.4 below.
3. Leave the samples exposed in air for about 3-7 days.
4. Observe and record any changes that may occur after a few minutes, one day, one week or more.
5. Suggest a reason for the change in appearance, if any.

Table 3.4: Observations made when salts are exposed in air

Salt	Appearance before staying overnight	Appearance after staying overnight	Reason for change
(a) Anhydrous copper (II) sulphate			
(b) Anhydrous calcium chloride			
(c) Iron(III) chloride			
(d) Hydrated sodium carbonate			

The facts

From your results you will notice that some changes have occurred in the salts. Some salts have absorbed water from the air while others have lost water into the air. Many substances absorb water from the air. For example ordinary impure sodium chloride (common salt) becomes damp when exposed to air. Impure salt contains magnesium chloride which absorbs water from the air. Pure salt does not absorb water. White anhydrous copper (II) sulphate turns blue when exposed to air.

Substances that absorb water from the air but do not form a solution are said to be **hygroscopic**. Examples of hygroscopic substances are listed in Table 3.5. Some salts e.g. calcium chloride absorb so much water from the air that it dissolves in it. Substances like this are said to be **deliquescent**. The process is known as **deliquescence**.

Substances that lose water of crystallisation to the air are said to **efflorescent**. The process is called **efflorescence**. For example, a crystal of hydrated sodium carbonate, when left exposed to air, slowly forms a white powder on its surface and finally the crystal changes to powder. In other words it loses 9 of its 10 water molecules of crystallisation to the air as shown in the following equation.

Sodium carbonate-10-water → Sodium carbonate-1-water + 9-water

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

Table 3.5:Summary of types of salts

Hygroscopic salts	Deliquescent salts	Efflorescent salts
Anhydrous copper (II) sulphate	Sodium nitrate	Sodium carbonate-10-water
	Calcium chloride Magnesium chloride	Sodium sulphate-10-water
	Iron(III) chloride	Iron(II) sulphate-7-water

Determining solubility of salts

When we dissolve substances such as sugar and common salt in water, we cannot see them in the colourless liquid. But we know they are there because we can taste the sugar or the salt in water. The solid has been broken into tiny particles which cannot be seen with the naked eye, nor with the most powerful microscope. The particles will pass through the pores of a filter paper. These particles and the water form a homogeneous mixture called a **solution**.

Substances such as salt and sugar which dissolve in water are described as **soluble**.

Solids such as sand and soil which do not dissolve are said to be **insoluble**. We say that a solution which contains only a little solute in a given amount of solvent is **dilute**. But if a solution contains a lot of solute in a given amount of solvent we say it is **concentrated**. If we evaporated more solvent, a solution becomes more concentrated. Eventually, the solution reaches a point that it cannot become any more concentrated while remaining a solution. If any more solvent is evaporated, some solid solute will come out of solution and form crystals. This is because the solution is **saturated**.

Activity 3.10: To make a saturated solution

In groups

Apparatus and chemicals

- beaker
- Bunsen burner
- stirrer/glass rod
- hydrated copper (II) sulphate or potassium nitrate or sodium chloride.

Using the knowledge from your previous lesson in secondary 1, write down the procedure on how to prepare a saturated solution of copper (II) sulphate.

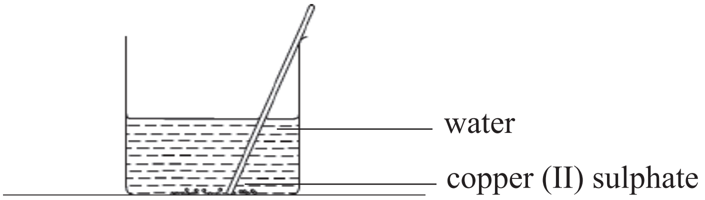


Fig 3.7: Making a saturated solution of copper (II) sulphate

The facts

The first portion dissolves and a faint blue solution is formed. As more and more portions are added and dissolved, the blue colour deepens. As further portions are added, it gets to a point when no more copper (II) sulphate will dissolve

This means that the solution has become **saturated.** The only way we can tell a solution is saturated, is when we see some undissolved solute. If the solution is warmed the undissolved solid dissolves. The hot solution dissolves more solute. This indicates that the solubility of solid increases with rise in temperature. A *saturated solution* is therefore a solution which contains as much solute as can be dissolved at a **particular temperature,** and has some **undissolved solute**.

Activity 3.11 To investigate solubility of some salts

Apparatus and chemicals

- test tube rack
- test tubes
- potassium nitrate
- sodium carbonate
- barium sulphate
- zinc carbonate
- copper carbonate
- copper (II) hydroxide
- sodium hydroxide
- barium chloride
- calcium carbonate
- potassium nitrate
- sodium chloride
- sodium nitrate
- ammonium sulphate
- lead (II)nitrate
- copper (II) nitrate
- zinc nitrate
- magnesium sulphate
- copper (II) sulphate
- silver nitrate
- silver chloride
- lead (II) carbonate

Procedure

1. Take a spatula endful of each of the salts listed above and put it in a test tube.
2. Add cold water to about half the test tube. (Do not warm).
3. Close the test tube with your thumb and shake well. What do you observe?
4. Place the test tube on the rack and observe what happens after a shortwhile.
5. Record observations in Table 3.6. What do you conclude?

Copy Table 3.6 in your notebook and put a tick(√) where salt is soluble and a cross (x) where the salt is insoluble.

Table 3.6 Some soluble and insoluble salts

Name of salt	**Soluble**	**Insoluble**
Potassium chloride Sodium nitrate Ammonium sulphate Sodium carbonate		
Lead (II) nitrate Copper (II)nitrate Zinc (II) nitrate Magnesium sulphate Calcium sulphate Barium sulphate Silver chloride Barium chloride Calcium carbonate Lead carbonate Zinc carbonate		

The facts

From your results you can see that some of the salts are soluble in water whereas others are insoluble. Table 3.7 shows the solubility patterns for various types of salts.

We must learn and remember the soluble and insoluble salts because this will assist us in indicating the state symbols when writing equations especially during precipitation. As mentioned earlier, precipitation is a method used to prepare insoluble salts from solutions of two soluble salts. For example:-

Sodium sulphate + lead (II) nitrate → lead (II) sulphate + sodium nitrate

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

If you look on the products side, lead sulphate, $PbSO_4(s)$ is given the solid state because it is among the insoluble sulphates. Sodium nitrate, $NaNO_3(aq)$ is given an aqueous state (aq) because all nitrates and sodium salts are soluble. We do not prepare nitrates and other soluble salts by precipitation.

Table 3.7 Solubility pattern of various types of salts

Salt	**Soluble**	**Insoluble**
Potasium salts	All are soluble	None is insoluble
Sodium salts	All are soluble	None is insoluble
Ammonium salts	All are soluble	None is insoluble
Chlorides	All are soluble hot water)	Except silver chloride, Lead (II) chloride (soluble in
Nitrates	All are soluble	None is insoluble
Carbonates	Only K, Na and NH are soluble	Others are insoluble
Sulphates	All are soluble	Except lead sulphate, barium sulphate, but calcium sulphate is slightly soluble

3.3 Action of heat on salts

As we have seen salts are compounds which contain metal or ammonium ions which, are positively charged (cations) and acid radicals which are negatively charged (anions). These opposite charges attract each other strongly. Strong electrostatic forces of attraction hold the ions together. In some salts, the forces of attraction are so strong that application of heat cannot break them. However, in some cases strong heating decomposes some salts and we get new products.

(a) Action of heat on carbonates and hydrogencarbonates

Carbonates and hydrogencarbonates are salts of carbonic acid, H_2CO_3. There are two solid hydrogencarbonates, namely, potassium hydrogencarbonate ($KHCO_3$) and sodium hydrogencarbonate ($NaHCO_3$).

The other known hydrogencarbonates exist only in solution. These are:

- calcium hydrogencarbonate, $Ca(HCO_3)_2$.
- magnesium hydrogencarbonate, $Mg(HCO_3)_2$

All metals form solid metal carbonates. However, aluminium and iron(III) carbonates do not exist.

The effect of heat on metal compounds can be linked to the reactivity series like the reaction of metals themselves.

Activity 3.12 To investigate the action of heat on carbonates and hydrogencarbonate.

In groups

Apparatus and chemicals

- Bunsen burner
- spatula
- test tube holder
- hard glass test tubes
- carbonates and hydrogencarbonates (as listed in the procedure)
- calcium hydroxide solution (lime water)

Procedure

1. Put into different dry test tubes about one spatula end of each of the following salts: Potassium carbonate, sodium carbonate, magnesium carbonate, zinc carbonate, copper carbonate, iron(II) carbonate, lead carbonate, potassium hydrogencarbonate and sodium hydrogencarbonate.
2. Note the colour of the salt before heating.
3. Heat the sample and test the gas produced with calcium hydroxide solution as in Fig 3.8. Copy Table 3.8 and record your observation.
4. Finally note the colour of the residue when hot and when it has cooled.
5. List the compounds which bring a change in calcium hydroxide solution when they are heated. These compounds will have decomposed.

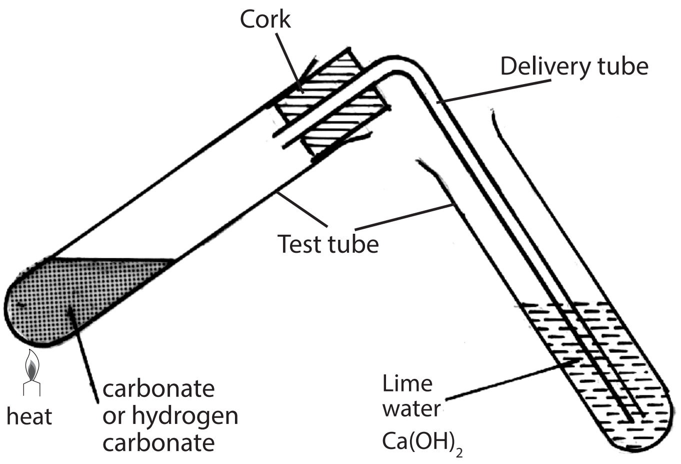


Fig 3.8: Testing decomposition of a carbonate/hydrogencarbonate on heating

The compounds that are coloured are copper carbonate and iron(II) carbonate. They are green solids. All the others are white solids.

Table 3.8 Decompositon of carbonates and hydrogencarbonates

Carbonate/hydrogen carbonate	**Colour before heating**	**Colour of the residue**	**Effect on calcium hydroxide solution**
Potassium carbonate			
Sodium carbonate			
Zinc carbonate			
Lead (II) carbonate			

The facts

From the results we recorded in Table 3.8, we notice that some carbonates are not decomposed by heat (there is no effect on calcium hydroxide solution. Potassium carbonate and sodium carbonate are not decomposed by heat).

Other metal carbonates decompose on heating to their respective metal oxides and carbon (IV) oxide.

Metal carbonate $\xrightarrow{\text{Heat}}$ metal oxide + carbon (IV) oxide

The carbon (IV) oxide produced reacts with calcium hydroxide solution forming a white precipitate of calcium carbonate as follows:

Calcium hydroxide + carbon (IV) oxide → calcium carbonate + water

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

The tiny particles of the precipitate are suspended in the liquid.

Generally:

Examples

Zinc carbonate decomposes on heating to form a yellow solid that turns white on cooling and a colourless gas is given off.

Zinc carbonate → zinc oxide + carbon (IV) oxide

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

(white) (yellow when hot

white when cold)

Lead (II) carbonate decomposes on heating to form a brown solid that turns yellow on cooling and colourless gas.

Lead carbonate → lead (II) oxide + carbon (IV) oxide

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

(white) (red-brown hot

yellow cold)

Black solid is formed when copper (II) carbonate is heated.

Copper carbonate → copper (II) oxide + carbon (IV) oxide

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

(green) (black)

White solid is formed when magnesium carbonate is heated.

Magnesium carbonate → magnesium oxide + carbon (IV) oxide

$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

(white) (white)

Note: Three products are formed when hydrogen carbonates are decomposed by heat.

For instance, potassium hydrogen carbonate and sodium hydrogen carbonate decompose to **metal carbonate, water** and **carbon(IV) oxide**.

Potasium hydrogen carbonate → potasium carbonate + water + carbon (IV) oxide

$2KHCO_3(s) \rightarrow K_2CO_3(s)$(white) $+ H_2O(g) + CO_2(g)$

Sodium hydrogen carbonate → sodium carbonate + water + carbon (IV) oxide

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

(white) (white)

However, ammonium carbonate when heated decomposes into two gases namely ammonia, carbon (IV) oxide and water vapour. The water vapour condenses on the cold part of the test tube.

It can be tested with anhydrous cobalt (II) chloride paper or anhydrous copper (II) sulphate.

Ammonium carbonate → ammonia + carbon (IV) oxide + water

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

Note: We mentioned earlier that sodium carbonate is not decomposed by heat.

The hydrated sodium carbonate when heated loses water of crystallisation as shown in the following equation.

Hydrated sodium carbonate → sodium carbonate + water

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

The reactivity series helps to explain the action of heat on carbonates. Potassium and sodium are very reactive metals (top of the reactivity series). Hence as mentioned earlier their carbonates do not decompose on heating. Calcium and magnesium follow in the series. Their carbonates decompose but at a higher temperature. The other carbonates lower in the series decompose easily. For example copper metal is near the bottom of the series; copper carbonate decomposes easily on warming.

Activity 3.13: To investigate the effect of heat on nitrates

In groups

Apparatus and chemicals

- ignition tubes
- wood splint
- sodium nitrate
- calcium nitrate
- lead (II) nitrate
- burner
- test tube holder
- potassium nitrate
- zinc(II)nitrate
- magnesium nitrate

anhydrous copper (II) sulphate or dry cobalt chloride paper

Procedure

1. Place a spatula end-full of sodium nitrate into an ignition tube.
2. Note the colour of the nitrate and heat the sample strongly.
3. Insert a glowing splint into the test tube and note what happens.
4. Copy Table 3.9 in your notebook and record your observations.

5. If any liquid condenses in the test tube, test with anhydrous copper (II) sulphate or dry cobalt chloride paper.
6. Repeat steps 1 - 4 using the rest of the nitrates listed above.

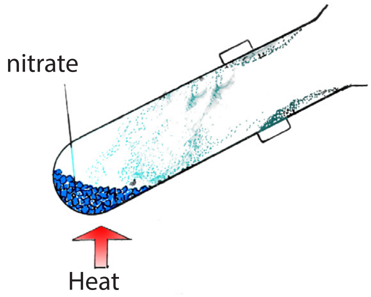


Fig 3.9: Heating nitrates

Table 3.9 Action of heat on nitrates

Name of nitrate used	Colour before	Change of colour during heating (if any)	Colour of gas heating	Effect on glowing splint produced	Effect on dry cobalt chloride paper
Sodium nitrate					
Potassium nitrate					
Calcium nitrate					
Zinc (II) nitrate					
Lead (II) nitrate					
Magnesium nitrate					

Study questions

1. List the nitrates which behaved the same in this experiment.
2. Which nitrates evolved a gas that relights a glowing splint?
3. Which nitrates produced residues that showed colour changes?
4. Which nitrates were easily decomposed?
5. Which nitrate makes a lot of noise when heated?

The facts

Results from the experiments can be used to classify the nitrates according to products formed when they are heated.

1. Nitrates that form two products only

- Sodium nitrate
- Potassium nitrate
- Ammonium nitrate

Sodium nitrate and potassium nitrate melt to form colourless liquid. On further heating, they slowly decompose forming metal nitrite and oxygen.

Sodium nitrate → sodium nitrite + oxygen

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

Potassium nitrate → potassium nitrite + oxygen

$$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$$

Note: Ammonium nitrate melts to a colourless liquid, which decomposes to form nitrogen (I) oxide and water.

Ammonium nitrate → nitrogen (I) oxide + Water

Caution: Do not heat ammonium nitrate in the laboratory because when a little of the nitrate is left, it explodes on strong heating.

2. Nitrates that decompose to form three products

These nitrates decompose to the metal oxide, nitrogen(di) oxide, (red-brown gas) and oxygen.

- Calcium nitrate
- Magnesium nitrate
- Zinc (II) nitrate
- Copper (II) nitrate
- Lead (II) nitrate

Note: • All the above nitrates are hydrated except lead nitrate.

• All the crystals are white except copper nitrate.

Examples

Calcium nitrate → Calcium oxide + nitrogen (IV) oxide + oxygen

$2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$

(white)

Magnesium nitrate → magnesium oxide + nitrogen (IV) oxide + oxygen

$2Mg(NO_3)_2(s) \rightarrow 2MgO(s) + 4NO_2(g) + O_2(g)$

(white)

Zinc nitrate → zinc oxide + nitrogen (IV) oxide + oxygen

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

(white) (yellow when

hot; white when cold)

Copper nitrate → Copper (II) oxide + Nitrogen (IV) oxide + Oxygen

$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$

(blue) (black)

When lead nitrate is heated, reddish brown gas forms inside the crystals and splits them. This **crackling** noise is called **decrepitation**. The reddish brown solid formed turns yellow when cold. The oxygen produced relights a glowing splint.

Lead nitrate → Lead (II) oxide + Nitrogen (IV) oxide + Oxygen

$2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$

(reddish brown) (red brown)

Note: We have mentioned that lead nitrate is formed in anhydrous form. The other nitrates are hydrated. When we heat the hydrates we drive off water and on further heating they decompose just like lead nitrate.

e.g. hydrated metal nitrate → metal nitrate + water

$$Cu(NO_3)_2.3H_2O(s) \rightarrow Cu(NO_3)_2(s) + 3H_2O(l)$$

Silver nitrate and mercury nitrate

Silver nitrate and mercury nitrate form metal oxides, nitrogen(IV) oxide and oxygen just like the other heavy metal nitrates. But silver oxide, and mercury oxide when heated usually decompose to free metal and oxygen. So the oxide formed during decomposition of the nitrate will decompose to the metal and oxygen.

$$\text{Silver nitrate} \xrightarrow{\text{heat}} \text{silver} + \text{nitrogen(IV) oxide} + \text{oxygen}$$

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

$$\text{Mercury nitrate} \xrightarrow{\text{heat}} \text{mercury} + \text{nitrogen(IV) oxide} + \text{oxygen}$$

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

Therefore, nitrates of metals high in the reactivity series decompose on heating to give corresponding nitrates and oxygen.

Moderately reactive metal nitrates give corresponding metal oxides, nitrogen (IV) oxide and oxygen on heating.

Least reactive metals give corresponding metal, nitrogen (IV) oxide and oxygen.

The action of heat on nitrates vary according to the position of the metal in the reactivity series.

$\left.\begin{array}{l}\text{K}\\ \text{Na}\end{array}\right\}$ On heating nitrates of these metals we get the nitrite and oxygen.

Ca
Mg
Al ⎬ On heating nitrates of these metals we get the metal oxide, nitrogen(IV) oxide and oxygen.
Fe
Pb

Cu
Hg ⎬ Nitrates of these metals decompose to the metal, nitrogen(IV) oxide and oxygen.
Ag

Action of heat on sulphates

Activity 3.14

To investigate the effect of heat on sulphates

Apparatus and chemicals

- Ignition tubes
- Litmus papers
- Hydrated iron(II) sulphate
- Bunsen burner
- Test tube holder
- Hydrated copper (II) sulphate

Procedure

1. Place a little of the sulphate at a time in the ignition tube.
2. Heat gently first and record your observation.
3. Then heat strongly and test any gases produced with litmus papers.

   Record your observations.

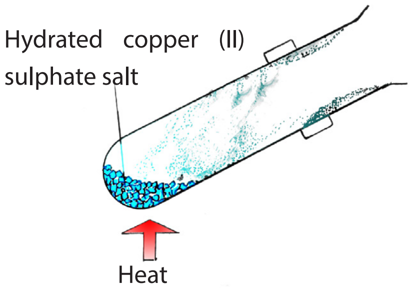


Fig 3.10: Heating nitrates

4. Record the initial and final colour of the residue.
5. Repeat the experiment with the other sulphates.

The facts

Hydrated salts first lose their water of crystalisation. Then the anhydrous salt decomposes on very strong heating.

For example

Iron(II) sulphate-7-water → iron(II) sulphate + water

$$FeSO_4.7H_2O(s) \rightarrow FeSO_4(s) \text{ (white powder) } + 7H_2O(g)$$

Then:

Anhydrous iron(II) sulphate → iron (III) oxide + sulphur (IV) oxide oxide + sulphur(VI) oxide

$$2FeSO_4(s) \rightarrow \underset{\text{(brown powder)}}{Fe_2O_3(s)} + SO_2(g) + SO_3(g)$$

Blue hydrated copper (II) sulphate will also first lose water of crystallisation. It will form white anhydrous copper (II) sulphate.

Copper (II) sulphate-5-water → anhydrous copper (II) sulphate + 5-water

$$\underset{\text{(blue)}}{CuSO_4.5H_2O(s)} \rightarrow \underset{\text{(white)}}{CuSO_4(s)} + 5H_2O(g)$$

On further heating, anhydrous copper (II) sulphate decomposes to copper (II) oxide and sulphur(VI) oxide.

Anhydrous copper (II) sulphate → copper (II) oxide + sulphur (VI) oxide

$$\underset{\text{(white)}}{CuSO_4(s)} \rightarrow \underset{\text{(black)}}{CuO(s)} + SO_3(g)$$

Uses of salts

Activity 3.15: Research Activity

In pairs

Propose some common uses of salts in our daily life. Write a report and present to the class.

The facts

1. *In controlling soil pH and plant growth.*

   Plant growth is affected by the alkalinity or acidity of the soil.
   The soil pH is affected by the use of fertilisers and the acid rain.

   If the soil is too acidic, the effect can be neutralised by calcium carbonate or calcium hydroxide. Calcium oxide is also used to neutralise soil acidity.

2. *Fertilisers*

   The following salts are used as fertilisers.

   - Ammonium nitrate
   - Ammonium phosphate
   - Ammonium sulphate
   - Sodium nitrate

3. *Indigestion* can be relieved by salts.

   Examples are calcium carbonate and sodium hydrogencarbonate.

4. *Baking*

   Sodium hydrogencarbonate is used in baking. When heated it produces carbon (IV) oxide which causes bread to rise.

5. *Softening of hard water*

   Sodium carbonate is used to soften hard water.

6. *Melting ice on roads*

   Sodium chloride and other salts are poured on roads to lower the freezing point of ice in countries that experience winter.

7. *Extraction of metals*

   Calcium chloride is used in the extraction of sodium metal.

Check your progress 3.1

1. What is a salt?

2. Briefly explain the difference between the following:

    (a) acid salts and normal salts (b) hydrated salts and anhydrous salts.

    (c) deliquescent, efflorescent salts and hygroscopic substances.

3. What do you understand by the term double decomposition.

4. What is decrepitation?

5. (a) A student wanted to prepare lead (II)sulphate. He chose dilute sulphuric (VI) oxide acid and lead (II) carbonate. He put the acid in a beaker and added lead (II) carbonate. But after a short while the reaction stopped.
    Explain why the reaction stopped?

    (b) What is the appropriate method he can use to prepare lead (II) sulphate.

6. Explain how you would prepare normal salts in the laboratory illustrate each method by writing an equation for the reaction.

7. Describe in detail how you would prepare in the,laboratory, crystals of zinc sulphate from zinc.

8. The chloride salts can be prepared by the following methods:

    (a) Direct combination (b) Precipitation (c) Neutralisation

9. Select suitable, but different methods and briefly state how you can prepare each of the following salts.

    (i) Sodium chloride (ii) Silver chloride

    (iii) Lead (II) chloride (iv) Iron (III) chloride

10. Akot was heating some salts in the laboratory.On testing the gases produced using a glowing splint one did not relight the glowing splint whereas the other one relight the splint.Write down the decomposition reaction of the salts that were being heated

11. Copy the following table and put a tick (√) if a precipitate is formed when the solutions are mixed and a cross (x) if no precipitate is formed.

Sodium sulphate	Silver nitrate	Sodium sulphate
Sodium chloride		
Barium chloride		
Lead nitrate		
Calcium chloride		
Magnesium sulphate		

12. Name uses of salts?

3.4 Electrolysis

Activity 3.16: To investigate the conductivity of aqueous solutions

In pairs

Apparatus and chemicals

- battery/dry cells
- connecting wires
- bulb (6V) or ammeter
- beaker 250ml
- graphite rods
- distilled (pure) water
- acidified water
- aqueous solutions of: sugar, urea, copper (II) chloride, sodium chloride, mineral acids (hydrochloric, nitric and sulphuric acids).

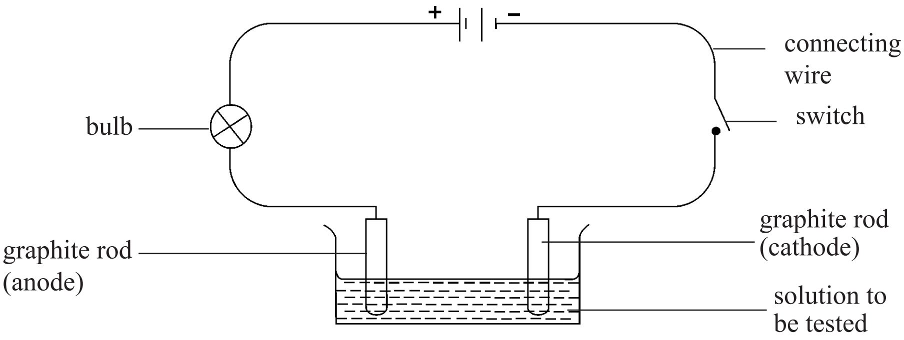


Fig. 3.11: The apparatus for testing the electrical conductivity of an aqueous solution

Procedure

1. Arrange the apparatus as in Fig. 3.11.
2. Switch on the current.
3. Before dipping the graphite rods into the solution, let the two electrodes come into contact and check whether the bulb lights up.
4. Place the graphite rods apart in the beaker.
5. Switch off the current.
6. Pour pure water into the beaker.Predict what will happen.Does the bulb light or not?
7. Switch on the current. Does the bulb glow? Compare the result to your prediction.
8. Switch off the current.Add a few drops of concentrated sulphuric acid to the water in the beaker.
9. Switch on the current.Does the bulb light?
10 What do you observe on the graphite rods? What does this indicate?
11. Do the above experiment using graphite rods and the solutions of the following compounds: sugar, urea, copper (II) Chloride, sodium chloride,mineral acids.
12. Note the changes (if any) occurring around the graphite rods and in the contents of the beaker.Rise the beaker and rods thoroughly with distilled water before each new solution is used. Why do you need to rinse the beaker before carrying out another test?
13. Copytable 3.10 and fill your results, include all the other substances being investigated.

Table 3.10 Electrolytes and non-electrolytes

Solution/substance	Does the bulb light		Any observation made on the anode (+ve), cathode (–ve)	
	Yes	No	Anode (+ve)	Cathode (-ve)
Sugar				
Sodium chloride				
Mineral acids				
Sodium chloride				
Copper (II) chloride				
Pure water				
Acidified water				

The facts

Liquids such as acidified water, sodium chloride solution, mineral acids and copper (II) chloride solution conduct electricity. Molten or aqueous solutions which conduct electric current and decomposed by it are called **electrolytes**. The current decomposes the electrolyte as it passes through it. On the other hand, Aqueous solutions such as sugar and urea do not conduct electricity. They are known as **non-electrolytes**.

What is electrolysis

Electrolysis is the decomposition of a compound by passing an electric current through it. The set-up of apparatus shown in Fig 3.12 is known as an **electrolytic cell**. The graphite rods which carry the current into and out of the electrolytes are called **electrodes.** Graphite is chosen because it does not react with the electrolyte or with the products formed during electrolysis. Unreactive metal electrodes such as platinum and copper can also be used as electrodes. The electrode connected to the positive terminal of the battery is called **anode**. The electrode connected to the negative terminal is called **cathode.**

Activity 3.17: To investigate the conductivity of molten substances

In groups

Apparatus and chemicals

- crucible
- connecting wire fitted with crocodile clips
- bunsen burner
- pipe-clay triangle
- sulphur
- graphite rods
- battery/dry cells with cell holder
- tripod stand
- lead (II) bromide or lead (II) iodide
- paraffin wax

Procedure

Note: Lead (II) bromide or lead (II) iodide produce poisonous fumes. The experiment should be performed in a fume cupboard or in an open space, like a field.

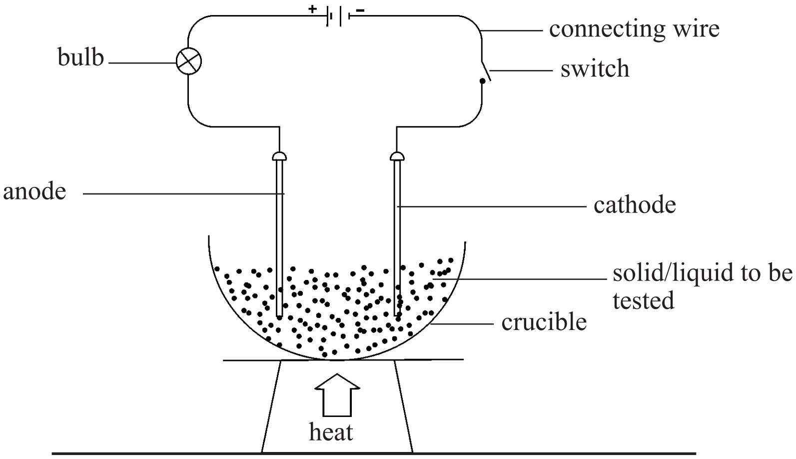


Fig. 3.12: Conductivity of molten substance

1. Set up the apparatus as in Fig 3.12.
2. Fill the crucible with either lead (II) bromide or lead (II) iodide to approximately two thirds.
3. Place the crucible and its contents on a pipe clay triangle supported by a tripod stand.
4. Insert the graphite electrodes into lead (II) bromide and close the switch. Does the bulb light?
5. Open the switch.
6. Gently heat the crucible until lead (II) bromide melts.
7. Adjust the burner so that the substance remains in molten state as the experiment progresses. Close the switch.Does the bulb light?
8. Observe whether there is any chemical reaction around the electrodes.
9. Record your observations as shown in Table 3.11
10. Repeat the experiment with other substances such as paraffin, wax naphthalene, calcium chloride, sugar. Use a fresh crucible for each substance and clean the electrodes thoroughly before using it with the next substance.

Table 3.11 Electrolysis of molten substances

Substance	Does bulb light?		Observations at the electrodes	
	Solid	Molten	Anode	Cathode
Lead (II) bromide	No	Yes	brown fumes of bromine	Grey bead of lead metal
Sulphur				
Wax				
Lead (II) iodide etc.				

The facts

When we insert the electrodes in solid lead (II) bromide, the bulb does not light. But the bulb lights when lead (II) bromide is heated until it melts. This indicates that molten lead (II) bromide conducts electricity. It is an electrolyte. Metal compounds in molten state are electrolytes. It is also observed that other substances like sulphur, paraffin wax do not conduct electricity when in solid or molten form. These are non-electrolytes.

Other examples of electrolytes and non-electrolytes are given in Table 3.12.

Table 3.12 Other examples of electrolytes and non-electrolytes

Electrolytes	Non-electrolytes
Sodium chloride solution	Pure water
Copper (II) chloride solution	Sugar solution
Sodium hydroxide solution	Paraffin wax
Molten lead (II) bromide	Molten sulphur
Molten lead (II) iodide	Solid lead (II)
bromide	
Hydrochloric acid	Solid sodium chloride
Sulphuric acid	Ethanol (alcohol)
Ethanoic acid	Urea
Most salts in molten or aqueous form	

Why electrolytes conduct electricity

In activites 3.15 and 3.16 we tested the ability of a range of liquid compounds and solutions to conduct an electric current. We have seen that some compounds when dissolved in water or when in molten state are electrolytes. These include sodium chloride and lead (II) bromide and most other salts. Compounds such as sugar do not conduct electricity when in solution or molten form.

Electrolytes are ionic compounds. They contain charged particles called *ions*. In solid form the ions are not free to move, therefore a solid ionic compound does not conduct electricity. The ions move if the solid is **melted** or **dissolved in water.**

Study Figs. 3.13 and 3.14 (a & b). The ions carry the charge through the liquid and the current flows.

The positively charged ions are called cations and the negatively charged ions are anions. During electrolysis the cations move to the negative electrode (cathode) and the anions move to the positive electrode (anode).

Ions not fee to move

Ions are free to move

Fig. 3.13: State of lead (II) bromide in solid and molten form

(a)

(b)

Solid sodium chloride

Aqueous sodium chloride

Ions not free to move

Ions free to move

Fig. 3.14: State of sodium chloride in solid and solution form

Metals form cations which move to the –ve cathode when an electricity is passed though. Non-metallic elements form anions and these move to the +ve anode. Positive radicals and hydrogen ions move to the cathode. Negative radicals move to the anode.

Why non-electrolytes do not conduct electricity

Bonding in non-electrolytes is covalent. When dissolved, there are no charged particles (ions) to carry an electric current. We learnt in unit 2 that a covalent bond is formed when atoms share a pair of electrons. In this way they form discrete molecules. Therefore non-electrolytes exist in the form of molecules.

Examples of covalent compounds that do not conduct

Compound	**Structure**
Sugar	$C_{12}H_{22}O_{11}$
Urea	$(NH_2)_2CO$
Ethanol	C_2H_5OH
Distilled water	H_2O

3.5 Electrolysis of molten compounds

An electrolytic cell is used to pass an electric current through a molten ionic compound. After a compound has been melted, heat must be supplied to keep the salt in molten state. Fig. 3.15, show how ions occur in a binary compound and how they move after melting the compound.

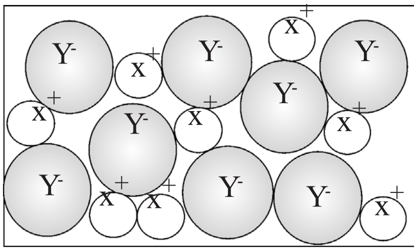
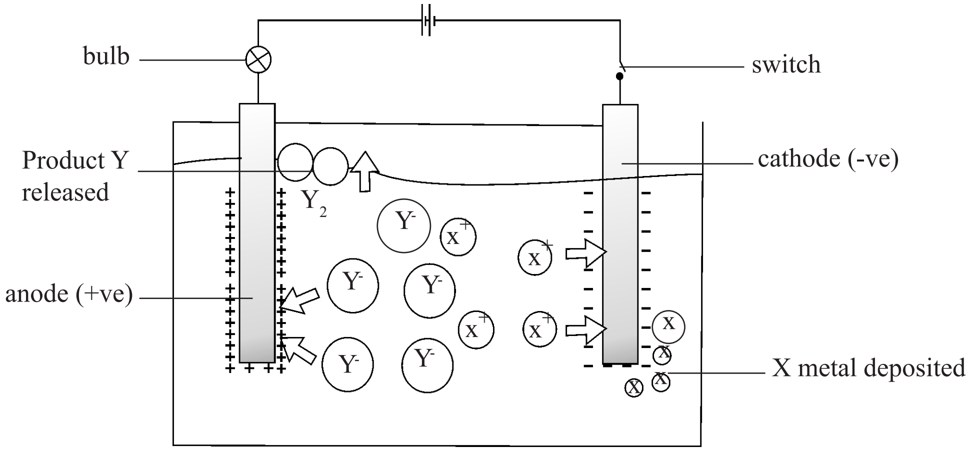


Fig. 3.15: Ions in a binary compound XY

When an ionic compound e.g. X^+Y^- is melted it produces X^+ ions and Y^- ions as follows:

$$X^+Y^-(s) \xrightarrow{\text{heat}} X^+(l) + Y^-(l)$$

When the switch is closed to complete the circuit, all +ve ions move to the cathode and all -ve ions move to the anode. As the current flow, a chemical reaction occurs at the electrodes. The ions lose the charge (they are discharged). Y– ions lose electron(s) while X+ ions gain electron(s). The respective products are deposited on the surface of each electrode, as shown in Fig. 3.16.



NON-METAL

- Negatively charged (-ve) ions are called anions and are attracted to the anode.

METALS

- Positively charged (+ve) ions are called cations and are attracted to the cathode.

Fig. 3.16: Discharging of anions at anode and cations at the cathode

Discharging ions

$X^+(l) \ + \rightarrow \ e^- \ X(s)$

$Y^- \ (l) \ - \ e \ \ Y(g)$

Let us now look at a particular salt and we compare it with the XY salt.

Electrolysis of molten Lead (II) bromide

Caution: Perform this experiment in a fume cupboard or in the open.

Lead (II) bromide is an ionic compound. When melted the electrolyte consists of two types of ions: - lead (II) ions and bromide ions.

Lead (II) bromide $\xrightarrow{\text{Heat}}$ lead (II) ions + bromide ions

$$PbBr_2(s) \xrightarrow{\text{Heat}} Pb^{2+}(l) \quad + \quad 2Br^-(l)$$

Lead (II) bromide is a binary salt. In molten state it produces two types of ions to be considered for discharging.

Activity 3.18

To investigate conductivity of molten lead (II) bromide.

In groups

Procedure

1. Set up the apparatus as in Fig. 3.12.
2. Fill a crucible to about two-thirds with lead (II) bromide.
3. Heat the salt until it melts. Keep on supplying heat to keep the salt in molten state.
4. Observe the bulb to see whether it lights.
5. Let the current pass for about 10 minutes.
6. Pay particular attention to any changes taking place around the electrodes. Record any observation made.
7. Remove the electrodes and examine them carefully. Record any observation.
8. After the crucible cools observe its contents. What do you observe around each electrode? Write equations to explain your observations.
9. Using the same procedure explain what will happen when an electric current is passed through lead (II) iodide.
9. Write equations to show products formed at the +ve and –ve electrodes.

The facts

The bulb lights brightly because molten lead (II) bromide conducts an electric current i.e. it is an electrolyte. Brown vapours are observed at the anode. These is bromine fumes produced during electrolysis. Beads of a grey solid are deposited on the cathode and at the bottom of the crucible. This is lead metal.

We can conclude that as the electric current passes through molten lead (II) bromide, it decomposes it to form lead metal and bromine. The following equations show what happens at the electrodes.

At the anode	**At the cathode**
Br^-(l) ions	*Pb^{2+}(l) ions*
Move to the anode and are discharged by losing an electronand become a bromine atom. $Br^-(l) \longrightarrow Br_2(g) + e$	Move to the cathode and discharge by gaining electrons o become lead atoms i.e.grey lead metal is formed. The metal is deposited on the cathode.
Bromine atoms combine in pairs to form molecules $Br(g) + Br(g) \rightarrow Br_2(l)$	$Pb^{2+}(l) + 2e \longrightarrow Pb(s)$

It is common to write one equation once we understand that there are very many Br^- ions in the melt as follows:

$2Br^-(l) \rightarrow Br_2(l) + 2e$

The compound is thus decomposed into the elements lead and bromine by the electric current:

$$\begin{array}{ccc} \text{Lead (II) bromide} & \xrightarrow{\text{electric current}} & \text{Lead} + \text{Bromine} \\ PbBr_2(l) & & Pb(s) + Br_2(l) \end{array}$$

Weak and strong electrolytes

Electrolytes are classified into weak electrolytes and strong electrolytes. Weak electrolytes are those electrolytes which allow small amount of an electric current to pass through them, thus they produce few ions in water. Examples of weak electrolytes include: Ammonium chloride, methanoic acid, ethanoic acid, citric acid.

Strong electrolytes are those electrolytes that allow large amount of an electric current to pass through them, thus they produce many ions in water. Examples of strong eletrolytes include: sodium chloride, sulphuric(VI) acid, hydrochloric acid, nitric(V) acid, potassium chloride.

Arrhenius theory states that: When an electrolyte is dissolved in water, it gets separated into electrcally positive and negative charged ions as cations and anions.

Preferential discharge of ions

When more than one type of ion migrates to an electrode, one of them is easily discharged at the expense of another. This is referred to as **preferential discharge**.

The factors affecting the preferential discharge of ion in electrolysis include:

1. The position of the ion in the electrochemical series.
2. The concentration of the competing ions in the electrolyte.
3. The nature of the electrode.

1. Position of ion in the electrochemical series

For cations, the higher an ion is in the electrochemical series the more difficult it is to discharge. Ions of less reactive elements are more easily discharged than those of the more reactive elements i.e ions which are lower in the electrochemical series.

Table 3.13: The electrochemical series for anions

Anion	**Ease of discharge**
SO_4^{2-}	**Least readily discharged**
NO_3^-	
Cl^-	
Br^-	
I^-	
OH^-	**Most readily discharged**

The lower an anion is in the series, the easier it is for it to be discharged.

For example, in a cell made of sodium chloride solution as the electrolyte. The ions at the anode will be chloride ions (Cl^-) and hydroxyl ion (OH^-). The hydroxyl ion will be preferentially discharged at the anode. At the cathode we shall have sodium ion (Na+) and Hydrogen ions (H^+), the hydrogen ions shall be preferentially discharged at the cathode.

Table 3.14: Electrochemical series for cations

Cation	**Ease of discharge**
K^+	**Least easily discharged**
Ca^{2+}	
Na^+	
Mg^{2+}	
Al^{3+}	
Zn^{2+}	
Cr^{3+}	
Fe^{2+}	
Pb^{2+}	
H^+	
Cu^{2+}	
Ag^+	**Most readily discharged**

2. Relative concentration of ions

This factor is very important for anions. Thus it mainly affects the products formed at the anode. The higher the concentration of the ions, the more easy for the ion to discharge. For example, in the electrolysis of dilute sodium chloride, the hydroxyl ions are preferentially discharged at the anode. In the case of concentrated sodium chloride solution, the chloride ions are preferentially discharged.

3. Nature of electrodes

The nature of electrode used influence the choice of ion for discharge. This can be explained as follow; In an electrolysis of NaCl solution, using platinum or mercury cathode, different products will be formed. When platinum cathode is used, H^+ ions are discharged in preference to Na^+ to give hydrogen gas.

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

Activity 3.19: To investigate preferential discharge in the electrolysis of dilute sodium chloride solution, dilute sulphuric acid and magnesium sulphate

In groups

Apparatus and chemicals

Beakers, 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, magnesium sulphate, dilute sulphuric (VI) acid.

Procedure

1. Put 200 cm^3 of dilute sodium chloride solution into a beaker and connect the electrodes to the connecting wires as shown in set up in Fig 3.17

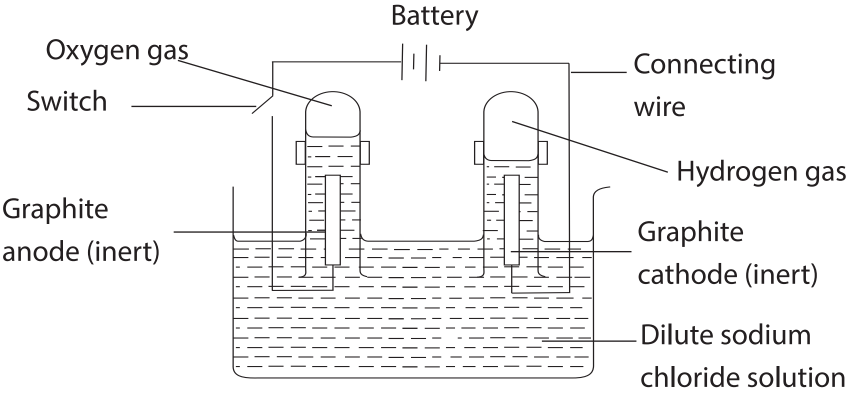


Fig. 3.17: Electrolysis of dilute sodium chloride solution

2. Connect the switch. After some time, disconnect the switch. Record any observations made at the electrodes.

- Write balanced questions for the reactions occuring at each of the electrodes.

3. Test the gases collected with a glowing and burning splints.

4. Repeat the experiment but using magnesium sulphate solution and dilute sulphuric acid in turn. Record observations made in each case in a table similar to Table 3.11

Table 3.11: Electrolysis of dilute sodium chloride solution

Gas collected at	Colour	Test with glowing/ burning splint
Anode		
Cathode		

The facts

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 hydrogen.

Ions present in the electrolyte

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 they are attracted by the oppositely charged electrodes.

At the cathode	At the anode
Ions present Na^+(aq), H^+(aq) Both ions move to the cathode. H^+ ion is discharged in preference to Na^+ ion because it is lower than Na^+ ion in the electrochemical series. It gains an electron to form H atom. $H^+(aq) + e^- \longrightarrow H\ (g)$ The hydrogen atoms combine to form H_2(g) molecules. $H(g) + H(g) \longrightarrow H_2(g)$ Therefore; $2H^+(aq) + 2e^- \longrightarrow H_2(g)$ But the $4e^-$ produced at the anode must react hence overall equation is $4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$	Ions present Cl^-(aq), OH^-(aq) Both ions move to the anode. OH^- ion is discharged because it is lower in electrochemical series OH^- is discharged by losing an electron $OH^-\ (aq) \longrightarrow OH\ (g) + e^-$ Four OH^- ions are discharged, the OH groups interact to form one mole of oxygen and two moles of water as shown below, releasing 4 electrons which move to the cathode. Overall reaction $4OH^-\ (aq) \longrightarrow O_2\ (g) + 2H_2O(l) + 4e^-$

Na^+ and Cl^- ions remain in solution forming sodium chloride..

Electrolysis of dilute sulphuric acid

Electrolysis of dilute sulphuric acid is also known as electrolysis of acidified water.

Ions present in the electrolyte
From sulphuric (VI) acid, $H^+(aq)$, SO_4^{2-} (aq)

From water $H^+(aq)$, OH^- (aq)

At the cathode	**At the anode**
Ions $H^+(aq)$ H^+ions moves to the cathode and gain an electron to form H atom. $H^+(aq) + e^- \longrightarrow H\ (g)$ Since hydrogen is a diatomic gas, the atoms combine to form a hydrogen molecule. $H\ (g) + H\ (g) \longrightarrow H_2\ (g)$ The $4e^-$ from the anode are taken up by H^+ ions giving 2 moles of hydrogen gas. Overall ionic equation $4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$	Ions $SO_4^{2-}(aq)$, $OH^-(aq)$ Both ions move to the anode OH^- ion is lower in the electrochemical series and is therefore discharged preferentially to SO_4^{2-} ion by loss of one electron. $OH^-\ (aq) \longrightarrow OH\ (g) + e^-$ The OH(g) ions interract to form 1 mole of O_2 and 2 moles of water Overall ionic equation. $4OH^-(aq) \longrightarrow 2H_2O(l) + O_2\ (g) + 4e^-$

The volume of hydrogen liberated at the cathode is *twice* the volume of oxygen produced at the anode. That is because as more water dissociates to replace the discharged $OH^-(aq)$ ions the concentration of H^+ (aq) ions at the cathode increases doubling the volume of hydrogen gas produced at cathode.

Electrolysis of magnesium sulphate solution

The solution contains Mg^{2+}(aq), SO_4^{2-} (aq) from magnesium sulphate and H^+(aq), OH-(aq) from water.

During electrolysis of magnesium sulphate solution, the movement of ions is as follows.

Anode	Cathode
OH^- and SO_4^{2-} ions migrate here OH^- ions are preferentially discharged because they are lower in the electrochemical series. Oxygen is liberated as shown in the equation below $4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$	H^+ and Mg^{2+}ions migrate here H^+ ions gain electrons and are discharged preferentially to Mg^{2+} ions because they are lower in the electrochemical series forming hydrogen gas. $2H^+(aq) +$ e- $\longrightarrow H_2(g)$. Account for the $4e^-$ from the anode.

Electrolysis of brine

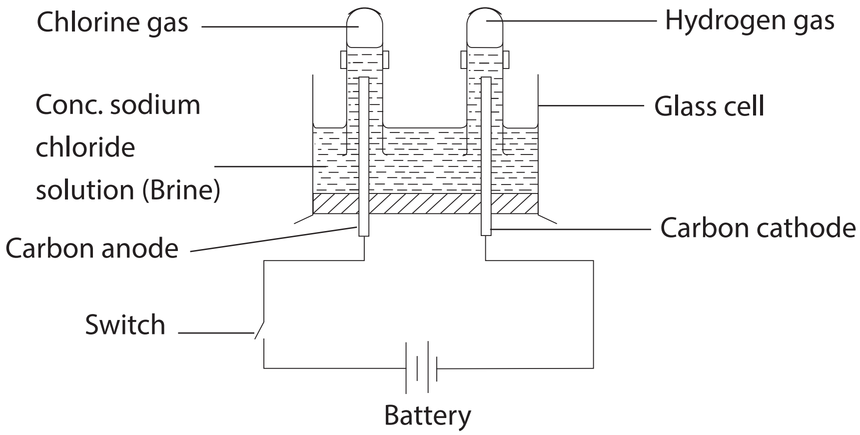


Fig. 3.18: Electrolysis of brine

At the anode, green-yellow gas with a choking irritating smell evolves. The gas is chlorine. At the cathode, a colourless gas which burns with a ‘pop’ sound is collected. This gas is hydrogen.

Ions present

From sodium chloride	$Na^+(aq)$, Cl^- (aq)
From water	$H^+(aq)$, OH^- (aq)

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 the following

(a) The **concentration** has determined the product at the cathode.

(b) H^+ and Cl^- ions are discharged leaving Na^+ and OH^- ions.

Chlorine would react with sodium hydroxide solution to form sodium hypochlorite which is a bleaching agent, and sodium chloride.

Ionic equation.

$$Cl_2\,(g) + 2OH^-(aq) \longrightarrow Cl^-\,(aq) + OCl^-\,(aq) + H_2O\,(l)$$

Electrolysis of copper (II) sulphate using different electrodes

Use of different electrodes in electrolysis of copper (II) sulphate solution has some significant effect on the electrodes and the products.

Electrolysis of copper (II) sulphate solution using inert electrodes (carbon rods)

Activity 3.20: Aim: To investigate how nature of electrode affects electrolysis of copper (II) sulphate solution

Inert electrodes are substances like carbon (graphite) or platinum. They take no part in electrolysis except to carry electrons to and from the electrolyte.

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. 3.19.

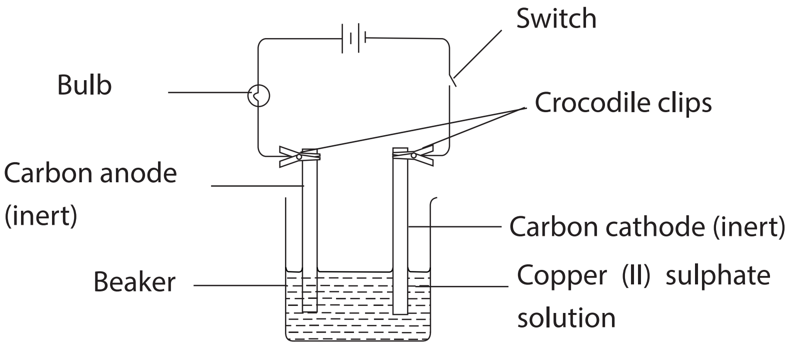


Fig. 3.19: Electrolysis of copper (II) sulphate using inert electrodes

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. From your experimental results, should carbon rods be used in the electrolysis of copper (II) sulphate solution?

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: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	• 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 the blue colour. As they are discharged and deposited at the cathode as copper solid, the colour of the solution fades and if we continue with electrolysis, the solution finally turns colourless and acidic.
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 3.21: Aim: Electrolysis of copper (II) sulphate solution using copper electrodes.

In groups:

Apparatus and chemicals

As in activity 3.19 but using copper electrodes, balance (sensitive), distilled water, propanone.

What do you suppose will happen if the copper electrodes used in electrolysis of copper (II) sulphate solution?

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.
4. Set up the apparatus as shown in Fig. 3.20 and switch on electric current.

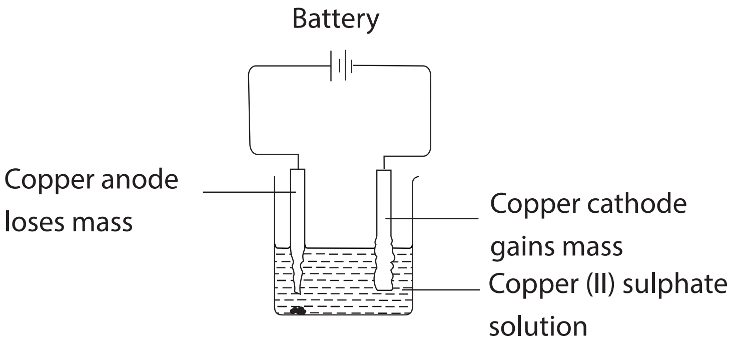


Fig. 3.20: Electrolysis of copper (II) sulphate using copper electrodes

5. After 10 minutes, note the colour of the electrolyte and rinse the electrodes with distilled water.
6. Dip the electrodes in propanone and then let propanone dry.
7. When dry, re-weigh the electrodes.
8. Record the mass of electrodes 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
$Cu^{2+}(aq)$, $H^+(aq)$ 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 on the cathode and it causes an increase in 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 mass.

The electrodes must be cleaned to remove any oxide that might be on the surface. The electrodes are washed with water and propanone to ensure they are completely dry.
Loss in mass of the anode electrode should be equal to the mass gained by cathode electrode.

3.6 Quantity of electricity

We have seen that when we pass electricity through an ionic compound in molten or solution form, it is decomposed and we get products at the electrodes. Experiments have shown that the mass or volume of elements of the products formed depends on the following factors:

1. amount of electricity passed.
2. length of time taken to pass the steady current.
3. charge on the ions of the element making electrolyte.

The quantitative laws relating the amount of electricity passed during electrolysis, to the mass or volume of elements formed at the electrodes were first formulated by Michael Faraday.

Fig. 3.21: Michael Faraday (1716-1867).

The coulomb

The coulomb is the practical unit used to measure the quantity of electricity or charge. An electric current is produced in the opposite direction to the flow of electrons in a circuit.

By definition, one coulomb (1 C) of charge is transferred if a current of 1-ampere (1 A) flows through a circuit for 1 second (1 s)

i.e. $1C = 1$ amp/sec or (As^{-1})

Generally, if Q coulombs of electricity flow along a wire for (t) seconds, the electric current (I),produced is given by :

$I =$ we can re-arrange the equation to get Q as:-

$$Q = I \times t$$

(Charge in) coulombs	Current in Amperes	time in

OR

coulombs amperes × seconds

C = A × s

In this section, we are going to look at these laws and learn how to use them in determining the quantity of substances deposited during electrolysis.

Example

Calculate the quantity of electricity passed when a current of 0.6 ampere flow for 50 minutes.

Solution

$Q = It$ Or $C = A \times s$

Substituting values given

$C = 0.6 \times (50 \times 60)$

$= 1800$ C

Therefore,the quantity of electricity passed is 1800 C

Farady's first Law of electrolysis

Activity 3.22:

In groups

To determine the amount of copper liberated during electrolysis using different quantities of electricity.

Apparatus and chemicals

- sensitive balance (electrical/digital)
- 2 strips of copper electrodes (same size)
- battery / dry cells 6V
- rheostat (variable resistant)
- connecting wires
- emery paper
- ammeter /stop clock
- copper (II) sulphate solution

Procedure

1. Set up the circuit as shown in Fig 3.22.
2. Fill the beaker to about two-thirds with dilute copper (II) sulphate solution.

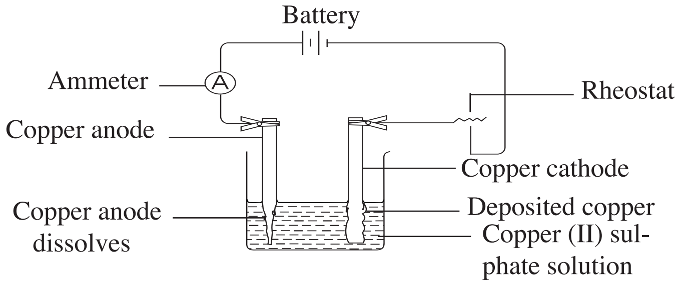


Fig. 3.22:Investigating the amount of copper deposited by different quantities of electric current

1. Clean the copper electrodes with emery paper until they shine.
2. Weigh the electrode attached to the cathode.
3. Place both electrodes into the solution.
4. Start the stop clock/stopwatch and simultaneously, complete the circuit by closing the switch.
5. Adjust the rheostat to obtain a steady current of about 0.2 to 0.5A.
6. Switch off the current after every fifteen minutes.
7. Record the current passed (in amperes) and time (in seconds) in a table similar to Table 3.16.
8. Remove the cathode electrode from the electrolyte, rinse it with distilled water, then dip it in propanone or ethanol and let it dry.
9. Weigh the cathode and determine the change in mass. Record this as mass of copper deposited in Table 3.16.
10. Repeat the experiment using different quantities of electricity as shown in Table 4.7.

Table 3.16 Specimen results

Current (A)	time(s) (15min × 60s)	Quantity of electricity Q = It	Mass of copper deposited (g)
0.5	600	300	0.098
1.0	600	600	0.196
1.5	600	900	0.392
2.0	600	1200	0.784

Draw a graph of copper deposited against quantity of electricity from the results in Table 3.16 above.

- What relationship do you notice between the quantity of electricity passed and the mass of copper deposited?

The facts

When the quantity of electricity is doubled, the mass deposited is also doubled. The relationship between the amount of copper deposited and the quantity of electricity passed can be shown graphically as shown in Fig 3.23.We get a straight line, which passes through the origin. This indicates that the mass of copper deposited is directly proportional to the quantity of electricity passed.

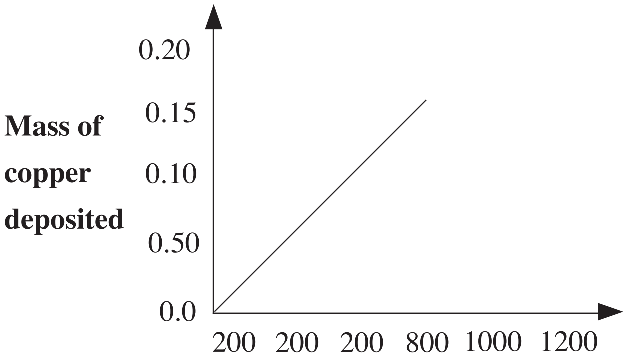


Fig. 3.23: Graph of mass of copper deposited against quantity of electricity passed

The famous scientist Michael Faraday did much of his work on electrolysis. He expressed the results of the many experiments he did in the law named after him. **Faraday's first law** of electrolysis can be stated as follows.

The amount of substance consumed or produced at one of the electrodes during electrolysis is directly proportional to the amount of electricity that is passe through the electrolytic cell.

From the above experiment, we have seen that it is possible to measure the mass of an element produced at an electrode during electrolysis. From such experiments, we can work out the quantity of electricity that would give 1 mole to the element.

Example

A student set up an experiment for the electrolysis of copper (II) sulphate solution. A current of 0.5A was passed through the circuit for 3860 seconds. A mass of 0.64g of copper was deposited.

(i) Which electrode gained mass?

(ii) How many coulombs are needed to deposit one mole of copper?

(R.A.M of Cu = 64)

Solution

(i) The cathode gained mass

(ii) Quantity of electricity(Q) = Current × time

(in amps) (in seconds) i.e

$Q = It$

$= 0.5 \times 3860$

$= 1930$ C

This means 1930 C of electricity deposited 0.64 of copper.

However, we know that 1 mole of copper has a mass of 64g. From the relative atomic mass (RAM) given we then say:

If 1930 C can deposit 0.64g of copper

x C would deposit 64g of copper.

Cross multiply and solve for x.

x = 1930 C × 64 g

0.64 g

= 193,000 C

OR

Working from mole conversions:-

1 mole of Cu = 64g

y mole of Cu = 0.64g.

Cross multiply and solve for y.

y = 0.64 g = 0.01mol

64 g

1930C deposited 0.01 mol

x C would deposit 1 mol

x = 1930 C × 1 mol

0.01 mol

= 193,000 C

We have seen how to work out the quantity of electricity (charge) needed to deposit 1 mole of an element. Now, let us see how we use the quantity of electricity (coulombs) to work out the charge of an ion.

Determining charges carried by ions using quantity of electricity used

The charge on an ion, whether a cation or anion, depends on its valency. It depends on the number of electrons lost, for cations or number of electrons gained, for anions.

The Faraday's Constant

It has been found through many experiments that when molten compounds and aqueous solutions are electrolysed, the quantity of electricity needed to produce one mole of atoms of any element is always 96500 coulombs or a multiple of 96500C. For example, in the previous example, it was found that 193,000C of electricity were required to deposit 1 mole of copper. This is twice 96500C i.e. (96500C × 2) = 193000C. This is also true with other elements. Table 3.17 shows how much charge is required to produce one mole of some of the atoms

Table 3.17 The charge needed to produce 1 mole of different elements

Element	Ion	Charge on ion	Charge needed to produce 1 mole (C)	Number of Faradays
Ag	Ag^+	+1	96500	1F
Na	Na^+	+1	96500	1F
H	H^+	+1	96500	1F
Cu	Cu^{2+}	+2	$193000 = 2 \times 96500$	2F
Mg	Mg^{2+}	+2	$193000 = 2 \times 96500$	2F
Al	Al^{3+}	+ 3	$289500 = 3 \times 96500$	3F
Br	Br^-	-1	96500	1F

Note that 96,500 coulombs is called the **Faraday's constant**. It's symbol is F. It was named in honour of Michael Faraday.

Now, let us analyse the equations for discharging some ions in Table 3.17 that would take place at the electrodes.

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

1 mol of Ag^+ ion requires 1 mole of electrons or 1F to deposit 1 mole of Ag atoms

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

1 mole of Cu^{2+} ions requires 2 moles of electrons or 2F to deposit 1 mole of copper atoms.

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

1 mole of Al^{3+} ions requires 3 moles of electrons or 3F to deposit 1 mole of Al atoms.

We can put the relationship in Table 3.17 to read

1F = 1 mole of electrons = 96500C.

Using this relationship, we can calculate the charge on an ion as shown in the following examples.

Faraday's second law of electrolysis

When the quantity of eletricity is passed through several electrolytes,the mass of the substance deposited are proportional to their respective chemical equivalent or equivalent weight.

Example 1

During electrolysis of copper (II) sulphate solution, a current of 0.5A was passed for a period 3860s.The amount of metal deposited was 0.64g.

(i) Calculate the charge carried by the metal ion.

(ii) Write an ionic equation for the reaction.

(R.A.M of Cu = 64, 1F = 96500C)

Solution

Quantity of charge = current × time (in seconds)

(Q) (I) (t)

$$= 0.5 \times 3860$$

$$= 1930\ C$$

So, 1930 coulombs deposited 0.64 g of copper

x coulombs would deposit 64 g of copper.

Cross multiply and solve for x

$$x = \frac{1930 \times 64}{0.64}$$

$$= 193000 \text{ coulombs}$$

$$1F = 96500\ C$$

$$y\ F = 193000\ C$$

$$y = \frac{193000\ C}{96500\ C}$$

$$= 2F$$

2F is the same as 2 moles of electrons.

The charge on the ion is +2.

Ionic equation

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

Example 2

During electrolysis of an aqueous solution of a salt of metal M, a current of 2.0A was passed for 32 minutes 10 seconds.The mass of metal M deposited was 2.24g.

(i) On which electrode was the metal deposited?

(ii) Calculate the quantity of charge needed to deposit 1 mole.

(iii) Calculate the charge carried on the ion.

(iv) Write an ionic equation to show how the ions of M are discharged at the electrode. (R.A.M of metal M is 112)

Solution

(i) At the negative electrode (cathode).

(ii) Quantity of charge = current × time (in seconds)

$Q = It$

On substituting the values = 2.0 × {(32 x 60) + 10}s

= 2.0 × 1930

= 3860 coulombs

3860 coulombs deposited 2.24g

x coulombs would deposit 112g

Cross multiply and solve for x

x = 3860C × 112g

2.24g

= 193000 coulombs

193000 coulombs are needed to deposit 1 mole

(iii) $1F = 96500C$

$yF = 193000\ C$

$y = 193000\ C$

$96500\ C$

$= 2F$

Since 1F = 1 mole of electrons, 2F means 2 moles of electrons were required. Therefore, the charge on the ion is +2.

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

3.7 Applications of electrolysis

Activity 3.23:Research Activity

Use reference material to find out why electrolysis plays an importtant role in our day to day life. Write a report and present to the class.

The facts

Some applications of electricity include:

1. Extraction of metals

Extraction of metals by the electrolysis process is known as **electro-metallurgy**. This process is used in extracting highly reactive metals such as sodium, magnesium and aluminium, e.g.

- Sodium is extracted from molten sodium chloride.
- Aluminium is extracted from molten bauxite (aluminum oxide, Al_2O_3).

2. Refining of metals

Copper is an excellent conductor of electricity widely used 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 illustrated in Fig 3.24.

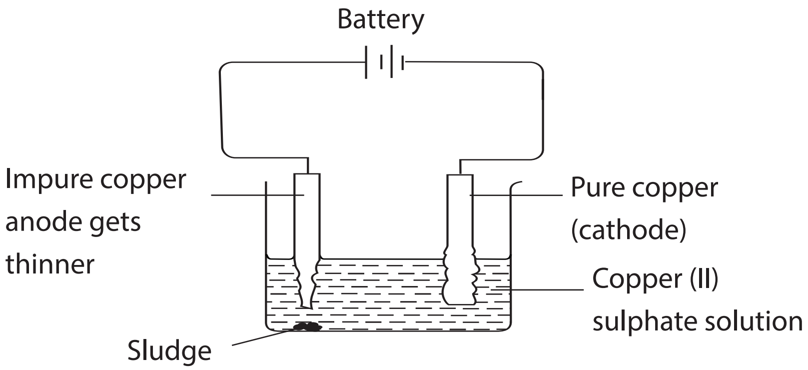


Fig 3.24: Refining impure copper

During refining of copper

- the cathode is made of thin sheets of pure copper.
- the anode is the impure copper.
- the electrolytic solution contains 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, accepting electrons and get discharged. They deposit on the cathode as copper metal.

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

Thus, the impure block at the anode dissolves and pure copper get 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 where a thin layer of desired metal is applied on a given object surface. It is done on some objects to make them look shiny and attractive to prevent rusting.

Examples

(i) Most metal objects made of iron or steel are nickel–chromium plated e.g. bicycle handle bars and car rims among others.

(ii) Silver or gold plating of 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 rusting.

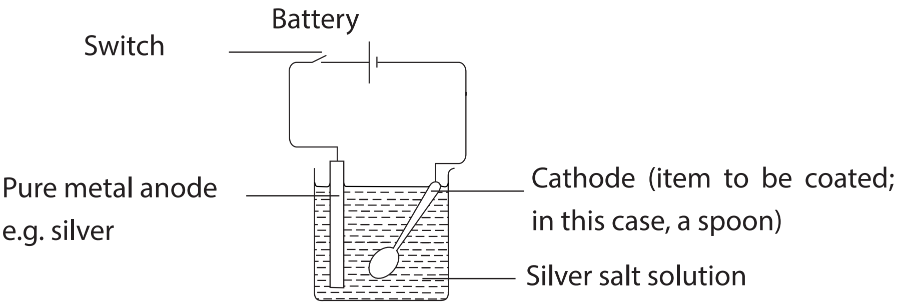


Fig. 3.25: Conventional electrolytic cell that can be used for electroplating

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 not alternating current(AC).
- The item to be coated is made the cathode. The coating metal is 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. 3.26 shows the set-up for the electrolysis process.

In this industrial process, mercury is made to flow along the cell as shown in Fig. 3.23. This mercury acts as the cathode where sodium is discharged.

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

Graphite is used as the anode because it is an inert electrode. It resists corrosion by chlorine being discharged and collected at the anode.

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

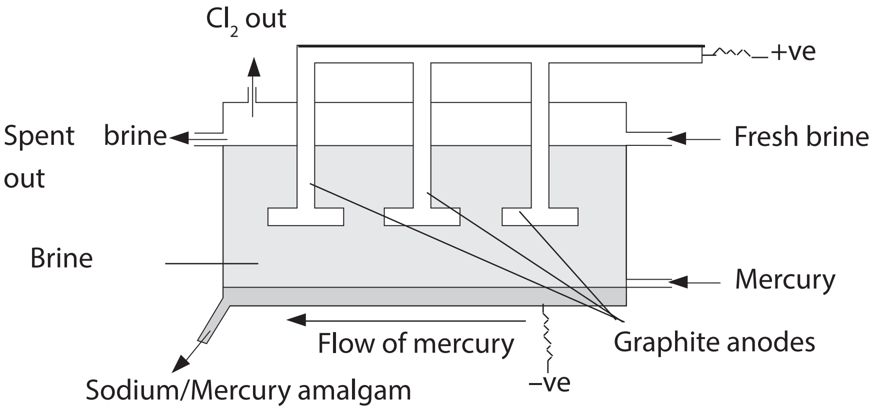


Fig. 3.26: The flowing mercury cell for the manufacture of sodium hydroxide, chlorine and hydrogen

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.

$$2NaHg\,(l) + 2H_2O\,(l) \longrightarrow 2NaOH(aq) + H_2\,(g) + 2Hg(l)$$

(sodium-mercury)

(amalgam)

The mercury released is recycled.

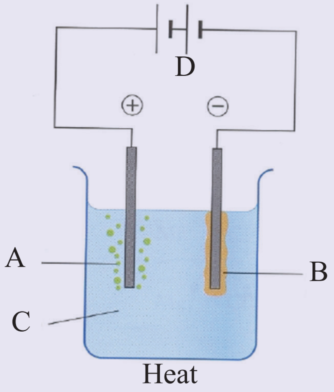
Caution: Mercury is toxic and can cause poisoning to animals and humans. All measures must be taken to ensure no mercury is discharged to the environment by any means.

Check your progress 3.2

1. Explain what you understand by the following terms:
   (i) Electrolysis
   (ii) Cathode and Anode
   (iii) Strong electrolyte and weak electrolyte
   (iv) Electrolysis and non-electrolysis. Give each case.

2. Explain the difference between a metallic conductor and an electrolyte

3. Assume the diagram below represents the set-up used during the electrolysis of molten lead (II) iodide using graphite electrodes. What is observed at:
   (a) (i) A
   (ii) B



(b) Why is it necessary to melt lead (II) iodide during electrolysis process?

4. Draw a labelled diagram of a simple electrochemical cell, which has zinc and copper immersed in their respective electrolytes. Show the direction of electron movement. Briefly state how the cell generates an electromotive force. State how such a cell will finally fail to generate power.

5. During electrolysis of water acidified with sulphuric acid, two gases were produced at the electrodes.

(a) Name the gas produced at the anode. How can you test for the gas produced?

6. In the electrolysis of sodium chloride solution the hydroxyl ions is preferetially discharged at the cathode.

  (a) Explain why this is so using the reactivity series.Support your your answer using an ionic equation.

  (b) Explain how aqueous copper (II) sulplate can be separated using eletrolysis

7 Molten lead (II) iodide is said to be an electrolyte.Explain how the salt is separated by electrolysis.

8. State some application of electrolysis.

9. Copper (II) sulphate solution is eletrolysed using inert electrodes. A current of 5A was passed for half hour.How much copper was deposited at the cathode?(R.A.M of Cu = 63.5

10. A current of 5.0 A was passed through molten anhydrous calcium chloride,(Ca Cl_2), for 10 minutes using inert graphite electrodes.Calculate the mass of each product liberated at each electrode (R.A.M of Ca = 40,Cl=35.5)

UNIT 4

Carbon, its atomic structures and compounds

Learning Outcome

Knowledge and understanding	Skills	Attitudes
• Describe the composition of atmosphere.	• Research second hand evidence from a wide range of sources. • Investigate and report on the significance of carbon compounds. • Present reasoned explanations for phenomena, patterns and relationships.	• Appreciate the importance of carbon to all life on Earth.

4.1 Introduction

Activity 4.1

In groups,

Study the photographs below.

A

B

C

D

E

F

Fig 4.1: Photographs showing some sources and uses of carbon and some of its compounds.

1. In small groups, discuss what the substances above are made of.
2. What are uses of the substances?

Work to do

1. Discuss on the elements making up our bodies. Is carbon one of them?
2. Suggest other compounds that contain carbon.
3. Discuss why carbon is an essential element in the body of living things.

Did you know?

Carbon occurs in all living organisms in the universe?

The facts

All living things contain carbon compounds. These include carbohydrates, proteins, fats among others. Petrol, oil and wood are carbon compounds. Carbon is incorporated in plants from carbon(IV) oxide during photosynthesis.

Carbon is a non-metal which occurs naturally as a free element or in its combined state as in carbonates and hydrogen carbonates. In the periodic table of elements, carbon appears in the second period of group 4 elements with atomic number 6.

Group	I	II		III	IV	V	VI	VII	VIII
Period 1	H 1								He 2
2	Li 3	Be 4		B 5	$^{12}_{6}C$ **Carbon**	N 7	O 8	F 9	Ne 10
3	Na 11	Mg 12		Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
4	K 23	Ca 20							

Fig 4.2 a: Position of carbon in the periodic table

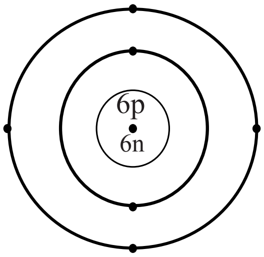


Fig 4.2 b: Atomic structure of carbon

4.2 Allotropes of carbon

Work to do

In pairs

1. Carbon exhibits allotropy and occurs different forms suggest the different form of carbon and the allotropes of carbon.
2. Suggest other elements that exhibit allotropy.

The facts

Some elements can take two or more different forms where atoms are arranged differently by chemical bonds. These different forms of an element in the same physical state are called **allotropes**. The existence of an element in more than one form but in the same physical state is what is referred to as **allotropy**. Carbon exists in two allotropic forms namely Graphite and Diamond.

i. Graphite

Work to do

1. Research on how the carbon atoms in graphite are joined together within the structure.
2. Which are the forces joining these carbon atoms together?

Activity 4.2

In groups,

Your teacher will provide you with a sample of grease and a carbon/graphite rod.

1. From Unit 3, You learnt that graphite rods are used as electrodes in electrolysis. Meaning they allow passage of electric current. Find out from the internet why this is possible.
2. Mention other non-metals which can conduct electricity.
3. Hold a small portion of grease between your fingers and gently move your fingers.
4. Describe the feeling.

The facts

A carbon atom has four valence electrons. In graphite, each carbon atom is bonded to three other carbon atoms using the covalent bonds. This means that, only three of the four valence electrons takes part in bonding leaving one non-bonding electron in each carbon atom. Within the structure, there are so many delocalized electrons. This explains why, graphite conducts electricity.

The interactions of the carbon atoms covalently bonded forms a network of hexagonal layers. In addition, these hexagonal layers are held together to one another by a weak force of attraction called the Van der Waal force making them slide over each other. For this reason, graphite tends to have a slippery feeling, soft and slippery, graphite can be used as a lubricant or in pencil leads.

The arrangement of carbon atoms in graphite forms a giant atomic structure in which carbon atoms are joined together by strong covalent bonds. Graphite has a density of 2.7g/cm^3 and boils at 4,830 °C

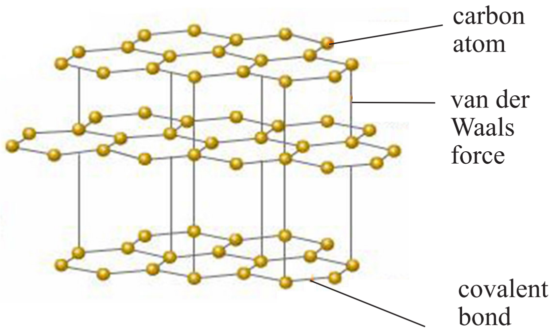


Fig 4.3: Structure of graphite

ii. Diamond

Activity 4.3

In groups,

- Try to drill through a rock using a wooden piece. Have you succeeded?
- Using a scalpel blade, try to cut an iron nail into two equal parts. Is it possible?
- Suggest another tool which can be used to cut the iron nail into two parts.

Work to do

1. Research on the places where diamond is mined in your country.
2. Suggest some materials used in making some of the most expensive jewellery.
3. Have you ever been to a place where metals are cut?

The facts

Diamond is an allotrope of carbon in which carbon atoms are covalently bonded to one another. During bonding, each carbon atom is bonded to four other carbon atoms. All the four valence electrons in each carbon atom take part in bonding. Therefore, there are no delocalized electrons in a diamond structure. For this reason, diamond does not conduct electricity.

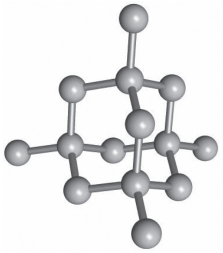


Fig 4.4: Structure of diamond

Arrangement of carbon atoms in diamond forms a tetrahedral three dimensional structure with strong covalent bonds to form a giant atomic structure.

The close packing of carbon atoms and the strong covalent bonds makes diamond the hardest substance known. It is used in cutting metals and drilling through hard rocks. In addition, the high refractive index of diamond makes it useful in making jewellery. Diamond exists as a shiny crystalline structure with a density of 3.5g/cm^3 and a very high boiling point.

Diamond and graphite are both giant covalent substances made entirely of carbon atoms. Both have chemical formula, but they do not have a molecular formula.

Table 4.1 Differences between graphite and diamond

	Graphite	**Diamond**
Structure	It is also Giant covalent structure, with each carbon covalently bonded to three other carbon atoms in a hexagonal arrangement.	Giant covalent structure, with each carbon covalently bonded to four other carbon atoms in a tetrahedral arrangement to form a rigid structure
Hardness	Soft. Layers of hexagonally arranged carbon atoms can slide over one another.	Extremely hard. Due to rigid, tetrahedral arrangement of carbon atoms.
Electrical conductivity	Conductor. Three out of four valence electrons are used for covalent bonding with other carbon atoms. Remaining valence electrons can be delocalised across the planes of carbon atoms.	Insulator. Mobile electrons are absent. All four valence electrons are used in covalent bonds.

Economic value

Diamond is a precious gem with high economic value. The economy of a nation lies in its ability to tap such natural resources.

Check your progress 4.1

1. Define the term allotropy.
2. Carbon in its pure state occur in two allotropic forms. Which one?.
3. Suggest a substance which can be used to make drilling objects. Especially, hard rock drills.
4. Compare the properties of the allotropes of carbon.

4.3 Amorphous carbon

Activity 4.4

In groups,

Separately burn the following substances: dry grass, saw dust and sugar. Note the color of the residue obtained after each heating.

a) Take a piece of charcoal and set it on fire
b) Allow it to burn until the fire dies off.
c) Was there a solid residue?
d) What was the color of the flame?

Work to do

1. Have you ever used charcoal as a fuel?
2. How do you think charcoal is made?
3. Would you prefer charcoal to other forms of fuel?

The facts

Armophous carbon refers to the impure forms of carbon which includes:- charcoal, lamp, black soot, coal and coke. Unlike graphite and diamond, amorphous carbon consists of minute crystals of graphite, hence their irregular shapes.

Charcoal can be made by heating wood in the absence of air. The slow combustion of part of the wood converts the rest to charcoal.

My environment, my life!

Excessive cutting of trees may lead to deforestation. Let us conserve our environment by planting more trees.

Armophous carbon has a wide range of use:-

- Carbon black is used in making carbon papers.
- Powdered carbon is used in gas masks to adsorb poisonous gases.
- Carbon black is used as a pigment in inks and paints.
- Carbon black is used as filler in rubber which helps it resist deterioration by U.V rays as well as giving it the black appearance.

Properties of carbon

Activity 4.5

In groups,

1. Burn a small piece of charcoal in an open crucible. Note the color of the residue if any.
2. Repeat the procedure using other forms of carbon.

Reaction with oxygen

The facts

All forms of carbon burn in oxygen to form carbon (IV)oxide and a lot of heat is produced. For the impure forms of carbon, ash is also formed.

$$\text{Carbon} + \text{Oxygen} \longrightarrow \text{Carbon (IV) oxide}$$

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

Reducing property of carbon

Activity 4.6

In groups,

What happens when carbon is heated together with black copper (II) oxide?

Procedure

1. Set up your apparatus as shown in the figure below.
2. Heat a mixture of powdered carbon and copper (II) oxide strongly in a combustion tube.
3. Note any colour changes.
4. What do you think will happen to copper (II) oxide in the presence of carbon after heating?

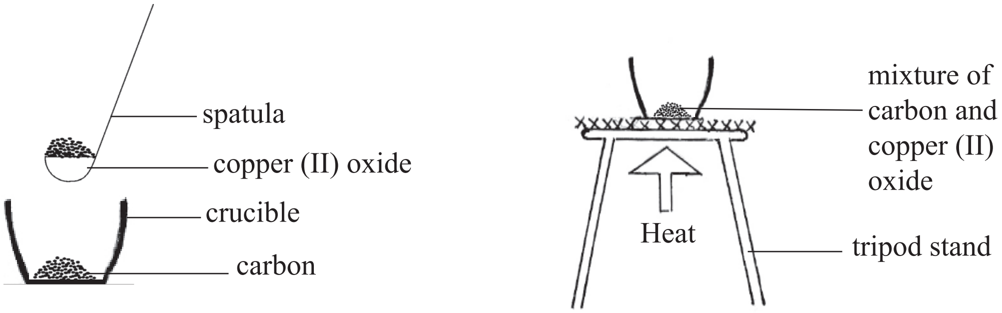


Fig 4.5: Diagram showing carbon reducing copper (II) oxide

1. What is the color of the solid mixture before heating?
2. What is the colour of the solid after heating?
3. What will be observed in the boiling tube containing calcium hydroxide solution?
4. Write the chemical equation for above experimaental observation.

The facts

Carbon can remove the combined oxygen from heated metal oxides of those metals below the reactivity series into their respective metals. The process of removing the combined oxygen from a substance is referred to as **reduction.**

Carbon reduces heated copper (II) oxide (black) into brown copper metal. This explains the color change in the combustion tube from black to brown.

Copper (II) oxide + Carbon $\longrightarrow$ Copper + Carbon (IV) oxide.

$$2CuO(s) + C(s) \longrightarrow 2Cu(s) + CO_2(g)$$

The carbon (IV) oxide produced forms a white precipitate with calcium hydroxide solution in the boiling tube.

Lead (II) oxide is reduced by carbon into silvery balls of lead metal while carbon itself is oxidised to carbon (IV) oxide.

Lead (II) oxide + Carbon $\longrightarrow$ Lead + Carbon (IV) oxide

$$2PbO(s) + C(s) \longrightarrow 2Pb(s) + CO_2(g)$$

The reducing property of carbon makes it widely used in the extraction of metals such as: - zinc, iron, lead and copper from their natural sources.

My environment, my life!

High concentration of carbon (IV) oxide in the atmosphere causes greenhouse effect which may lead to global warming. Let us take care of our environment.

Other reducing action of carbon.

Cold acids have no effect on carbon. However, hot concentrated acids react with carbon to form carbon (IV) oxide among other products.

CAUTION!

Experiments involving concentrated acids should only be performed under strict guidelines from your teacher.

Carbon reduces concentrated Sulphuric (VI) acid and concentrated Nitric (V) acid into sulphur (IV) oxide and nitrogen (iv) oxide respectively.

Concentrated sulphuric (IV) acid + Carbon $\longrightarrow$ Sulphur (IV) oxide + Carbon (IV) oxide + Water

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

Concentrated nitric (V) acid + Carbon → Nitrogen (IV) oxide + Carbon (IV) oxide + Water

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

Health check

Inhaling of poisonous gases poses danger to your general health. Stay safe always by wearing gas mask when dealing with such gases.

Check your progress 4.2

1. What do you understand by term armophous carbon?
2. State and explain the observations made when a mixture of carbon and copper (II) oxide is strongly heated in a crucible.
3. Other than fuel, suggest other uses of carbon.

4.4 Oxides of carbon

Work to do

1. Try to cover a burning candle with a container for a minute. What do you notice?
2. Think of a planet without trees or any vegetation cover. Can other living organisms survive in such a planet?
3. Have you ever used baking powder in your home? How does the dough containing baking powder behave when cooked?
4. Read the label on a fire extinguisher that will be presented to you by your teacher. What do you see?
5. What sound is produced when you open the bottle of a soda?
6. Have you ever seen an aircraft moving in the clear sky especially during the dry seasons leaving behind some kind of smoke that spreads uniformly in the sky? Ever wondered what the smoke like thing is?

The facts

Carbon (IV) oxide exists as free molecules in the atmosphere forming about 0.03% of air by volume. Carbon containing substances burn in air to produce Carbon (IV) oxide as one of the products. It's also a product of metabolism in animals. Plants use carbon (IV) oxide during photosynthesis process.

Laboratory preparation of carbon (IV) oxide

Activity 4.7

In groups

Apparatus and reagents.

- Marble chips (calcium carbonate)
- Gas jars
- Dropping funnel
- Lime water
- Water trough.
- Dilute hydrochloric acid
- Conical flask
- Candle
- Wooden splint

Set up your apparatus as shown below.

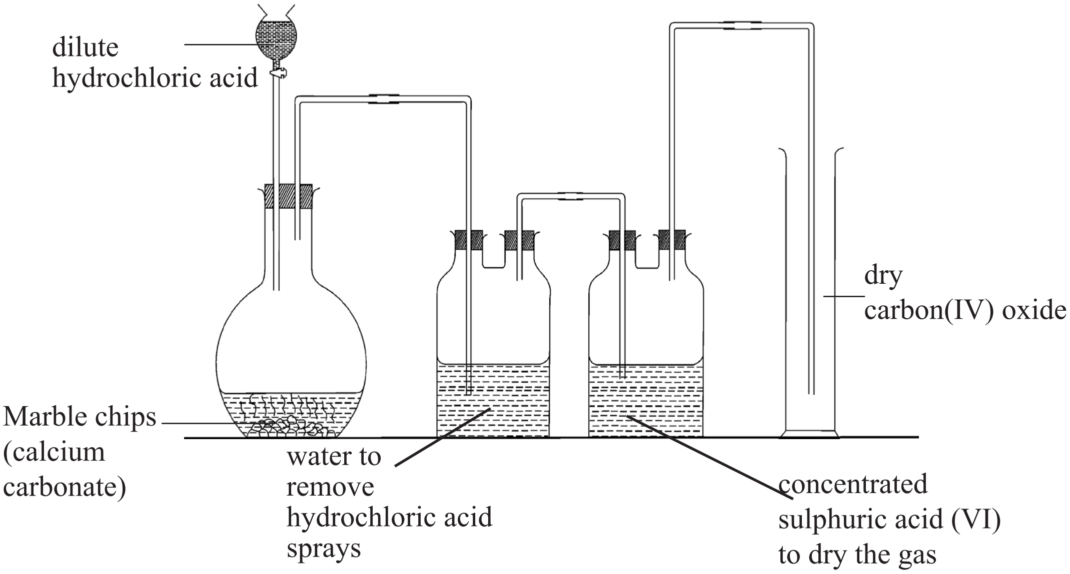


Fig 4.6: Diagram showing laboratory preparation of Carbon (IV) oxide.

Collect several jars of carbon (IV) oxide for use in the following experiments.

Note the colour and the smell of the gas.

Experiment	**Observations**
What is the purpose of water in the first bottle?What is the purpose of concentrated sulphuric(VI) acid in the second bottle?	
Why is the gas collcted by the method shown?	
Bubble carbon (IV) oxide through lime water for some time and then in excess. Note any color changes.	
Lower a burning piece of magnesium into a gas jar full of carbon (IV) oxide.	

The facts

Action of a dilute acid on a suitable metal carbonate produces carbon (IV) oxide. When dilute hydrochloric acid is added to calcium carbonate, effervescence occurs with the production of carbon (IV) oxide.

Calcium carbonate + Dilute hydrochloric acid ⟶ Calcium chloride + Carbon (IV) oxide + Water

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

Carbon (IV) oxide is then passed through water to remove the traces of hydrogen chloride fumes and then collected by downward delivery method. When required dry, carbon (IV) oxide is passed through concentrated sulphuric (VI) acid which acts as a drying agent.

Generally, carbon (IV) oxide can be prepared in the laboratory by the action of dilute acid on a suitable metal carbonate. However, reaction between dilute sulphuric (VI) acid with calcium carbonate takes place for a few seconds and then eventually stops. This is due to the formation of an insoluble salt calcium sulphate which forms a coating on the surface of the unreacted carbonate preventing further action of the acid.

Other carbonates that form insoluble compounds with dilute acids includes:-

1. Action of dilute hydrochloric acid on lead carbonate.
2. Action of dilute sulphuric acid on barium carbonate.

Physical properties

Properties of carbon (IV) oxide

i. It is a colourless gas.
ii. It is odourless.
iii. It is denser than air, hence collected by downward delivery method. Carbon (IV) oxide is 1.5 times denser than air.
iv. It readily liquefies and solidifies.

Chemical properties of carbon (IV) oxide

1. Combustion

Carbon (IV) oxide neither burns nor supports combustion. A burning splint is readily extinguished when lowered into a gas jar full of carbon (IV) oxide. Being more denser than air, carbon (IV) oxide is commonly used in fire extinguishers. When sprayed over the flames of a burning fire, it forms a 'blanket' that prevents entry of oxygen thereby extinguishing the fire.

However, a burning piece of magnesium continues to burn when lowered into a gas jar containing carbon (IV) oxide forming a white solid and black specks of carbon. Burning magnesium produces a lot of heat that breaks carbon (IV) oxide into carbon and oxygen. Carbon forms as black specks whereas the oxygen produced ensures the continued burning of magnesium to form magnesium oxide.

$$\text{Magnesium} + \text{carbon (IV) oxide} \longrightarrow \text{Magnesium oxide} + \text{Carbon}$$

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

2. Reaction with water

Carbon (IV) oxide slightly dissolves in water to form a weakly acidic solution of carbonic acid. Carbon (IV) oxide is an acidic gas.

$$\text{Carbon (IV) oxide} + \text{Water} \longrightarrow \text{Carbonic acid}$$

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

Rain water is slightly acidic due to the dissolved carbon (IV) oxide in the atmosphere.

3. Reaction with alkalis.

When carbon (IV) oxide is bubbled through calcium hydroxide solution, a white precipitate is formed. ***This is the confirmatory test for carbon (IV) oxide in the laboratory***. Being an acidic gas, carbon (IV) oxide reacts with alkali to form salt and water.

Calcium hydroxide + Carbon (IV) oxide ⟶ Calcium carbonate + Water

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

When excess Carbon (IV) oxide is bubbled through the resulting solution, a white precipitate dissolves to form a colourless solution.

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

Carbon (IV) oxide reacts with other alkalis to form their corresponding carbonates. Sodium carbonate and potassium carbonate are soluble. Therefore, carbon (IV) oxide does not form precipitates with alkaline solutions of sodium and potassium.

Check your progress 4 .3

1. A sample of rain water on analysis was found to have a pH value less than 7. Give a reason.
2. Would you expect a burning candle to continue burning when carbon (IV) oxide is ‘poured’ over the flame?
3. What is the significance of the property of carbon (IV) oxide displayed in (2) above?
4. Fire caused by oil spillage should be extinguished using the carbon (IV) oxide fire extinguisher. Give a reason.

Uses of carbon (IV) oxide

Carbon (IV) oxide can have some commercial uses which includes;

1. Manufacturing of carbonated drinks such as soda and beer.
2. Manufacture of fire extinguishers.
3. Solid carbon (IV) oxide *(dry ice)* is used as a coolant by the ice-cream vendors. This is because, solid carbon (IV) oxide sublimes leaving no dampness (wetness).
4. It is used in welding to provide an inert environment that prevents oxidation.
5. Used in rain seeding. Solid carbon (IV) oxide is sprayed in the sky by an aircraft to accelerate the rate of condensation of water vapour in the sky. This is a new technology for making rain.
6. It is used in the manufacturing of fertilisers.

Carbon (II) oxide (CO)

Health Check

Carbon (II) oxide is highly poisonous. When inhaled it causes suffocation which may lead to death. Practicals involving it should not be attempted in the laboratory. Stay safe.

Work to do

1. Think of a charcoal stove in use.
2. Discuss the advantages of using a charcoal store over using wood in cooking.

Occurrence

Carbon (II) oxide is formed when there is incomplete combustion of carbon such as using a charcoal stove in a poorly ventilated room.

$$2C(s) + O_2\,(g) \xrightarrow{\text{Limited supply}} 2CO\,(g)$$

Laboratory preparation of carbon (II) oxide.

Carbon (II) oxide is prepared by dehydration of organic compounds using suitable de hydrating agents. Dehydration is the process of removing elements of water (i.e) H and O from a compound to form H_2O.

In the laboratory Carbon (II) oxide is prepared by dehydrating methanoic acid or ethanedioic acid using concentrated sulphuric (vi) acid. During the process concentrated sulphuric (vi) acid remains chemically unchanged at the end of the reaction.

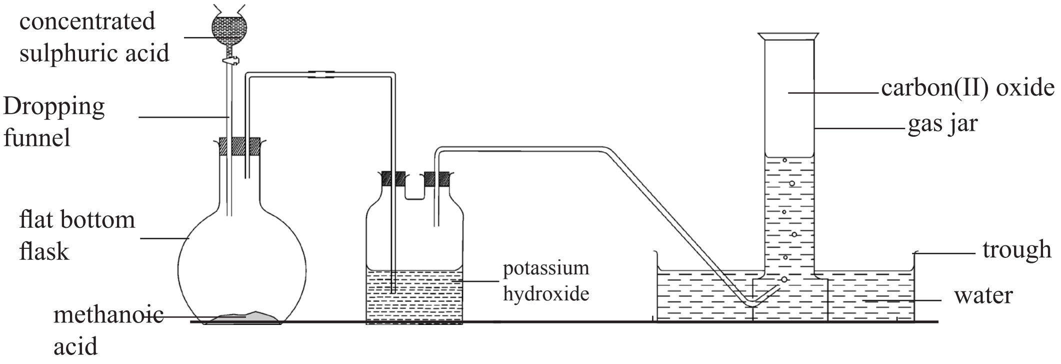


Fig 4.7: Diagram showing dehydration of methanoic acid.

$$\text{Methanoic acid} \longrightarrow \text{Carbon (II) oxide} + \text{Water}$$

$$HCOOH\,(aq) \longrightarrow CO\,(g) \quad + \quad H_2O\,(l)$$

When ethane (Oxalic acid) dioic acid is used a gaseous Carbon (II) oxide and carbon (IV) oxide is formed. The gaseous mixture is then passed through concentrated sodium hydroxide solution or concentrated potassium hydroxide solution to remove carbon (iv) oxide.

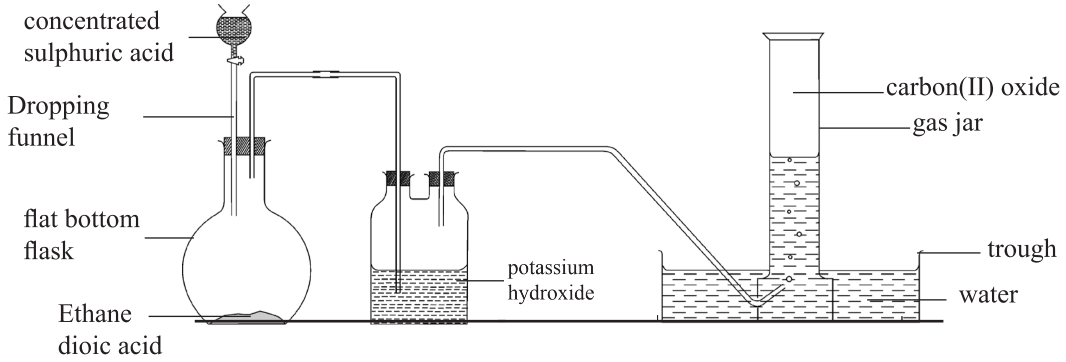


Fig 4.8: Showing dehydration of ethane dioic acid.

Ethane dioic acid + Concentrated Sulphuric (VI) acid ⟶ Carbon (II) Oxide + Carbon (IV) oxide + Water

$$H_2C_2O_4(aq) \xrightarrow{\text{Conc. Sulphuric (VI) acid}} CO(g) + CO_2(g) + H_2O(l)$$

Other sources of carbon (II) oxide are;

1. Carbon (IV) oxide can also be obtained by passing dry Carbon (IV) oxide over heated carbon.

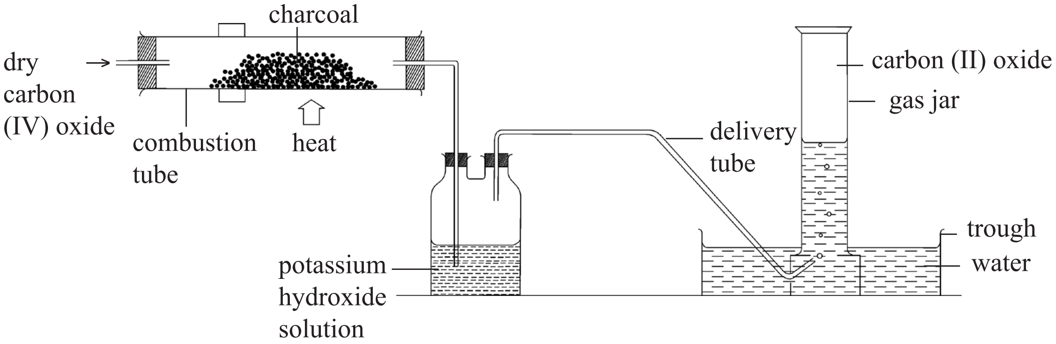


Fig 4.9: Preparation of carbon (II) oxide from carbon (IV) oxide

The facts

Carbon has reducing properties. When dry carbon (IV) oxide is passed over heated carbon, it is reduced to carbon (II) oxide.

Carbon (IV) Oxide + Carbon ⟶ Carbon (II) Oxide

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

Some properties of carbon (II) oxide

- It is a colourless gas.
- It is odourless.
- It is tasteless.
- It is slightly soluble in water hence collected by over water method.
- It is neutral. It has no effect on litmus papers.

Chemical properties of carbon (II) oxide

Work to do

1. While at home, observe a charcoal in use. Note the colour of the flame produced.

The facts

Combustion

Carbon (II) oxide burns quietly in air with a pale blue flame to form carbon (IV) oxide.

Carbon (II) oxide + Oxygen ⟶ Carbon (IV) oxide

$$2\,CO\,(g) + O_2\,(g) \longrightarrow 2CO_2\,(g)$$

Reducing properties of carbon (II) oxide

Carbon (II) oxide reduces heated metal oxides lower in the reactivity series into their corresponding metals.

E.g. When passed over heated copper (II) oxide, it reduces the black copper (II) oxide into brown copper metal.

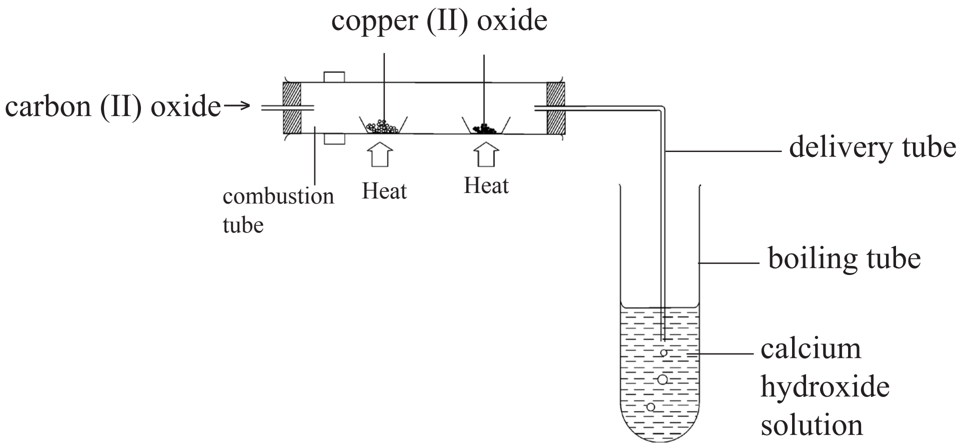


Fig 4.10: Action of carbon (II) oxide on copper (II) oxide

Copper (II) Oxide + Carbon (II) Oxide ⟶ Copper + Carbon (IV) Oxide.

$$\underset{\text{(Black)}}{CuO\,(s)} + CO\,(g) \longrightarrow \underset{\text{(brown)}}{Cu\,(s)} + CO_2\,(g)$$

Carbon (IV) oxide produced forms a white precipitate with calcium hydroxide solution in the boiling tube. Excess carbon (II) oxide should be burned or pumped through the fume chamber to prevent it from filling the room as it is highly poisonous. When inhaled, carbon (II) oxide combines with haemoglobin in the blood to form a stable compound carboxy – haemoglobin. This compound reduces the capacity of red blood cells to transport oxygen to the body tissues which may lead to suffocation and eventually death.

Health Check

Your health should be given the first priority. Never use a charcoal stove in poorly ventilated room as this may lead to suffocation or even death.

The charcoal stove

In the modern society a charcoal stove has been widely used as a heating tool. Commonly used for roasting meat or green maize.

But how does a charcoal stove work?

Consider a charcoal stove in use as shown in the figure below

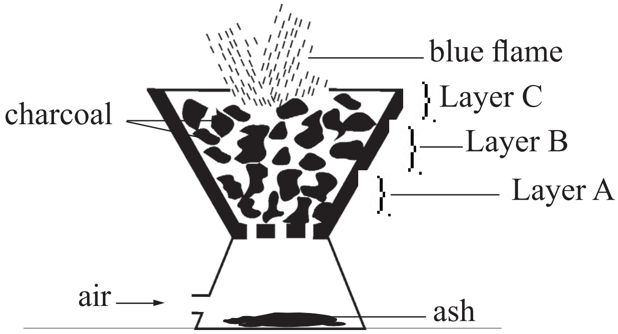


Fig 4.11: Burning charcoal in a charcoal stove

The facts

Layer A

In this layer, there is sufficient supply of air. Carbon will completely burn in this region to form carbon (IV) oxide

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

Layer B

In this region there is insufficient supply of air.

Two processes take place in this region.

(i) Incomplete combustion of carbon due to limited supply of air.

$2 C (s) + O_2 (g) \longrightarrow 2 CO (g)$

(ii) Reduction of carbon (IV) oxide from Layer A by the hot carbon to form carbon (II) oxide.

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

Layer C

In this layer, there is sufficient supply of air. Carbon (II) oxide from Layer B burns with a pale blue flame to form carbon (IV) oxide.

$2CO (g) + O_2 (g) \longrightarrow 2 CO_2 (g)$

Note: Using a charcoal stove in a poorly ventilated room may result in accumulation of Carbon (II) oxide which when inhaled may cause carbon (II) oxide poisoning through suffocation which may cause death.

It is also important to note that car engines produces some considerable amount of carbon (II) oxide. However some modern cars are fitted with catalytic converters which converts carbon (II) oxide into less poisonous carbon (IV) oxide.

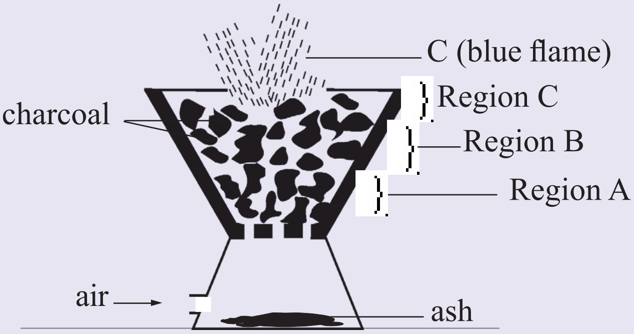
Uses of carbon (II) oxide

Carbon (II) oxide has commercial uses which includes:

1. It is used in the extraction of some metals.
2. A mixture of carbon (II) oxide and hydrogen (water gas) is used as a fuel.
3. It is used in the manufacture of methanol which is used as a solvent.

Check your progress 4.4

1. Both carbon (II) oxide and carbon (IV) oxide are colourless gases. Suggest a chemical test that can be used to distinguish the two.
2. In limited supply of air, carbon (II) oxide may cause suffocation. Explain.
3. What are some of the uses of carbon (II) oxide?
4. Below is a diagram showing a charcoal store in use. Use it to answer the questions that follows.



(a) Write two equations to describe the processes taking place in region B.

5. Compare the physical and chemical properties of carbon (IV) oxide and carbon (II) oxide.

4.5 Carbonates

Work to do

1. Make a list of carbonates that you know.
2. From your list of carbonates name the elements making each of the carbonate. Which element (s) seems to appear in almost all of the listed carbonates?
3. Test the solubility of these carbonates in water. What do you notice?

The facts

Carbonates are carbon compounds resulting from the action of carbonic acid or carbon (IV) oxide on an alkali.

For example,

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

$Ca\ (OH)_{2\ (aq)} + CO_2(g) \longrightarrow CaCO_3(s) + H_2O\ (l)$

Almost all carbonates are insoluble in water except group I carbonates and ammonium carbonate.

Action of heat on carbonates

Activity 4.8: To investigate the effect of heat on carbonates.

In groups

Apparatus and Reagents

- Sodium carbonates
- Lead carbonate
- Freshly prepared
- Potassium carbonate
- Copper carbonate
- Lime water
- Calcium carbonates
- Bunsen burner
- Boiling tubes

Procedure

1) In separate boiling tubes, heat a small portion of sodium carbonate, potassium carbonate, calcium carbonate, lead carbonate and copper carbonate strongly until there is no further change.
2) Note any colour change in each experiment. Test for any gas produced using freshly prepared lime water.

Set up your apparatus as shown below.

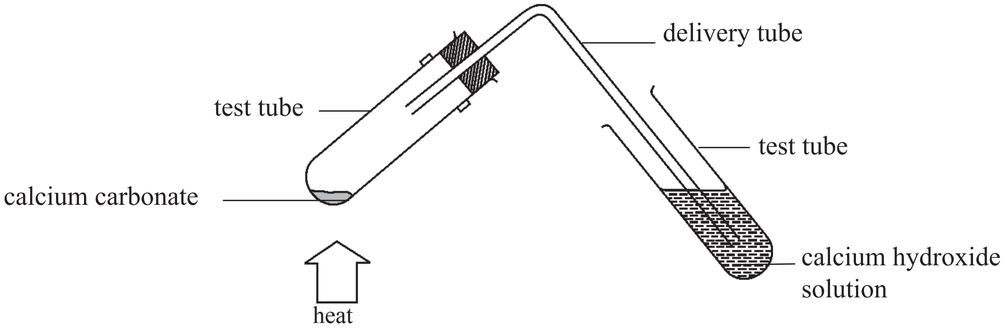


Fig 4.12: Action of heat on carbonates.

Table 4.2 Copy and complete the table below.

Carbonate heated	**Appearance**		**Effect of the gas on lime water**
	Before Heating	**After heating**	
Sodium Carbonate			
Potassium Carbonate			
Calcium (II) Carbonate			
Lead (II) Carbonate			
Copper Carbonate			

The facts

Carbonates of metals higher in the reactivity series do not decompose on heating.

K_2CO_3 (s) $\xrightarrow{\text{Heat}}$ No effect

Na_2CO_3 (S) $\xrightarrow{\text{Heat}}$ No effect

However, hydrated carbonates of sodium and potassium on heating losses water of crystallisation to form anhydrous compounds.

$$Na_2CO_3.10H_2O(s) \xrightarrow{\text{Heat}} Na_2CO_3(s) + 10H_2O\ (l)$$

Carbonate of metals lower in the reactivity series decomposes on heating to form the corresponding metal oxide and carbon (IV) oxide is produced. Carbon (IV) oxide forms a white precipitate with lime water.

Example

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

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

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

Ammonium carbonate readily decomposes on heating to form ammonia, carbon (IV) oxide and water. Ammonia turns moist red litmus paper blue.

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

Action of dilute acids on carbonates and hydrogen carbonates

Activity 4.9: To investigate the action of dilute acids on carbonates and hydrogen carbonates

In groups

Apparatus and reagents

- Sodium carbonate
- Calcium carbonate
- Potassium hydrogen carbonate
- Freshly prepared lime water
- Dilute sulphuric acid
- Test tubes
- Sodium hydrogen carbonate
- Zinc carbonate
- Copper carbonate
- Dilute hydrochloric acid
- Dilute nitric acid
- Wooden splints.

Procedure

1) In separate test tubes, add a spatulaful of sodium carbonate, sodium hydrogen carbonate, potassium carbonate, potassium hydrogen carbonate, calcium carbonate, zinc carbonate and copper carbonate in 5 cm^3 of dilute hydrochloric acid.
2) Repeat the procedure using 5 cm^3 portions of dilute sulphuric acid and dilute nitric acid in place of dilute hydrochloric acid. Test for any gas produced using lime water and a burning splint.

3. Copy and complete the table below.

Action of dilute acid on	**Effect of the gas on**	
	Lime Water	**Burning splint**
Sodium carbonate		
Sodium hydrogen carbonate		
Zinc carbonate		
Copper carbonate		
Potassium carbonate		
Potassium carbonate		
Potassium hydrogen carbonate		
Calcium carbonate		

The facts

Dilute acids reacts with carbonate and hydrogen carbonates to form a salt, carbon (IV) oxide and water.

Example

$Na_2CO_3 + 2HCl\ (aq) \longrightarrow 2Nacl\ (aq) + CO_2\ (g) + H_2O\ (I)$

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

$ZnCO_3\ (s) + 2HCl\ (aq) \longrightarrow ZnCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$

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

Carbon (IV) oxide forms a white precipitate with lime water and also extinguishes a burning splint.

Check your progress 4 .5

1. Name the carbonates that do not decompose on heating.
2. Calcium carbonate reacts with dilute sulphuric acid for a few seconds and the reaction eventually stops. Why do you think this is so?
3. Dry carbon (IV) oxide was bubble through freshly prepared calcium hydroxide solution (lime water) for two hours. What would be the expected observations after five minutes and at the end of the experiment? Use chemical equations to support your answer.
4. What will you observe when you heat the carbonates below:

   (a) Na_2CO_3, $ZnCO_3$ and $KHCO_3$

(b) Write chemicals equations for the reaction.

5. What is the effect of adding carbonates to dilute acids?

4.6 Large scale manufacture of sodium carbonate

Activity 4.10 Investigating uses of Sodium carbonate

In pairs

1. Look at the paper making of your exercise book. Think about how these papers were manufactured.
2. Do you think treatment of drinking water is necessary? Is there anything that will happen if we fail to treat water meant for drinking?
3. How is water treated in a water treatment plant?
4. Your teacher will take you for a trip in a water treatment plant. Find out the whole process of water treatment.

Fig 4.13: Water treatment plant

The facts

In 1872, Ernest Solvay developed a process of manufacturing sodium carbonate in large scale called Solvay process

Activity 4.11 Simple demonstration of Solvay process in a school laboratory

Work in pairs

Reagents and apparatus

- Concentrated solutions of ammonia
- Sodium chloride
- Boiling tube
- Dry carbon (IV) oxide

Procedure

1) Mix equal volumes of concentrated ammonia solution and sodium chloride in a boiling tube.
2) Bubble dry carbon (IV) oxide through the mixture. Note any colour changes.

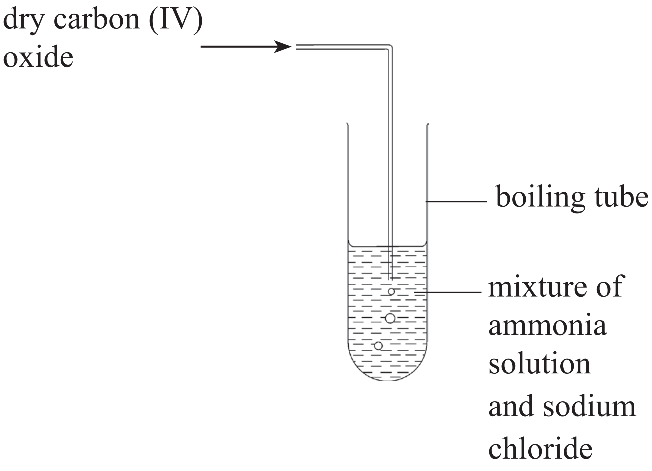


Fig 4.14: Demonstration of Solvay process

The facts

A white suspension will form when carbon (IV) oxide is bubbled through the mixture. As a result of the formation of sodium hydrogen carbonate.

The Solvay process

Raw materials:

1. Lime stone ($CaCO_3$) 2. Brine (Concentrated Nacl)

3. Ammonia from Harber process 4. Water 5. Coke

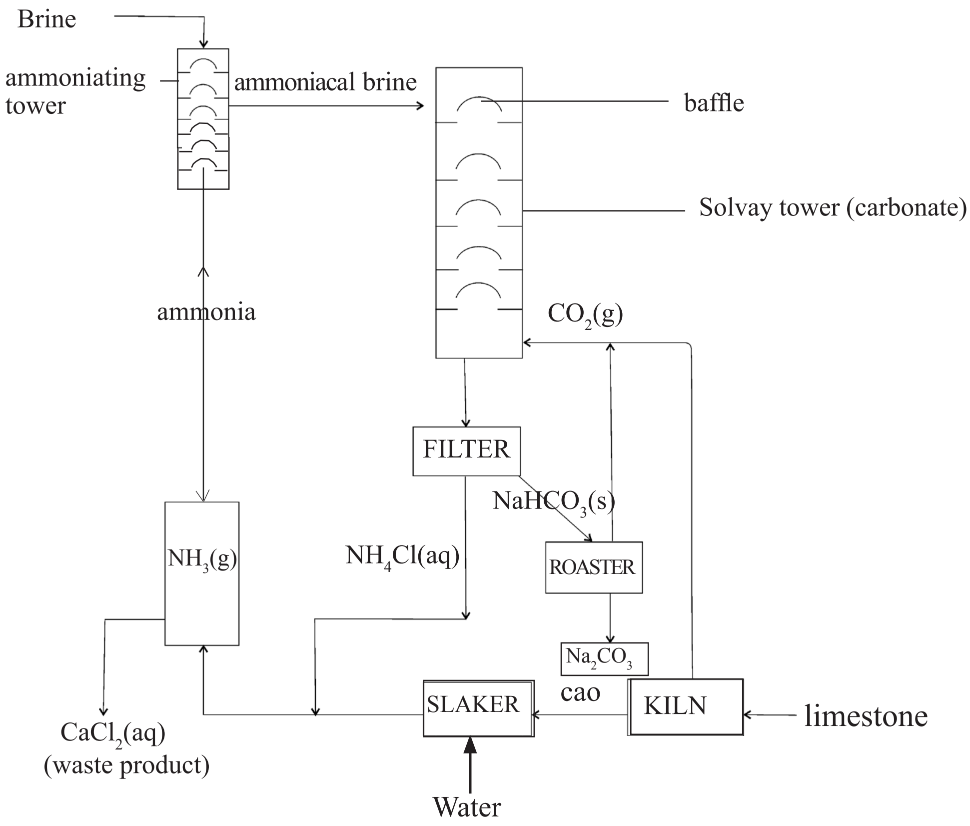


Fig 4.15: The Solvay process

A mixture of brine and ammonia is fed from the top of the Solvay tower (Ammoniating tower). In the Solvay tower, brine is saturated with ammonia gas from the generator. The process generates a lot of heat, hence the need to cool the tower from time to time.

The Ammoniated (ammoniacal) brine then flows down the Solvay tower into the carbonator. Carbon (IV) oxide from the kiln is pumped from the bottom of the carbonator where it reacts with the ammoniacal brine.

In the carbonator, the following reaction takes place.

$$NaCl\ (aq) + CO_2\ (g) + NH_3\ (g) + H_2O\ (l) \longrightarrow NaHCO_3(s) + NH_4Cl\ (aq)$$

$NaHCO_3$ forms as a suspension where it is allowed to settle and separated from ammonium chloride through **filtration**.

In the kiln, **thermal decomposition** of limestone and burning of coke takes place. The two processes produces carbon (IV) oxide.

$CaCO_3\ (s) \xrightarrow{Heat} CaO\ (s) + CO_2\ (aq)$

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

Calcium oxide is then reacted with water in the slaker to form calcium hydroxide (Slaked lime)

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

Sodium Hydrogen carbonate is then heated in the roaster to obtain sodium carbonate

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

Carbon (IV) oxide produced is recycled back into the carbonator.

Calcium hydroxide from the slaker is reacted with ammonium chloride from the carbonator to regenerate ammonia which is recycled back into the ammoniating tower.

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

Calcium chloride is the only by product which is not recycled during the process. It is mainly used us a drying agent for gases and also when added to snow it lowers in melting point.

Reactions in the Solvay chambers produce a lot of heat. For this reason, Solvay plants are usually located near water bodies to facilitate cooling of the hot chambers.

Solvay process is very economical since most of the by- products such as carbon (IV) oxide and ammonia are recycled.

Potassium carbonate cannot be manufactured using the same process. This is because potassium hydrogen carbonate and ammonium chloride have almost the same solubility in water making it impossible to separate the two by filtration.

Uses of sodium carbonate

1. It is used in the manufacture of glass and ceramic.
2. It is used in water and sewage treatment.
3. It is used in the manufacture of detergents.
4. It is used in paper and textile industries

Uses of sodium hydrogen carbonate

1. Manufacture of baking powder
2. Manufacture of carbonated drinks.

Most of the sodium carbonate is obtained from a mineral called **trona**- $Na_2CO_3 . NaHCO_3 . 2H_2O$ which is cheaply available in various parts of the world. Such as Magadi in Kenya.

The carbon cycle

Work to do

Consider the flow of money in the economy. People earn money in form of salaries and commissions and then spend that on goods and services. The process of earning and spending ensures that there is no too much supply of money in the economy which may lead to inflation and there is no too little to spend. Certain processes add money to the economy while others take money from the economy. This is called the money cycle.

The facts

Carbon compounds in humans play a very significant role in the ecosystem. Processes such as respiration add carbon (IV) oxide into the atmosphere whereas processes such as photosynthesis removes carbon (IV) oxide from the atmosphere. These processes ensures that the level of carbon (IV) Oxide in the atmosphere is kept at a constant over ages. Some of the processes through which carbon and carbon compound passes are as shown below.

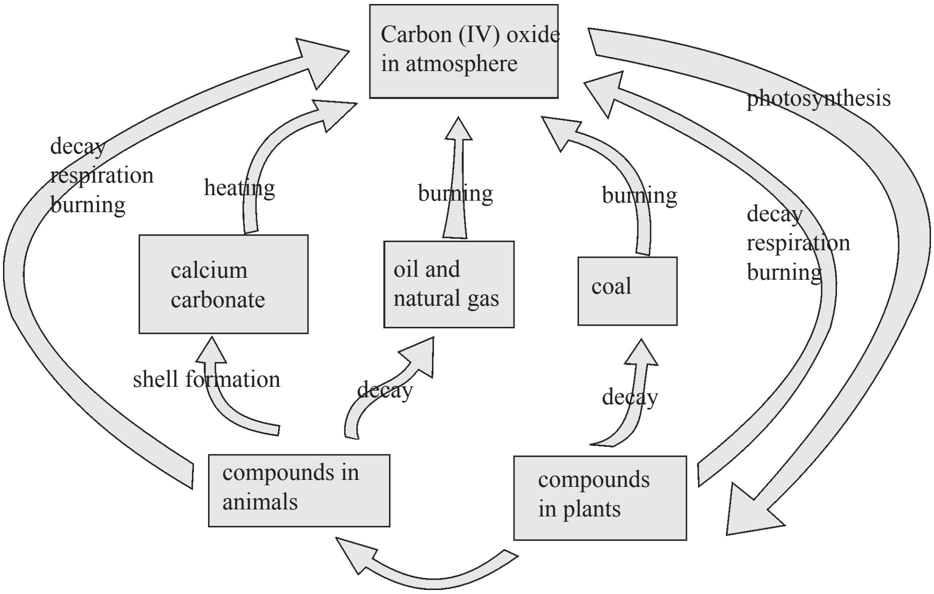


Fig 4.16:The carbon cycle

My environment my life

High levels of carbon (IV) oxide in the atmosphere may cause global warming. Plant more trees to ensure a clean environment.

Significance of carbon (IV) oxide and carbon compounds in the eco- system

1. Plants use carbon (IV) oxide during photosynthesis to form glucose. When animals feed on plants, glucose is converted into other essential substances such as proteins and starch through metabolism.
2. Carbon (IV) oxide in the atmosphere dissolves in rain water to from Carbonate ions (CO_3^{2-}). Some sea animals combines these carbonate ions with calcium from rocks to form calcium carbonate which they use to make shells.
3. Decayed plant leaves and other organic matter form humus with the soil. Humus is highly nutritious and rich in minerals needed for plants growth. The structure of humus enables it to act as a buffer against soils that are too acidic or to alkaline.

Fig 4.17: Decayed plant leaves and organic matter form humus

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

Industrial processes releases carbon (IV) oxide in the atmosphere. Increase in the amounts of carbon (IV) oxide in the atmosphere may cause;

Greenhouse effect which may lead to global warming. Greenhouse gases such as carbon (IV) oxide forms a blanket in the atmosphere. Sunlight in form of ultra-violet radiations passes the atmosphere and warms the earth. Some of the radiations are transformed into heat which is reflected back into the space. As this heat goes up the atmosphere, some of it is trapped by the greenhouse gases. This will result in excess warming of the earth.

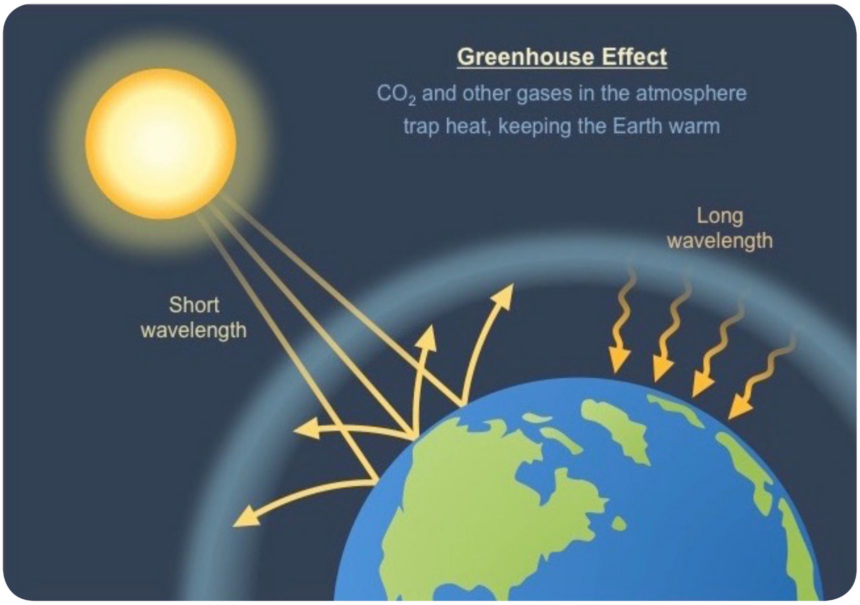


Fig 4.18: Greenhouse effect

1. The high temperatures causes the melting of ice caps in the poles and mountains and this may result in rising in the level of the oceans and flooding. Global warming also causes change in the climatic patterns among other disasters.
2. Car engines releases some reasonable amounts of carbon (II) oxide. Carbon (II) oxide may lead to suffocation and even death when inhaled. Car engines should be fitted with catalytic converters to convert carbon (II) oxide into less toxic carbon (IV) Oxide.

Greenhouse effect

When solar radiations reach the earth's atmosphere,

i. Some are reflected back into the space.

ii. Some energy is absorbed by the earth therefore heating it.

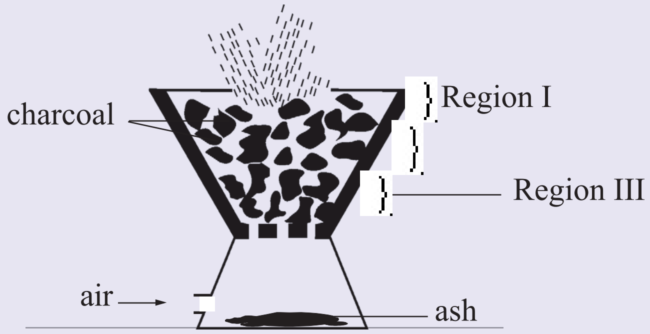
iii. Some of this heat is reflected back to the space.

iv. Greenhouse gases traps some of the reflected heat.

v. Extra heat trapped causes the earth's temperature to rise (global warming).

Check your progress 4 .6

1. The diagram below shows a charcoal stove in use. Study it and answer the questions that follows.



   (a) Identify the gas formed at region I.

   (b) State and explain the observation made at region III.

2. Diamond and graphite are allotropes of carbon

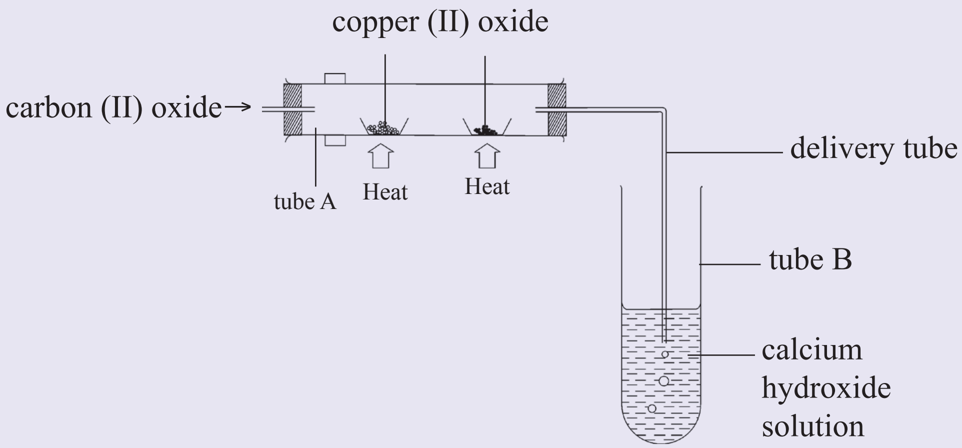
   (a) What are allotropes?

   (b) Which of the allotropes conducts electricity? In terms of its structure, explain why it conducts electricity.

3. Diamond is used in cutting of metals. In terms of its structure, explain how this is possible.

4. Name some uses of diamond and graphite.

5. A student wanted to investigate a property of carbon II oxide. He set up the apparatus as shown below ;

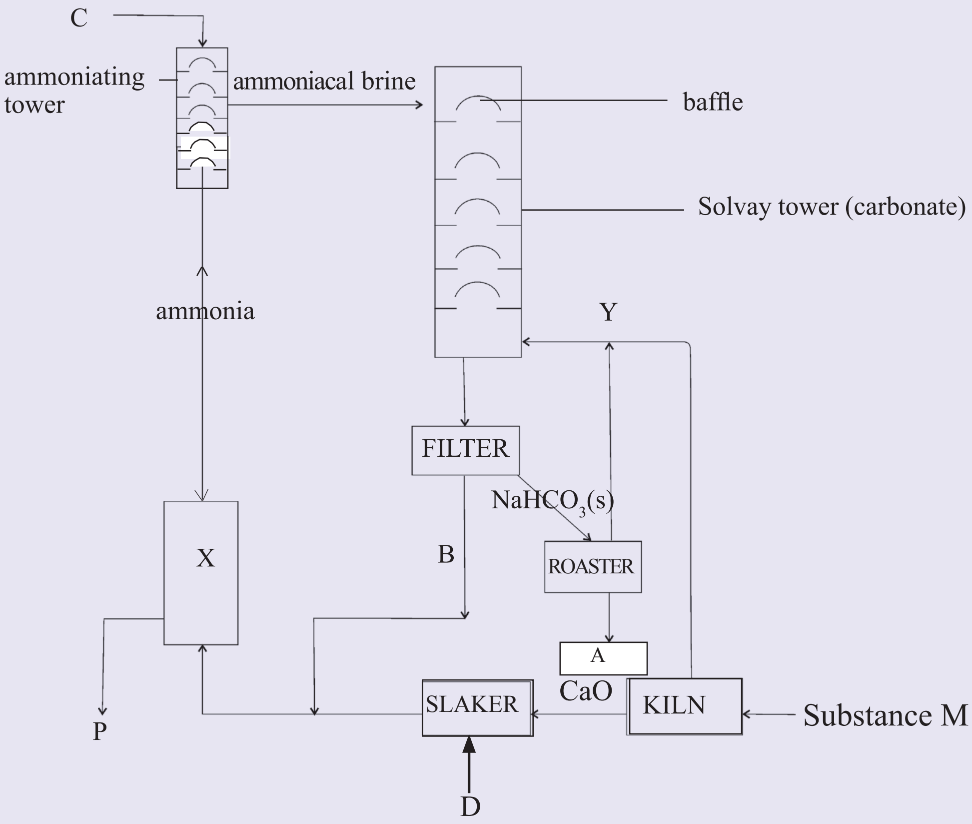


(a) What will be observed in tube B?

(b) Write an equation for the reaction that takes place in Tube A.

(c) Give one use of carbon II oxide.

6. Solid carbon (IV) oxide is preffered to ice as a cooling agent. Give a reason.
7. A certain colourless gas was suspected to be carbon (IV) oxide. Suggests a chemical test that can confirm its identity.
8. The flow chart diagram below shows some of the processes involved in the manufacture of sodium carbonate by Solvay process. Study it and answer the questions that follows.



(a) Identify the substances labelled A, B, C and D.

(b) Name Gas X and Gas Y

(c) State one use of the by-product P.

(d) Identify substances M, hence write equations for the two reactions taking place in the kiln

(e) Solvay tower is cooled from time to time give a reason.

(f) Give two uses of sodium carbonate.

(g) Potassium hydrogen carbonate and ammonium chloride cannot be prepared by the same method. Give a reason

9. If two unlabelled black powders were placed before you on the laboratory bench and you are told that they are copper (II) oxide and charcoal, what simple test would you perform to distiguish them?
10. A student mixed up the labels of solid sodium carbonate and solid sodium hydrogencarbonate. How did she go about differentiating the two solids?
11. Explain the following:
    (a) Carbon (IV) oxide is one of the causes of greenhouse effect.
    (b) Diamond and methane have the same type of bonding, yet methane is a gas at room temperature and pressure, while diamond is a very hard solid.
    (c) The reactiom between dilute sulphuric acid and barium carbonate is not suitable for preparing carbon (IV) oxide.
12. Carbon occurs in both pure and impure forms. Give examples of both pure and impure forms of carbon.
13. Global warming can be controlled using the carbon cycle. Explain how this is possible.

Glossary

A

Absorption:	Soaking up of one substance through the entire mass of another.
Acid salt:	A compound or salt from an acid in which not all the hydrogens have been replaced by a metal. eg. $NaHCO_3$, $NaHSO_4$ etc.
Active substance:	A substance which reacts vigorously with other substances, e.g., potassium which reacts vigorously with water.
Adsorption:	The acquisition of one substance, e.g. a gas by the surface of another, e.g. solid surface.
Allotropes:	Different forms of the same element in the same state, e.g. diamond and graphite is one allotrope of carbon.
Amphoteric:	Capable of acting either as an acid or as a base, e.g. aluminium oxide is an amphoteric oxide.
Amphoteric oxide:	A metal oxide which reacts with an acid or a base producing salts in both cases, e.g. aluminium oxide.
Anhydrous:	Containing no water of crystallisation.
Anion:	A negatively charged ion, e.g. chloride ion (Cl^-) ion.
Anode:	The positively charged electrode. It is joined to the positive terminal of the battery/cell. The anions move towards it during electrolysis process.
Atomic number:	The number of protons in an atom, e.g. the atomic number of carbon is 6 because it has six protons.

B

Baking soda:	Sodium hydrogencarbonate ($NaHCO_3$).
Basicity:	The number of hydrogen atoms in each molecule of an acid replaceable directly or indirectly by a metal or an ammonium radical, e.g. the basicity of sulphuric acid (H_2SO_4) is 2.

Binary compound:	A compound consisting of only two elements.
Bleaching:	The process by which colour is partially or wholly removed from a coloured substance.
Bleaching powder:	A chlorine containing substance used in bleaching textiles or sterilising water. It is made of slaked lime ($Ca(OH)_2$) and chlorine.
Boiling point:	The temperature at which equilibrium vapour of a liquid is equal to prevailing atmospheric pressure. The lower the atmospheric pressure, the lower the boiling temperature of a liquid at which the equilibrium vapour pressure are equal.

C

Carbon:	A non-metallic element of atomic number 6 and found in group 4 and period 2 of the periodic table.
Carbon cycle:	The process of removing and replacing of carbon(IV) oxide in the air thus keeping its percentage in air constant.
Carbon (II) oxide:	A colourless, odourless, tasteless, neutral and poisonous gas formed when carbon, hydrocarbons or organic matter burns in insufficient supply of air/oxygen. It burns with a blue flame.
Carbon (IV) oxide:	A heavy colourless and odourless gas which forms a white precipitate with calcium hydroxide solution.
Cathode:	A negatively charged electrode. It is joined to the negative terminal of the battery/cell. The cations move towards it during electrolysis process.
Cation:	A positively charged ion, e.g. Sodium ion (Na^+).
Charged particle:	A particle with either negative or positive charge, e.g. electron (–), proton(+), chloride ion (Cl^-), sodium ion (Na^+).

Chemical bond:	A force which holds particles (atoms, ions or molecules) together.
Chemical equation:	A statement for the overall results of a chemical change and represented by symbols or formulae of the reactants and products, e.g. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$
Chemical formula:	A shorthand notation using chemical symbols and numerical subscripts to represent the composition of substances,e.g. $Na_2\ CO_3$ (Sodium carbonate).
Chemical properties:	Those properties which pertain to the behaviour of a material in changes in which its identity is altered completely.
Combustion:	Any chemical reaction that occurs so rapidly that noticeable heat and light are produced, e.g. reaction between hydrogen and oxygen.
Conductor:	A substance which allows flow of an electric current, e.g.all metals and graphite. It can be in solid or molten form.
Covalent bond:	Bonding in which two atoms share a pair of electrons, e.g., the bond between hydrogen and chlorine atoms.

D

Dative bond:	Bonding in which two atoms share a pair of electrons. The pair of electrons are donated by one atom. The other atom only accepts the sharing. It is also called *co-ordinate bond*.
Dehydrating agent:	A substance which removes water vapour from another substance e.g., concentrated sulphuric acid.

E

Effervescence:	The process of giving off bubbes of a gas from a liquid in which it is dissolved or by reacting two reacting substances, e.g; reaction of calcium carbonate and dilute hydrochloric acid.
Electric current:	A flow of electrons from one point to another.
Electrode:	A metal or carbon rod (graphite) used to conduct electricity into (or out of) an electrolyte.
Electrolysis:	The decomposition of a compound in solution or molten form by electricity. D.C is used.
Electron:	A negatively charged particle revolving around the nucleus of an atom. It has negligible mass.
Electrolyte:	A substance which conducts electricity when in solution or molten form, e.g. aqueous or molten sodium chloride.
Electronegative element:	An element having a relatively strong attraction of electron of an atom within a molecule.
Endothermic:	A chemical recaction is which heat energy is absorbed.
Electronegativity:	The tendency of an atom to attract the electron(s) from another atom forming a chemical bond between it and that atom.
Electroplating:	The process of coating the surface of metallic surface with another metal using electrolysis process, e.g. elecroplating iron with tin.
Electropositive element:	An element having a relatively weak attraction for its outermost energy level electron(s) or outermost energy level electron(s) of other elements.
Energy level:	A region outside the nucleus of an atom in which electrons may be considered to circulate.
Exothermic:	Pertaining to a chemical reaction which occurs with the evolution of heat energy.

G

Group:	A vertical column of elements in the periodic table, e.g. Group I having, Li, Na, K, Rb, Cs and Fr.

H

Halogen:	Name given to the family of elements having seven electrons in the outermost energy level of their atoms. i.e group VII
Homogeneous substance:	A substance which has similar properties throughout.
Hydrate:	A crystalline substance that contains water of crystallisation, e.g. $CuSO_4 . 5H_2O$
Hydration:	The chemical attachment of water molecules to molecules of the solute, i.e., the process of adding water of crystallisation to a molecule.
Hydrogen bond:	A weak chemical bond between a hydrogen atom in one polar molecule and a negative atom in a second polar molecule e.g.

$$H^{\overset{+}{\partial}} \text{———} Cl^{\overset{-}{\partial}} \text{ --------- } H^{\overset{+}{\partial}} \text{———} Cl^{\overset{-}{\partial}}$$

Hydrolysis:	The reaction of a salt with water to form a solution which is acidic or basic, e.g. hydrolysis of aluminium chloride to form aluminium hydroxide and hydrogen chloride gas.
Hydroxonium ion:	A hydrated proton, i.e. H_3O^+ ion.

I

Inactive substance:	A substance which reacts passively with other substances e.g. nitrogen reacts with air only at very high temperature.
Inert substance:	A substance which does not react with other substances under the ordinary conditions of chemical reactions e.g. argon.

Insulator:	A non-conductor of electricity used to protect something from electricity e.g., plastic.
Insoluble:	Incapable of being dissolved.
Ion:	An atom or a group of atoms which possesses one or more positive or negative charges, e.g., Na^+, Cl^-, NO_3^-, SO_4^{2-}, Ca^{2+}, S^{2-} etc.
Ionic bonding:	Bonding in which one or more electrons are transferred from one atom to another.
Isotope:	One of two or more forms of atoms with the same atomic number but with different mass number which comes as a result of different number of neutrons in the nucleus e.g. $^{35}_{17}Cl$ and $^{37}_{17}Cl$
Isotopy:	The existence of elements in form of isotopes.

L

Lime water:	Calcium hydroxide solution, it used for testing carbon(IV) dioxide.
Litmus:	A dye extracted from lichen, which is used as an indicator. In the laboratory it is either in solution, crystals or soaked and dried into filter paper, usually known as litmus paper.

M

Mass:	The quantity of matter which a body possesses.
Mass number:	The number of protons plus the number of neutrons in an atom of an element.

N

Neutral particle:	A particle with no charge, e.g. neutron, atom.
Neutron:	A neutral particle found in the nucleus of an atom. It has a unit mass of one and no charge.

Non-conductor:	A solid which does not allow an electric current to pass through e.g. wood, plastic, etc.
Non-electrolyte:	A substance which does not conduct electricity in solution or molten form, e.g. wax, benzene.
Normal salt:	A compound consisting of metal or ammonium radical and non-metallic radical derived from an acid. The non-metallic radical is incapable of further yielding more hydrogen. e.g. sodium carbonate (Na_2 CO_3).
Nucleus:	The positively charged, dense and central part of an atom. It consists of protons($^+$), neutrons (no charge) and protons($^+$).

P

Period:	Horizontal row of elements in the periodic table, e.g. period 3 - containing, Na, Mg, AL, Si, Pi, S, Cl and Ar.
Periodic table:	A tabular arrangement of the chemical elements based on their atomic structure and number. Currently, the periodic table has 119 elements.
Physical properties:	These are properties which can be determined without causing a change in the identity of a substance e.g. density.
Polar molecule:	A molecule which has regions of positive and negative charge caused by unequal attraction of the shared electrons, e.g. $H^{\partial+}$ — $Cl^{\partial-}$
Precipitate:	A substance usually solid which separates from a solution as a result of some chemical or physical change e.g. lime water forms a white precipitate of calcium carbonate when reacted with carbon(IV) oxide. The white precipitate is calcium carbonate because it is insoluble in water.
Product:	An element or compound resulting from a chemical reaction e.g. the product of the reaction between hydrogen and oxygen is water.
Proton:	A positively charged particle found in the nucleus of an atom. It has one unit mass and one positive charge.

R

Radical:	A group of atoms which usually behaves like it were a single atom, e.g. SO^{2-}_4, NH^+_4.
Reactant:	An element or compound entering in a chemical reaction, e.g., hydrogen and oxygen are the reactants which form water.

S

Salt:	A compound consisting of metal or ammonium radical and non-metallic radical derived from an acid, e.g. sodium carbonate (Na_2CO_3), sodium hydrogencarbonate ($NaHCO_3$).
Soluble:	Capable of being dissolved.
Sublimation:	The process in which a solid changes to vapour or vapour to solid, without first forming liquid, e.g. iodine, ammonium chloride, solid carbon(IV) oxide.
Synthesis:	The process of combining of simple substances to make a more complex substance, e.g. synthesis of iron(III) chloride from iron and chlorine.

T

Triatomic molecule:	A molecule containing three atoms e.g. ozone (O_3), water (H_2O).
Trona:	A double salt of sodium carbonate and sodium hydrogencarbonate. It contains sodium chloride dissolved in its water of crystallisation.

V

Valence/ valency:	The number of electrons gained, lost or shared by an atom in chemical bonding with one or more atoms, e.g. The valency of iron(III) ion(Fe^{3+}) = 3 and chlorine ion (Cl^-) = 1. It is also the number of hydrogen atoms that an atom of an element or a radical can combine with, e.g., the valence of sulphate ion (SO_4^{2-}) = 2, nitrate ion (NO^-_3) = 1.

Valence or valency electrons:	Electrons in the unfilled outermost energy level.
Van der Waal's bond:	A comparatively weak bond (force) which holds molecules together in molecular crystals or liquids. The crystals have a low melting point and the liquids a low boiling point.

W

Water gas:	A fuel gas which is a mixture of carbon(II) oxide and hydrogen. It is made by allowing a blast of steam through a bed of red-hot coke.
Water of crystallisation:	The water present in the structure of the crystals in fixed amount, e.g. hydrated copper(II) sulphate ($CuSO_4.5H_2O$) has five molecules of water of crystallisation.
Word equation:	A brief statement which identifies the reactants in a chemical reaction and the products formed, e.g., Hydrogen + oxygen → water.

APPENDIX I
Periodic table

x Z y Name

x is the Relative Atomic Mass (RAM)
Z is the symbol of the element
y is the Atomic Numbers of the element

Groups

Periods	I	II											III	IV	V	VI	VII	O
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	54 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 Unununium	285 Uub 112 Ununbium	Uut 113	289 Uuq 114 unnuqnabium	Uup 115	292 Uuh 116 ununhexium	Uns 117	Uuo 118 ununoctium

Lanthanoid elements	257 La 57 Lanthanum	140 Ce 58 Cerium	141 Pr 59 Praseodymium	144 Nd 60 Neodymium	147 Pm 61 Promethium	150 Sm 62 Samarium	152 Eu 63 Europium	157 Gd 64 Gadolinium	159 Tb 65 Terbium	163 Dy 66 Dysprosium	165 Ho 67 Holmium	167 Er 68 Erbium	169 Tm 69 Thulium	173 Yb 70 Ytterbium
Actinoid elements	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

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