UNIT 1: ELEMENTS AND THE PERIODIC TABLE

In form 1, you were introduced to the periodic Table of elements. You learnt that the Periodic Table is a grid made up of columns called groups and rows called periods where elements are arranged in a systematic manner. You also learnt about the various families of elements in the periodic table.

In this class, you will learn about the physical and chemical properties of elements and their compounds in relation to the number of electronics that are found in their outermost energy levels. When elements are arranged in the Periodic Table in order of increasing atomic number, a regular change in the outermost electron arrangement is observed. In turn, a regular variation of properties is also observed.

This brings about gradual change in the physical and chemical properties of the elements. In order to illustrate these changes, we will consider a number of trends and patterns observed in the Periodic Table.

1.1 Blocks of elements in the periodic table

Most elements in the Periodic table are metals while others are non-metals. Some other elements show partly metallic and partly non-metallic characteristics. These elements are called semi-metals or metalloids. Table 1.1 shows the various blocks of elements in the Periodic table.

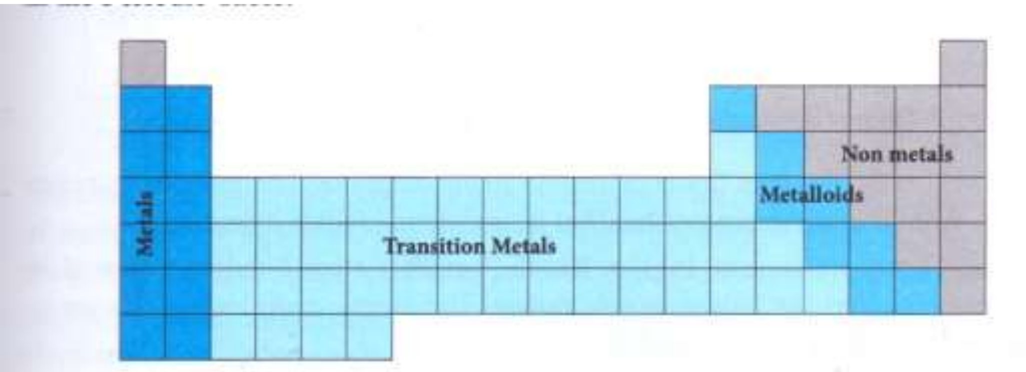


Table 1.1: Classification of elements in the Periodic Table

1.2 The general trends in the periodic table

We will study the general trends in the following properties of elements:

- Atomic radius
- Electron affinity
- Ionization energy
- Electronegativity

a. **Atomic radius**

The atomic radius is the distance between the nucleus and the outermost energy level in an atom. See fig 1.1. The more the number of energy levels, the greater the atomic radius and vice versa. Also, shielding effect by the electrons on the positive nucleus affects the size of the atom.

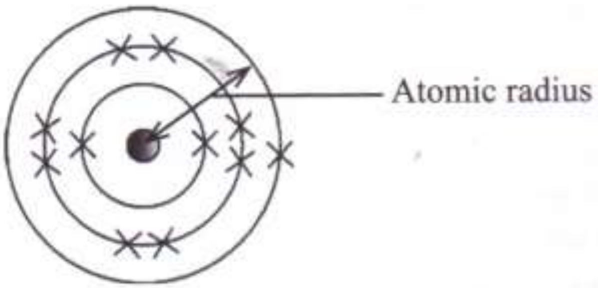


Fig. 1.1: Atomic radius of sodium

Consider lithium (Li), sodium (Na) and potassium (K) with 3, 11 and 19 electrons respectively. Their electronic configurations and atomic structures will be as follows.

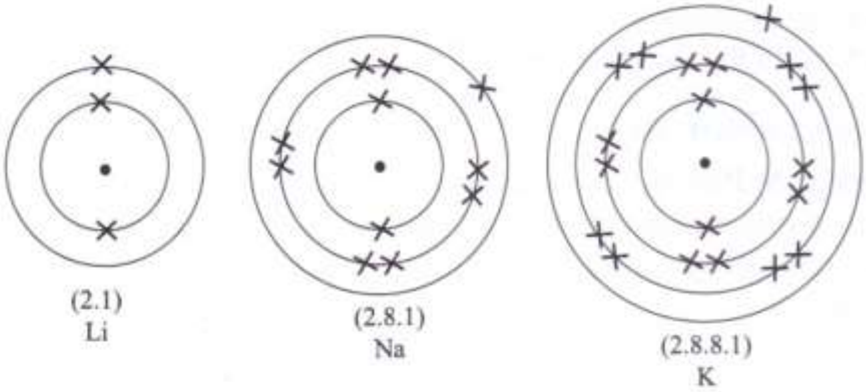


Fig. 1.2: Electron arrangement in Group I elements; lithium, sodium and potassium

You will note that the distance from the nucleus to the outer most energy level increases as we move from lithium to potassium. That is, the atomic radius increases.

Of the three Group 1 elements, lithium has the smallest atomic radius followed by sodium. Potassium has the largest atomic radius. The atomic radii increase as we go down the group as illustrated in gig. 1.3

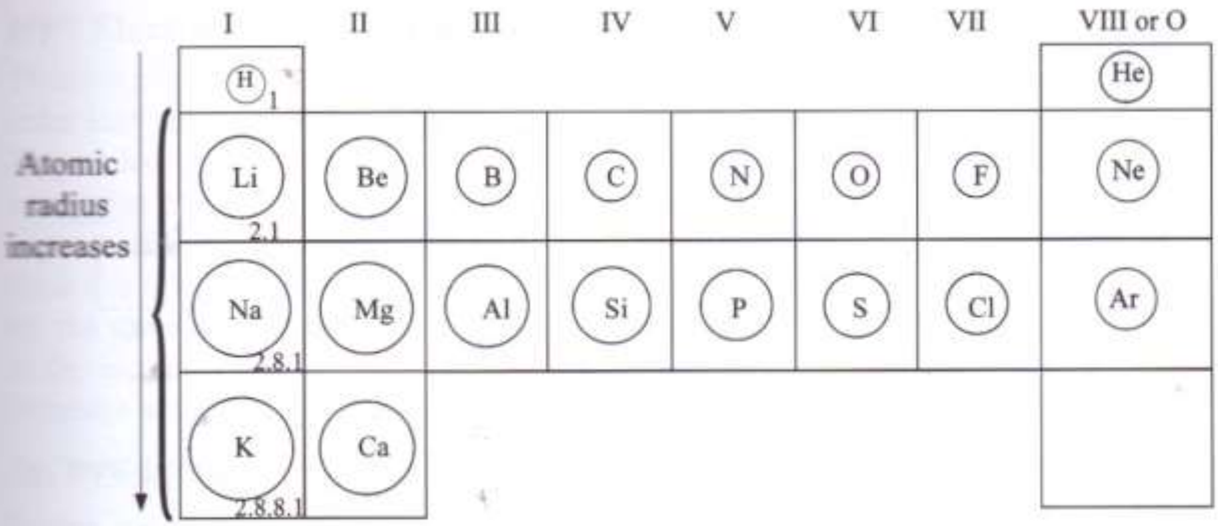


Fig. 1.3: Variation of atomic radii of the first twenty elements of the Periodic Table

Why do you think this is the case? Before we answer this question, let us first look at Fig. 1.4.

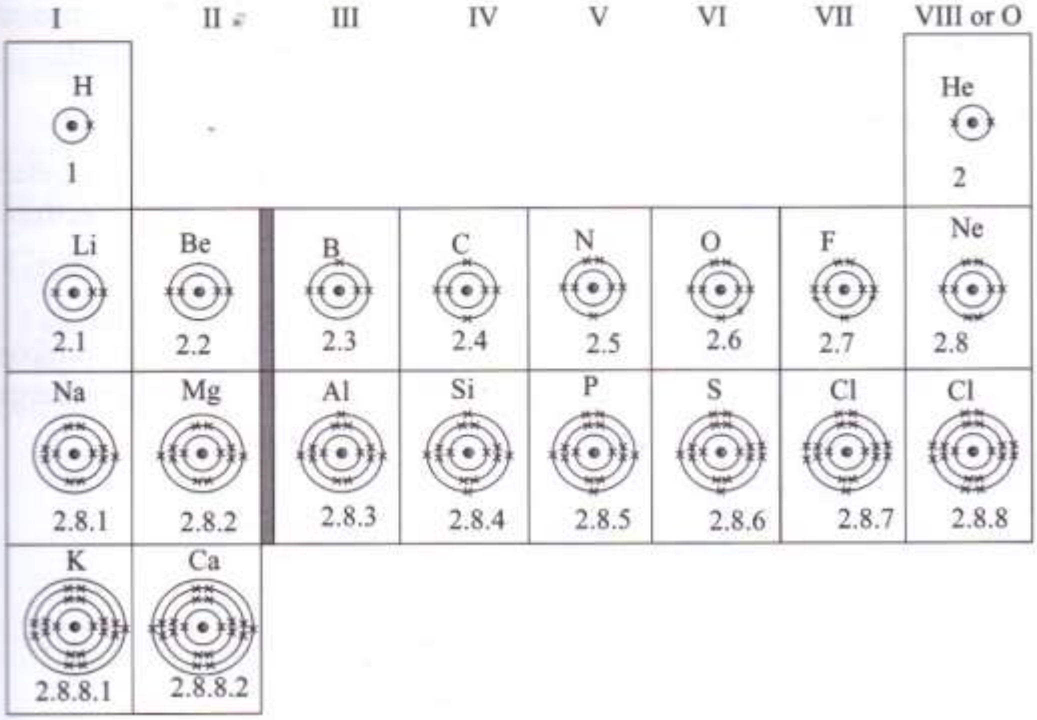


Fig. 1.4: Electron arrangement and structure of the first 20 elements of the Periodic Table

The atomic radii increase down the group because as we go down the group, the number of energy levels increases. The number of electrons and protons also increases down the group. Increase in the number of protons leads to increased nucleus charge.

Although the nuclear charge increases due to the increase in the number of protons, the electrons in the inner energy levels shield those in the outermost energy level from full attraction by the nucleus. This is called shielding effect.

In order to understand the shielding effect better, look at Fig 1.5 below

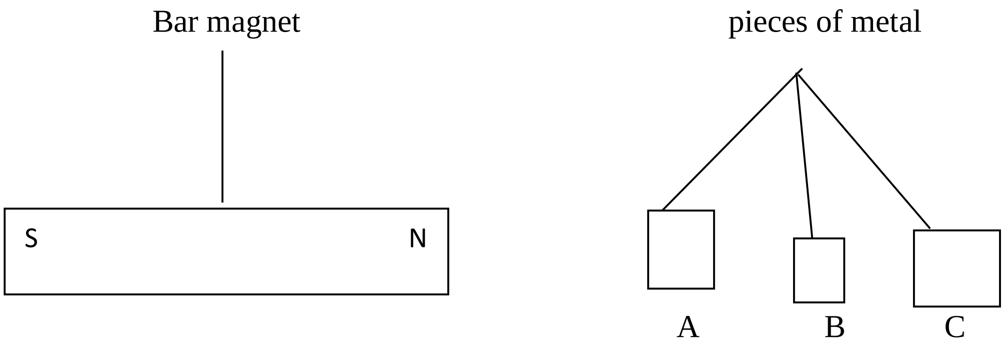


Figure 1.5

The magnet attracts metal A more than B or C. metal A is shielding the total "feel" of attraction for B and C.

Electrons in the inner energy levels behave in a similar manner. They shield the outermost electrons from the full feel of nuclear attraction. The size of an atom depends on how strongly the positive protons in the nucleus attract the electrons in the outermost energy level.

Shielding effect or screening effect to the reduced attraction by the nucleus for the outermost electrons caused by the electrons within the inner energy levels.

b. Ionization energy

Ionization energy is that energy that must be absorbed in order to remove the outermost electron from an atom. It is a measure of how difficult it is to remove an electron from an atom in gaseous state.

$$M\,(g) \longrightarrow M^{+}\,(g) + e^{-}$$

The higher the ionization energy the more the difficult it is to remove an electron from the atom. Look at the table below. What can you conclude about the ionization energies of the elements? How does it relate to the atomic size of the elements?

Table 1.2 ionization energies of some elements

Element	Ionization Energy (K)	Atomic radius

Lithium (Li)	520	0.15um
Sodium (Na)	496	0.19um
Potassium (K)	479	0.23um

Notice that the ionization energy decreases as you move down the group while atomic radius increases. This means that it is easier to remove an electron from the atom as we go down the group I. the smaller the atom the higher the ionization energy. Electrons are more attracted to the nucleus in a smaller atom. It becomes more difficulty to remove the electrons (s) from such atoms. Potassium has the lowest ionization energy because the outermost electron is far away from the nucleus that is it has the biggest atomic radius and therefore it is relatively easier to remove.

c. Electron affinity

Electronic affinity is concerned with the attraction of Electrons. An electron trying to enter into an atom will be repelled by the electrons in the outermost energy level. The nucleus of an atom with a small radius will attract the incoming electron more easily than one which has a bigger atomic radius. Since the atomic radii increase down a group of a periodic Table, electron affinity decreases down the group. Why do you think this is so? However, across the period, the atomic radii decreases because electrons fill the same energy level and the nuclear charge (protons) increases. The attraction of the incoming electron also increases across the period. Therefore electron affinity increases across a period.

d. Electronegativity

Earlier, we learnt that when two atoms of non-metallic atoms bond, they share a pair of electrons to form a molecule; each atom contributes on electron. But in molecules, instead of sharing the pair of electrons equally one atom tends to attract the pair of other atom slightly positive. This electron-attracting power of an atom is known as electronegativity.

Electronegativity increases across a period and decreases down the group of the periodic Table.

Let us now look at the trends in the physical and chemical properties of the various families of elements.

1.3 Group I Elements-Alkali Metals

Group I elements are also known as alkali-metals. They are the most reactive group of metals. They are found on the extreme left hand side of the Periodic Table as illustrated in fig 1.6. Members of this group include Lithium (Li), Sodium (Na) and potassium (K).

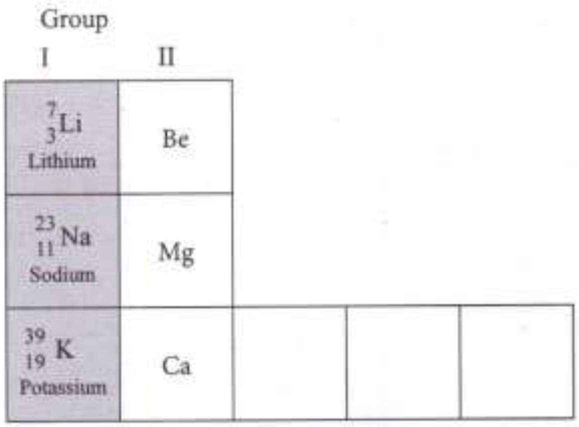


Fig. 1.6: Alkali metals

Trends in physical properties of alkali metals

Some of the physical properties to be studied in this section include:

- Melting and boiling points
- Physical appearance (colour, texture, malleability
- Thermal conductivity
- Electrical conductivity
- Density among others

Experiment 1.1

Aim

To investigate the physical appearance of alkali metals

Apparatus and chemicals

- Tile (ceramic)
- Filter paper
- Knife
- Pair of tongs
- Lithium
- Sodium
- Potassium

Procedure

1. Using a pair of tongs, remove a small piece of lithium from the bottle.
2. Place it on the tile and observe

- Describe the apparatus of the lithium from the bottle.

3. Cut the piece of lithium into two using a knife as shown in fig 1.7 to expose the inside.
    - What is the colour of the freshly cut part of lithium?

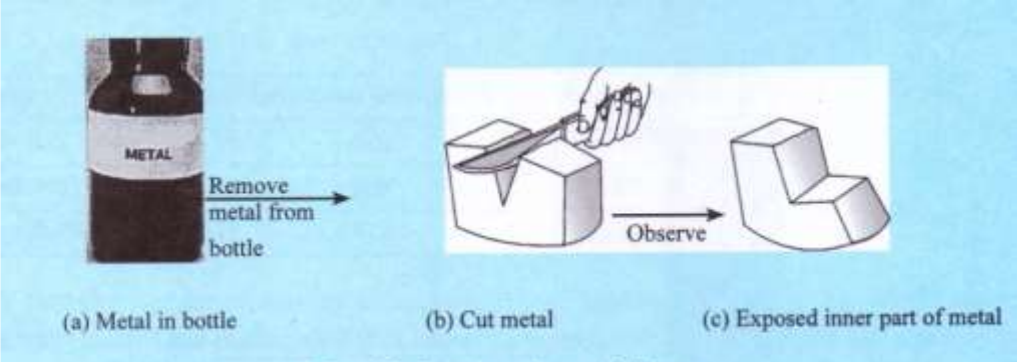


Fig. 1.7: Cutting a piece of lithium

4. Repeat the above procedure with sodium and potassium and answer the same questions.
5. Record your observations in a table like the one shown below.

Table 1.5 physical properties of alkali metals

	Appearance of metal form bottle	**Appearance of freshly cut metal**
Lithium		
Sodium		
Potassium		

- What do you conclude about the appearance of Group I elements?
- Why does the appearance of the metal change after a short while?

We saw in previous section that alkali metals have one electron in the outermost energy level. As a result, alkali metals have very similar physical and chemical properties.

When alkali metals are freshly cut, they look shiny and silvery but they tarnish immediately when exposed to air.

Other characteristics

- Very low density (they float on water)
- Very soft and can be cut with an ordinary knife easily
- Good conductors of heat
- Good conductivity of electricity
- Low melting points
- High boiling points
- Shiny surface when freshly cut

Table 1.4 gives a summary of the physical properties of alkali metals

Table 1.4 physical properties of alkali metals

Element	Symbol	Physical state at room temperature	Physical appearance	Electron arrangement	Melting point	Boiling point	Thermal conductivity	Electrical conductivity
Lithium	Li	Solid	Silvery metal	2.1	181^0C	$134.\ ^0C$	Good conductor	Good conductor
Sodium	Na	Solid	Silvery metal	2.8.1	98^0C	883^0C	Good conductor	Good conductor
potassium	K	Solid	Silvery metal	2.8.8.1	63^0C	759^0C	Good conductor	Good conductor

In general, down the group:

- The melting and boiling points decrease due to an increase in atomic size hence attraction between positive nucleus and the electrons are reduced
- Atoms become bigger due to increase in number of energy levels
- Density increase due to increased mass
- Ionization energy decreases down due to increase in atomic size hence weaker force of attraction between the possible nucleus and outermost energy level electrons.
- Electron affinity reduces as the atom becomes bigger
- Electronegativity decreases due to increase in atomic size

Trends in chemical properties of Alkali metals

Some of the chemical properties that will be suited include:

- Reaction with air or oxygen
- Reaction with chlorine
- Reaction with water

a. Reaction of alkali metals with air or oxygen

Experiment 1.2

Aim

To investigate what happens when alkali metals react with air

Apparatus

- Bunsen burner
- Pair of tongs
- Knife
- Lithium

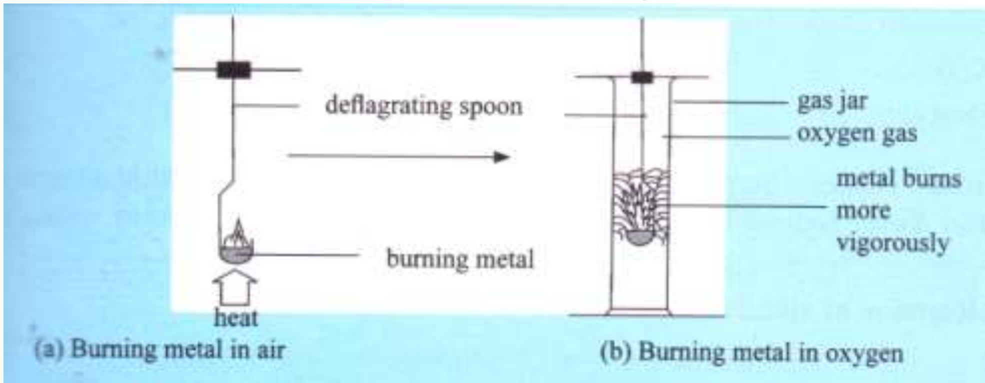
- Potassium
- Tile (ceramic)
- Filter paper
- Deflagrating spoon
- Sodium

Procedure

1. Place all the requirements on the demonstration bench
2. Using a pair of tongs, remove a small piece of lithium form the bottle
3. Place it on the ceramic tile covered with filter paper. Remove most of the oil with filter paper

Caution: do not touch the metal with bare hands

4. Cut a small piece of lithium 9about 2-3 mm3 or the size of a rice grain) and return the rest into the bottle
5. Transfer the small piece of lithium into a deflagrating spoon
6. Heat the lithium metals as illustrated in Fig 1.8



(a) Burning metal in air

(b) Burning metal in oxygen

Fig. 1.8: Reaction of alkali metals in air and oxygen

7. As lithium begins to burn, remove the spoon from the flame and observe the reaction.
8. Repeat the same procedure for sodium and potassium

9. Record the observations in your notebook.

    - What do you conclude about creativity of these metals with air (lithium, sodium and potassium)
    - Ali had an element X which he suspected to be in Group I. what would he expect to observe when burning X is inserted in a gas jar of oxygen?

Write a balanced equation for the reaction.

Group I elements react with air to form oxides and peroxides depending on the amount of oxygen available

- Lithium reacts with oxygen to form lithium oxide which is white in colour.

Lithium + oxygen → lithium oxide

$4Li(s)$ + O_2 (g) $2Li_2O$ (s)
(Silvery) (White)

- Sodium reacts with oxygen to form a mixture of oxide and peroxide

Sodium + oxygen → lithium oxide
$4Na(s)$ + O_2 (g) $2Na_2O$ (s)
(Silvery) white

Sodium + excess oxygen → sodium peroxide

$2Na$ (s) + O_2 (g) → Na_2O_2 (s)

(Silvery) (Yellow-white)

- Potassium also reacts with oxygen in air to form a mixture of oxide and peroxide.

Potassium + oxygen → lithium oxide
$4K(s)$ + O_2 (g) $2K_2O$ (s)
(Silvery) white

Potassium + excess oxygen → potassium peroxide

$2K$ (s) + O_2 (g) K_2O_2 (s)

(Silvery) (White)

All Group I elements burn with characteristic flames. For example, Lithium burns with a scarlet flame; sodium burns with yellow flame while potassium burns with a lilac flame.

Experiment 1.3

Aim

To investigate what happens when alkali metals react with water:

Apparatus and chemicals

- Glass trough
- Pair of tongs
- Tile (ceramic)
- Litmus solution or universal indicator solution
- Sodium
- Potassium
- Water

Procedure

1. Half-fill the trough on the demonstration bench with water
2. Add three drops of universal indicator solution. (You can also test the water with red an
3. d blue litmus paper instead of universal indicator solution or litmus solution)
    - What colour change do you observe? Indicate the pH of the solution as neutral, alkaline or acidic.
4. With pair of tongs, remove a small piece of lithium from the bottle and place it on the ceramic tile
5. Cut a small piece of lithium (the size of a rice grain) and return the rest to the bottle
6. Drop the small piece of lithium into the glass trough with water using a pair of tongs as shown in fig 1.9. What do you observe?

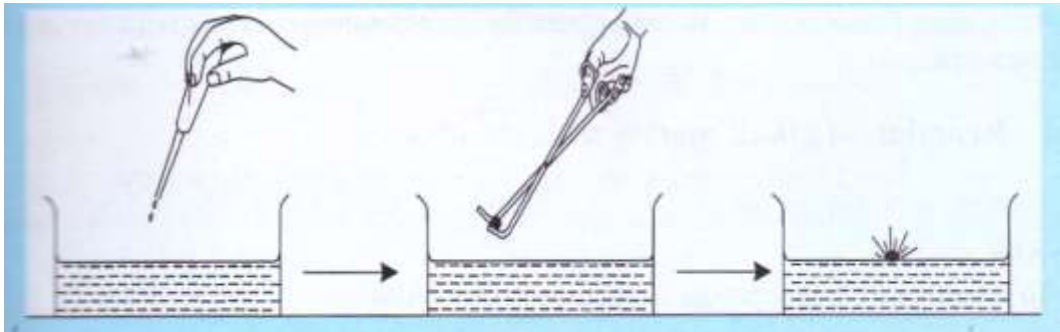


Fig. 1.9: Reaction of alkali metals with water

7. In case you used litmus paper, repeat the test with red and blue litmus papers. What do you observe?
8. Record your observations in your notebook
    - Arrange the metals in order of how reactive they are with water
    - State which metal is least reactive? Give the reason

Lithium reacts slowly with water. It does not melt because the reaction produces less heat that is not sufficient to melt it. Hydrogen gas is produced and lithium hydroxide is formed. Since lithium hydroxide is an alkaline solution, the indicator will change colour to show the presence of an alkaline solution.

$\text{Lithium + water} \longrightarrow \text{lithium hydroxide + hydrogen}$

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

Sodium reacts vigorously than lithium. Immediately it is dropped in water, it move on the surface of the water with a hissing sound. The water changes colour depending on the indicator solution used. When the sodium is placed on top of a floating piece of filter paper, it catches fire.

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

$\text{Sodium+ water} \longrightarrow \text{sodium hydroxide + hydrogen}$

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

Potassium reacts explosively with water. Immediately it is dropped in water it ignites spontaneously. The heat produced is so much that it ignites the hydrogen gas given out. Potassium hydroxide solution, which is an alkaline solution, is also formed.

Potassium+ water $\longrightarrow$ potassium hydroxide + hydrogen

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

For all the three metals, depending on the indicator used the colour change of the solution will be as follows:

indicator	**Colour in**	
	Acid solution	**Alkaline solution**
Litmus paper	Red	Blue
Bromothyl blue	Yellow	Red
Universal indicator	Red	Purple
Methyl orange	Red	Yellow
phenolphthalein	colourless	Pink

Note: Group I elements are stored in paraffin because they can react with oxygen in the atmosphere.

b. Reaction of alkali metals with chlorine

Experiment 1.4

Aim

To investigate the reaction of alkali metals with chlorine

Caution:

- Do not attempt this reaction with potassium
- Reaction with chlorine should be done in a fume cupboard or an open space.

Apparatus and chemicals

- Deflagrating spoon
- Bunsen burner
- Tile (ceramic)
- Two jars of chlorine
- Knife
- Pair of tongs
- Filter paper

- Sodium
- Lithium

Procedure

1. Using a pair of tongs, remove a small piece of lithium from the bottle and place it on the ceramic tile covered with a filter paper to remove much of the oil.
2. Using a knife, cut a small piece of lithium (size of a grain of rice) and return the rest to the bottle
3. Transfer the small piece of lithium into a deflagrating spoon
4. Hold the spoon directly into the Bunsen burner flame with your hand far away from the flame. *Protect your eyes too.*
5. As soon as the lithium catches fire, lower t quickly into the gas jar of chlorine.
    - What do you observe?
6. Repeat this procedure with sodium
7. Record your observations in a table like table 1.5. Which metal reacts with chlorine most readily? Explain.

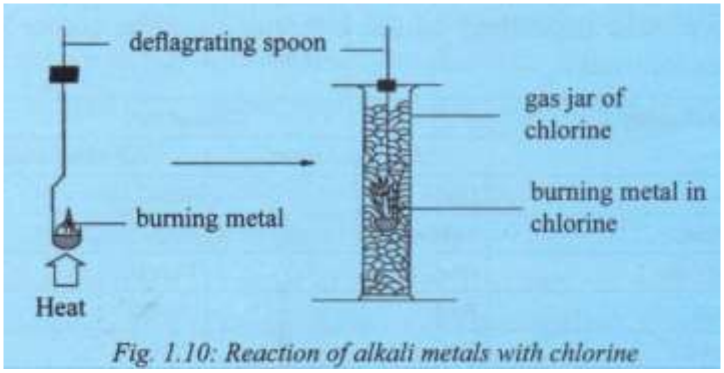


Fig. 1.10: Reaction of alkali metals with chlorine

Table 1.5 Reaction of alkali metals with Chlorine

Element	**Observation (include balanced chemical equations)**
Lithium	
Sodium	

Alkali metals react readily with chlorine to form compounds known as chlorides. Lithium forms a white solid of lithium chloride while sodium also forms a white solid of sodium chloride.

Lithium + chlorine ⟶ lithium chloride

$2Li\ (s) + Cl_2\ (g) \longrightarrow 2LiCl\ (s)$

Sodium + chlorine ⟶ sodium chloride

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

Note: the experiment should NOT be performed in the laboratory.

Potassium reacts explosively with chlorine to form a white solid of potassium chloride.

Potassium + chlorine $\longrightarrow$ potassium chloride

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

Self-assessment exercise 1.1

1. Define the following terms:
   a. Atomic radius
   b. Electron affinity
   c. Ionization energy
   d. Electronegativity
2. Explain why:
   a. Potassium is more reactive than sodium
   b. Magnesium is less reactive than sodium

Uses of alkali metals

a. Lithium

- Lubricating greases are produced from lithium
- Lithium is used in deoxidizing copper and copper alloys
- Lithium -6 is a main source for the production of titanium
- Lithium is used in batteries which contain more energy compared to other metals. These batteries are used in cell phones and computers.

b. Sodium

- Sodium metal is used in the preparation of tetraethyl lead, an important antiknock reagent in petrol. However, tetraethyl lead is being phased out in many countries because of lead pollution problems.
- Sodium metal is used in the extraction of titanium metal form titanium chloride
- Sodium vapour is used in lamps for street lighting.
- NaCl, a sodium compound, is widely used to season or add taste to food. It is not healthy to eat too much salt
- Sodium carbonate, is widely used in the manufacture of glass, detergents and for softening hard water
- Sodium hydroxide and sodium chloride are used in the manufacture of soaps and detergents.

c. Potassium

- Potassium chloride is used in making fertilizer, it is essential for the growth of plants
- Potassium chromate is used in tanning leather, in the manufacture of inks, gun powder and dyes

- Potassium hydroxide is used to make detergents
- Potassium nitrate is used as a food preservative.

1.4 Group II Elements (Alkaline-Earth Metals)

The group II elements are known collectively as alkaline-earth metals. They are found at the left hand side of the periodic table just after Group I elements as illustrated in Fig 1.11

Members include beryllium (Be), Magnesium (MG) and calcium (Ca)

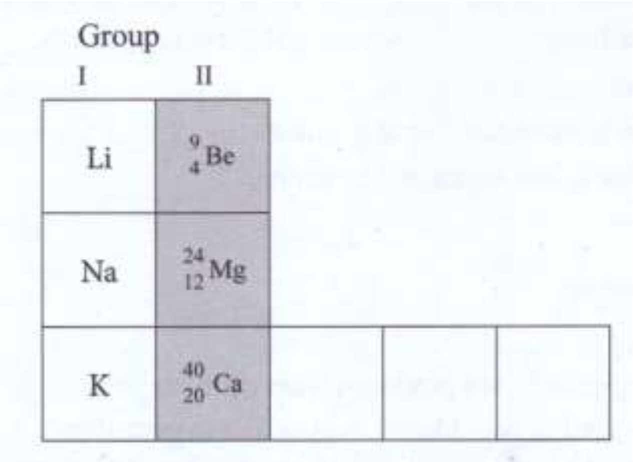


Fig. 1.11: Alkaline-earth metals

Trends in physical properties of alkaline-earth metals

Group II elements have two electrons in their outermost energy levels. They have similar physical and chemical properties. Beryllium is the only member of this group that has unique characteristics.

The properties of Group II elements are as summarized in Table 1.6 below.

Table 1.6: physical properties of alkaline-earth metals

Element	**Symbol**	**Electron arrangement**	**Melting point**	**Boiling point**	**1st ionization energy**	**Physical appearance**	**Thermal conductivity**	**Electrical conductivity**
Beryllium	Be	2.2	1278^{0}C	2970 ^{0}C	899	Silvery metal	Good conductor	Good conductor
Magnesium	Mg	2.8.2	669^{0}C	1107^{0}C	738	Silvery metal	Good conductor	Good conductor
Calcium	Ca	2.8.8.2	839^{0}C	1484^{0}C	590	Silvery metal	Good conductor	Good conductor

In general, their characteristics include:

- High melting and boiling points
- Grey silvery surface
- Good conductors of heat
- Good conductors of electricity

The following are important trends down Group II elements

- Melting points and boiling points generally decreases. This is because of increased atomic size which leads to a decrease in the forces of attraction between the positive nucleus and the electrons.
- Density increases due to increased mas of the atom
- The atomic radii increase down the group. This is because down the group, the number of energy levels increases. Another reason is that the electrons in the inner energy levels shield the electrons in the outer levels from full attraction by the nucleus
- Ionization energy decreases down the group because as one goes down the group the radii of the atoms increase. As a result, the attraction of electrons in the outermost energy level by the nucleus decreases hence the decrease in ionization energy.

Trends in chemical properties of alkaline-earth metals

Some of the chemical properties that will be studied include:

- Reaction with air or oxygen
- Reaction with water
- Reaction with chlorine
- Reaction with dilute acids

a. Reaction of alkaline-earth metals with air

Experiment 1.5

Aim

To investigate what happens when alkaline earth metals react with air

Apparatus and chemicals

- Bunsen burner
- Pair of scissors
- Pair of tongs
- Emery paper

- Magnesium ribbon
- Calcium granules

i. Reaction of magnesium with air

Procedure

1. Using a pair of scissors, cut about 4-5 cm of magnesium ribbon and observe it. What is the appearance of the magnesium ribbon?
2. Clean the surface of the magnesium with emery paper. What is the colour of magnesium ribbon after cleaning
3. Hold one end of the magnesium ribbon with the pair of tongs in a Bunsen flame as shown in Fig 1.12 until the magnesium catches fire.
4. Record your observation in your notebook.

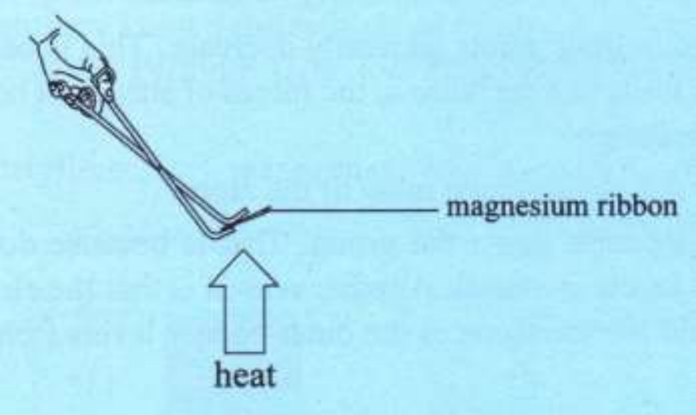


Fig. 1.12: Burning magnesium in air

Caution: looking directly at burning magnesium may affect your eyes.

ii. Reaction of calcium with air

Procedure

1. Hold a piece of calcium granule tightly with a pair of tongs. Observe the granule. What is the appearance of the calcium granule?
2. Hold the pair of tongs with the calcium granule directly in the Bunsen flame and heat strongly.
    - Record your observations in a table as in Table 1.7

Table 1.7: observation mad when alkaline-earth metals react with air

Alkaline-earth metals	Physical appearance of the metal	Observations of reaction with air	Description of the product	With an arrow show order of

			formed (write equations for the reaction)	decreasing vigour of reaction with air
Magnesium				
Calcium				

- Compare the reaction of magnesium and calcium with air
- Write equations for the reactions between the two metals and air

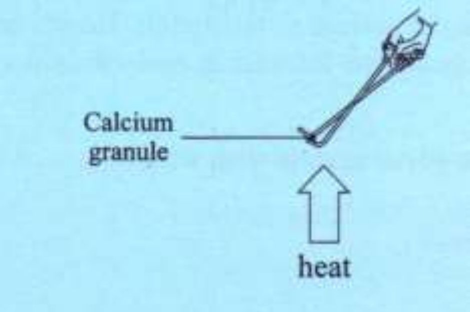


Fig. 1.13: Burning calcium in air

- What do you conclude about reactivities of magnesium and calcium with air?
- How would you expect the reaction of beryllium with air to compare with that of magnesium and calcium?

Generally, alkaline-earth metals burn in air to form simple metals oxides i.e.

Metal + oxygen ⟶ metal oxide

$2M\ (s) + O_2\ (g) \longrightarrow 2MO\ (s)$} M represents the metals

- Beryllium is a silvery metal. It has a strong but very thin layer of beryllium oxide on its surface which prevents any further attack from air. Beryllium is relatively stable compared to other metals in this group. It is known not to react with air even at over 600^0C. However, powdered beryllium metal does burn in air to give a white solid i.e. beryllium oxide.

  Beryllium + oxygen ⟶ beryllium oxide
  $2Be\ (s) + O_2\ (g) \longrightarrow 2BeO\ (s)$
- Magnesium burns readily with an intense bright white flame to produce a white powder of magnesium oxide.

  Magnesium+ oxygen ⟶ magnesium oxide
  $2Mg\ (s) + O_2\ (g) \longrightarrow 2MgO\ (s)$

- Calcium is also a silvery metal. The surface of calcium metal is covered with a thin layer of oxide that prevents the metal from further reacting with air. Without the oxide coating, calcium bursts into a white flame, which burns intensely to form a white solid of calcium oxide.

Calcium + oxygen calcium oxide →

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

Unlike alkali metals, beryllium, magnesium and calcium do not form peroxides when heated in air.

It might not be very clear to see the trend of reactivity of alkaline-earth metals with air due to an oxide coating on the surface of the metals. However, the reactivity increases as you go down the group due to increasing ease of removing the electrons in the outermost energy level.

b. Reaction of alkaline-earth metals with water

Experiment 1.6

Aim

To investigate what happens when alkaline-earth metals react with water

Apparatus

- Beaker
- Boiling tubes
- Pair of tongs
- Magnesium
- Calcium
- Water
- Burning splint

Procedure

1. Fill a boiling tube with water and invert it in a beaker half filled with water. Add a piece of calcium to the beaker. Adjust your boiling tube to collect the gas formed (see fig1.14)
2. When the boiling tubes is full, lift it out of the water, keeping it in an inverted position. Test the gas collected by bringing a burning splint to the mouth of the test tube
3. Observe the contents of the beaker. Test the solution in the beaker with red and blue litmus papers.
4. Add a small piece of cleaned magnesium ribbon to a beaker half-filled with water and observe. Heat the water to near boiling and observe any change in the reactivity of magnesium
5. Record your observations in a table as in table 1.8

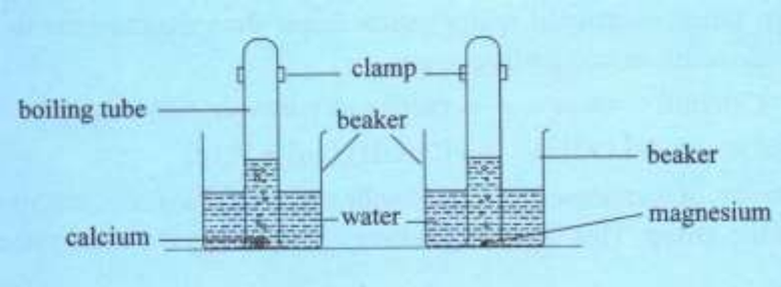


Fig. 1.14: Reaction of alkaline-earth metals with water

- Rank the metals tested with respect to their reactivity with water
- What gas was produced when calcium reacted with water? Give a reason for your answer. Where does this gas come from?
- When alkaline-earth metals react with water, are the resulting solutions acidic or basic? What class of metallic compounds is responsible for this property?

Table 1.8: observations made when alkaline-earth metals react with water

Alkaline-earth metals	**Colour change in litmus paper**	**Observations (write equations)**	**Reactivity (show increasing reactivity with an arrow)**
Magnesium			
Calcium			

- Although the metals tested varied in their rate of reactivity with water, they all produced hydrogen gas and an alkaline solution. Write word equations and chemical equations for the reactions of the metals with water.
- From your observations place the metals tested appropriately in the reactivity series.

Beryllium does not react with water or steam even when it is red-hot

Cleaned magnesium slightly reacts with cold water to form magnesium hydroxide solution which turns red litmus blue. Hydrogen gas is evolved in the process. The reaction stops after sometime because the magnesium hydroxide formed is almost insoluble in water and forms a barrier on the magnesium metal preventing further reaction.

Magnesium + oxygen → magnesium hydroxide + hydrogen

$Mg(s) + 2H_2O_2 (g) \longrightarrow 2Mg (OH)_2 (aq) + H_2 (g)$

The reaction with hot water is much faster and becomes even faster when magnesium reacts with steam.

- Calcium reacts with cold water much faster than magnesium to give calcium hydroxide solution and hydrogen gas.

Calcium + oxygen → calcium hydroxide + hydrogen

Ca(s) + $2H_2O_2$ (l) $\longrightarrow$ 2ca $(OH)_2$ (aq) + H_2 (g)

The reactivity of alkaline-earth metals with water shows increasing reactivity as you go down the group. This is mainly due to decreasing ionization energy down the group

c. Reactivity of alkaline-earth metals with chlorine

Experiment 1.7

Aim

To investigate what happens when alkaline-earth metals react with chlorine.

Caution: Chlorine is very poisonous gas that can cause irritation and trigger asthmatic attacks. This experiment should be done in a fume chamber or in an open space like a field.

Apparatus

- Bunsen burner
- Gas jars full of chlorine
- Pair of tongs
- Magnesium ribbon
- Sand paper
- Calcium turnings

Procedure

1. Clean the surface of magnesium ribbon to remove the oxide coating
2. Coil a small piece of metal (magnesium ribbon) on a deflagrating spoon and light using the Bunsen burner
3. Insert the burning magnesium into a gas jar of dry chlorine as illustrated in fig 1.15. What do you observe? What is the colour of the product?
4. For calcium, place the metal in a deflagrating spoon and burn
5. Insert the burning calcium into a gas jar of chlorine as illustrated in fig 1.16
6. Record you observations in a table as in Table 1.8

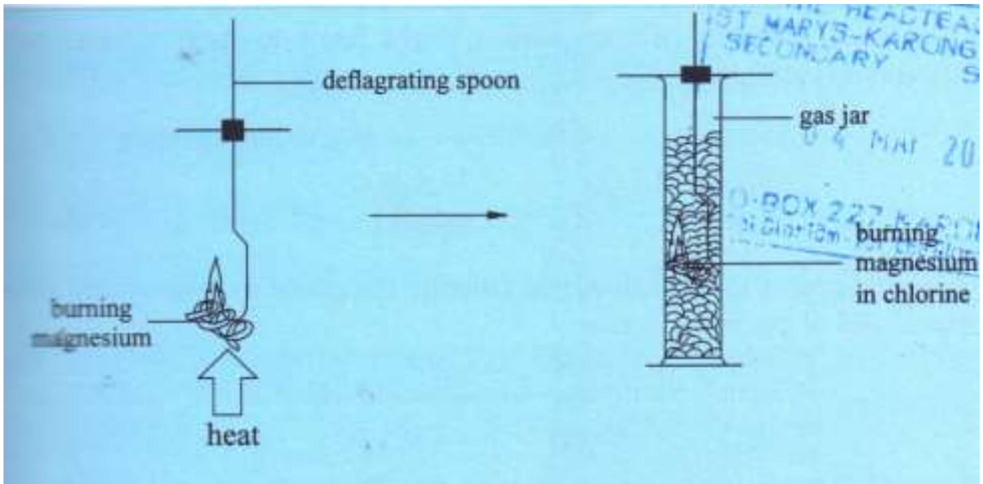


Fig. 1.15: Reaction of magnesium with chlorine

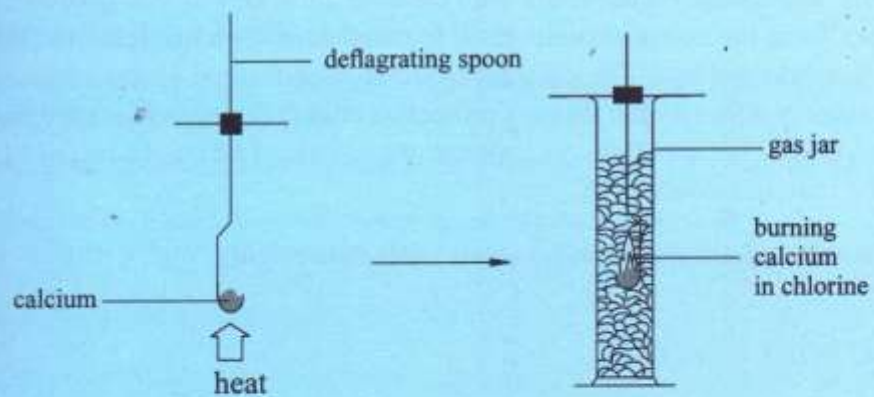


Fig. 1.16: Reaction of calcium with chlorine

- Compare the reactivities of magnesium and calcium with chlorine.

Table 1.9: reaction of alkaline-earth metals with chlorine

Element	**Observation**	**Equation of the reaction**
Magnesium		
Calcium		

Alkaline-earth metals react with chlorine to form chloride salts.

$$\text{Beryllium + chlorine} \longrightarrow \text{beryllium chloride}$$

$$Be\ (s) + CL_2\ (g) \longrightarrow BeCl_2\ (s)$$

- Magnesium burns in chlorine with a bright frame to form a white solid magnesium chloride.

$$\text{Magnesium + chlorine} \longrightarrow \text{magnesium chloride}$$

$$Mg\ (s) + CL_2\ (g) \longrightarrow MgCl_2\ (s)$$

- Calcium appears to burn slowly in chlorine compared to magnesium to form white solid of calcium chloride.

Calcium + chlorine $\longrightarrow$ calcium chloride

$Ca (s) + CL_2 (g) \longrightarrow CaCl_2 (s)$

Reactivity of alkaline-earth metals with chlorine increases as you go down the group. It appears from the above experiment that magnesium is more reactive than calcium. This is however not true. The calcium oxide coating formed on the surface as a result of its reactivity with oxygen forms a protective coat that makes the reaction of calcium to seem slower. The sinner silvery calcium metal cannot be easily reached by the chlorine hence the reaction slows down.

d. Reaction of alkaline-earth metals with dilute acids

Experiment 1.8

Aim

To investigate what happens when alkaline-earth metals react with dilute acids

Apparatus and chemicals

- Bunsen burner
- Test tubes
- Test tube rack
- Burning splint
- Magnesium ribbon
- Calcium
- Dilute hydrochloric acid
- Dilute sulphuric acid
- Dilute nitric acid

Procedure

1. In two separate test tubes, place a 2 cm strip of clean magnesium ribbon and a small piece of calcium
2. Add 5 cm^3 of dilute hydrochloric acid to each of the test tubes and observe what happens

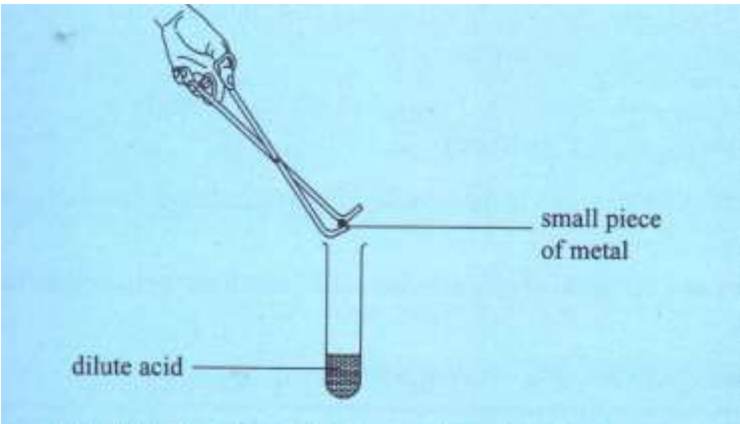


Fig. 1.17: Reaction of alkaline-earth metals with dilute acids

3. If a gas is produced, hold your thumb over the mouth of the test tube until you feel the pressure build up. Release you thumb and immediately bring a burning splint to the mouth of the test tube.
4. Repeat procedures 2 and 3 using dilute sulphuric acid then dilute nitric acid in place of the hydrochloric acid
5. Record your observations in a table as in Table 1.10

Table 1.10: observations made when alkaline-earth metals react with dilute acids.

Element	**Hydrochloric acid**	**Sulphuric acid**	**Nitric acid**
Magnesium			
Calcium			

- Did all the metals tested react with dilute hydrochloric acid? Which one was more reactive? Which was least reactive?
- In reactions that produced a gas, was the same gas formed in each case? Identify the gas
- Compare the reactivity of each metal in the different acids. Is there any difference in the rate of the reaction in the different acids? Do you notice any pattern?
- Compare the order of reactivity of the metals in each of the acid

When metals react with acids, two products are formed. One is a gas and the other is a salt. The general word equation and chemical equation for the reaction of a metal with an acid is:

Metal + Dilute acid ⟶ salt + h7ydrogen gas

Alkaline-earth metals react with different dilute acids to form corresponding slats and hydrogen gas only i.e.

$2M (s) + 2HCl (aq) \longrightarrow 2MCl (aq) + H_2 (g)$ (hydrochloric acid)

(Chloride)

M (s) + H_2SO_4Cl (aq) $\longrightarrow$ MSO_4 (aq) + H2 (g) (sulphuric acid)

(Sulphate)

2M (s) + $2HNO_3$ (aq) $\longrightarrow$ $2MNO_3$ (aq) + H_2 (g) (nitric acid)

(Nitrate)

The evolution of hydrogen gas is evident by the 'pop' sound by the burning splint.

Table 1.111 give a summary of the products of reactions between various acids and different metals.

Table 1.11: reaction of alkaline-earth metals with dilute acids

Alkaline-earth metal	**Dilute acid**	**Equation of the reaction**
Beryllium (Be) the surface of beryllium metal is covered with a thin layer of beryllium oxide that prevents the metal form further reacting with acids, but powdered beryllium metal dissolves readily in dilute acids to form a salt and hydrogen gas		
	Hydrochloric acid HCl (aq)	Beryllium + hydrochloric acid $\longrightarrow$ beryllium chloride + hydrogen Be (s) + 2HCl (aq) $\longrightarrow$ $BeCl_2$ (aq) + H_2 (g)
	Sulphuric acid H_2SO_4 (aq)	Beryllium + sulphuric acid $\longrightarrow$ beryllium sulphate + hydrogen Be (s) + H_2SO_4 (aq) $\longrightarrow$ $BeSO_4$ (aq) + H_2 (g)
	Nitric acid HNO_3 (aq)	Beryllium + nitric acid $\longrightarrow$ beryllium nitrate + hydrogen Be (s) + 2HNO(aq) $\longrightarrow$ $Be(NO_3)_2$ (aq) + H_2 (g)
Magnesium (Mg)	Hydrochloric acid HCl (aq)	Magnesium reacts with dilute hydrochloric acid to form magnesium chloride and hydrogen Mg (s) + 2HCl (aq) $\longrightarrow$ $MgCl_2$ (aq) + H_2 (g)
	Sulphuric acid H_2SO_4 (aq)	Magnesium reacts with dilute sulphuric acid to form magnesium sulphate and hydrogen Mg (s) +H2SO4 (aq) $\longrightarrow$ $MgSO_4$ (aq) + H_2 (g)
	Nitric acid HNO_3 (aq)	Magnesium reacts with dilute nitric acid to form magnesium nitrate and hydrogen Mg (s) + 2HNO3 (aq) $Mg(NO_3)_2$ (aq) + H_2 (g)
Calcium (Ca)	Hydrochloric acid HCl (aq)	Very dilute hydrochloric acids reacts with calcium vigorously to form calcium chloride and hydrogen gas. Calcium + hydrochloric acid $\longrightarrow$ calcium chloride + hydrogen gas Ca (s) + 2HCl (aq) $\longrightarrow$ $CaCl_2$ (s) + H_2 (g)

	Sulphuric acid H_2SO_4 (aq)	Dilute sulphuric acid reacts much slower with calcium compared with the other acids to form calcium sulphate and hydrogen gas. The reaction is slower because the calcium sulphate produced is slightly soluble and forms a coating on the surface of the calcium metal preventing any further reaction Calcium + sulphuric acid → calcium sulphate + hydrogen gas $Ca\ (s) + H_2SO_4\ (aq) \longrightarrow CaSO_4\ (s) + H_2\ (g)$
	Nitric acid HNO_3 (aq)	Very dilute nitric acid reacts, with calcium to form calcium nitrate and hydrogen gas. $Ca\ (s) + 2HNO_3\ (aq) \longrightarrow Ca(NO_3)_2\ (s) + H_2\ (g)$

Self- assessment exercise 1.2

1. Write down the word and chemical equation of the reaction between:
   a. Calcium and oxygen
   b. Lithium and chlorine
   c. Magnesium and sulphuric acid
2. Arrange the following elements in increasing order of reactivity: sodium, lithium, magnesium and calcium.

Uses of alkaline-earth metals

a. Beryllium

- Used in transmission of X-rays (beryllium transmits X-rays better that aluminum)
- An alloy of beryllium or copper is hard, strong and with high resistance to wear. It is therefore used to make computer parts, and other instruments with desirable lightness and stiffness.
- Alloys of beryllium are used as structural materials for high performance aircrafts, missiles, spacecraft and communication satellites among other things
- Beryllium oxide is used in the nuclear industry

b. Magnesium

- It is lighter than aluminum, hence it is used to make alloys used for manufacturing aircraft, parts of car engine casings and for missile construction.
- Used as a reducing agent in the production of uranium and other metals from their salts.
- Magnesium hydroxide (milk of magnesia), magnesium chloride and magnesium sulphate (Epsom salts) are used in the pharmaceutical industry in manufacturing medicine.

- Magnesium oxide is used as a brick-liner in furnaces.
- Used in computers of radio-frequency shielding

c. Calcium

- Used as a reducing agent in the preparation of metals such as thorium, uranium and zirconium
- Calcium forms calcium carbonate which is a component of cement. Cement is used in building and construction.
- Calcium carbonate is used in antacid tablets

1.5 Group VII Elements (Halogens)

The elements in group VII are collectively known as halogens. Their group name comes from Greek words ‘hal, meaning ‘salt’ and ‘gen’ meaning to ‘produce. Therefore, halogens are known for their salt producing property. They are found at the right hand side of the Periodic table just before noble gases as shown in fig. 1.18.

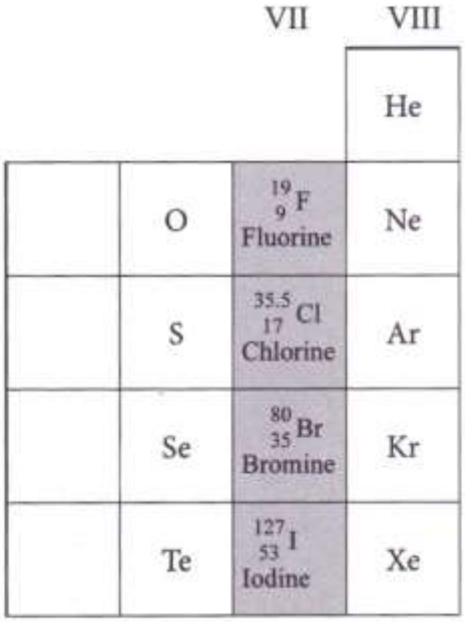


Fig. 1.18: Halogens

Trends in physical properties of halogens

All of the halogens are highly reactive and occur only as compounds in nature. The halogens show great similarities to each other in their properties as shown in Table 1.12

Table 1.12: physical properties of halogens

Element	**Symbol**	**Electron arrangement**	**Boiling point**	**Melting point**	**Physical appearance**	**Physical state at room temperature**
Fluorine	F	2.7	-223^{0}C	-187^{0}C	Pale yellow	Gas
Chlorine	Cl	2.8.7	-102^{0}C	-35^{0}C	Pale green-yellow	Gas
Bromine	Br	2.8.18.7	7^{0}C	59^{0}C	Deep red	Liquid
Iodine	I	2.38.18.18.7	114^{0}C	183^{0}C	Dark purple	Solid

Fluorine (F) and Chlorine (Cl), are diatomic gases (F_2, Cl_2), at room temperature. Fluorine has a pale -yellow colour, while chlorine is a pale green-yellow gas. Both have a chocking smell that is harmful to the lungs and nasal passages. Bromine is a dense deep-red liquid that exists as bromine molecule (Br_2) and iodine is a shiny, dark purple solid, existing as a diatomic molecule (I_2).

Note: bromine gives off red-brown fumes and iodine gives off purple fumes both of which are harmful to the respiratory system.

- The melting and boiling points of group seven elements are generally low but increase down the group. Fluorine has the lowest melting and boiling points whereas iodine has the highest. The intermolecular forces of attraction become stronger as one goes down the group due to increased nuclear charge as the number of protons increases down the group. This creates force of attraction on a stronger the electrons by the positive nucleus. It is harder to break this force of attraction down the group hence intermolecular forces of attraction are stronger in iodine. It needs more energy to break these forces to separate the molecules.
- Group VII elements do not conduct both electricity and heat. This is because they exist as molecules with stronger intermolecular forces of attraction bonding the atoms together.
- The atomic radii of halogens increases down the group. This is because down the group, the number of energy levels increases. However, the atomic sizes of halogens atoms are smaller compared to the atomic sizes of alkali and alkaline-earth metals in the same period. This is because across the period, these is increase in the effect of nuclear charge. This then causes on increase in force of attraction of the outermost electrons thereby reducing the size of the atom.
- Ionic sizes of halogens increase down the group but they are larger compared to their atomic sizes. Let us consider how their ions are formed. See Fig 1.19.

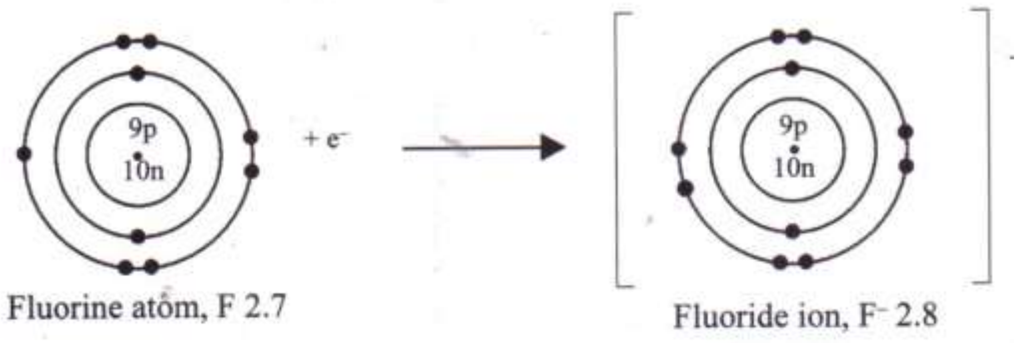


Fig. 1.19: Ion formation in fluorine

Fluoride ions have more electrons than protons. For example, in the above case it has gained one additional electron. This leads to increase in the overall size of the ion. The sudden increase in the ionic size of a fluoride ions is due to the increase in repulsion forces among the electrons in the complete outermost energy level because the nuclear charge remains the same. Furthermore, there is a larger number of electrons to be attracted by the same number of protons in the nucleus.

Trends in chemical properties of halogens

Caution:

- Bromine liquid is very corrosive and can cause burns. It has a very irritating vapour. In case of skin contact, add sodium thiosulphate and large amounts of water and seek medical attention quickly. For spillages, ad sodium carbonate and a lot of water.
- Chlorine is extremely irritating and sometimes may trigger off asthma attack

Chlorine should always be handled within a fume chamber or in an open space.

- Iodine solid is corrosive and may cause burns. In case of skin contact, wash with large amounts of water

a. Reaction of halogens with water

Experiment 1.9

Aim

To investigate the reaction of halogens with water:

Apparatus and chemicals

- A gas jar of Chlorine
- A small amount of bromine liquid
- A few small crystals of iodine solid

- Beakers
- Universal indicator solution or blue litmus paper
- Water
- Stirring rod
- Spatula
- Boiling tube

Procedure

i. Reaction of chlorine with water

Caution: this should be done in a fume of chamber or open space

1. To gas a jar full of chlorine, add a small amount of water and universal indicator solution. Litmus paper can also be used.
2. Replace the cover slip and shake as illustrated in fig. 1.20
3. Observe any changes
4. Record your observations as illustrated in Table 1.12
    - What is the colour of water with the indicator?
    - What do you observe when the water is added to the gas and shaken? Explain your observations.

ii. Reaction of bromine with water

1. Carefully place 1-2cm^3 of bromine liquid in a large beaker containing about 250cm^3 of water
2. Stir with glass rod as illustrated in Fig 1.21
3. Observe any changes
4. Put blue litmus paper in the solution formed
5. Record your observation as illustrated in Table 1.12
    - What is the initial colour of bromine liquid?
    - What happens when bromine liquid is dissolved in water?

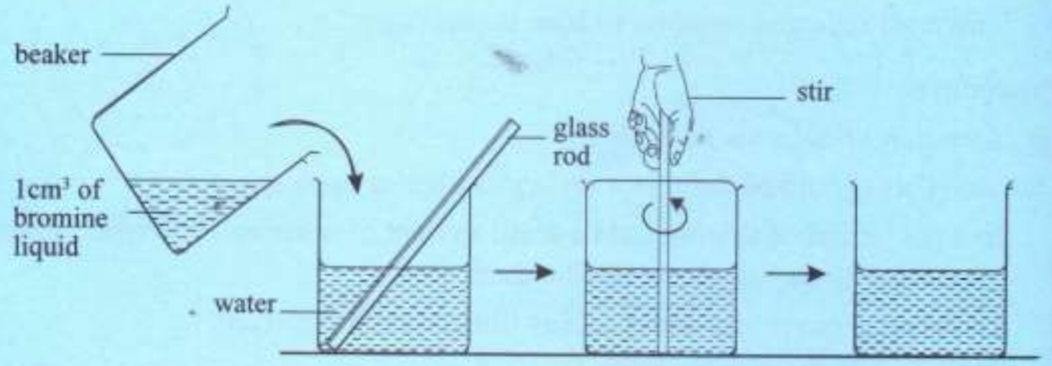


Fig. 1.21: Reaction of bromine with water

(iii) Reaction of iodine with water

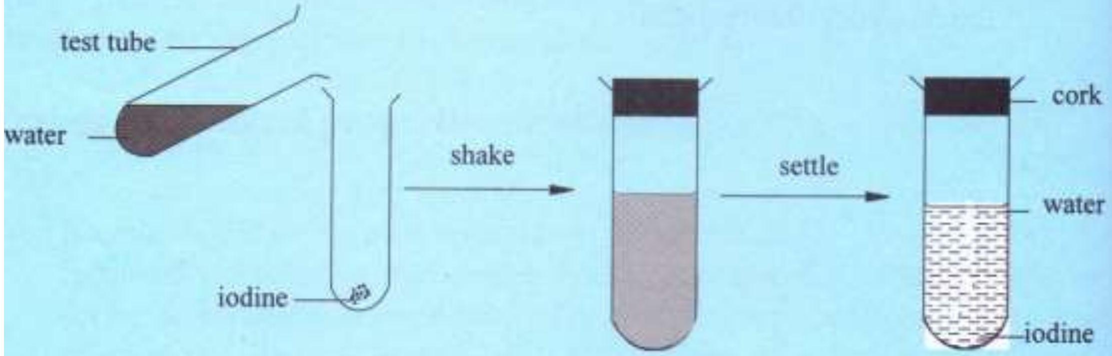


Fig. 1.22: Reaction of iodine with water

1. Put a spatula full of iodine solid into a boiling tube
2. Add about 10cm3 of water
3. Put a stopper on and shake
4. Allow to settle
5. Record your result in a table like table 1.13 below

Table 1.13 reaction of halogens in water

Element	**Physical appearance and state of element before reaction**	**Observation (reaction with water)**	**Names of products formed**
Chlorine			
Bromine			
Iodine			

Fluorine reacts with water to form hydrogen fluoride and oxygen gas.

Fluorine + water → hydrogen fluoride + oxygen

$2F_2\ (g) + 2H_2O\ (l) \longrightarrow 4HF\ (aq) + O_2\ (g)$

When chlorine dissolves in water, the colour of universal indicator rapidly changes. Chlorine in water dissolves to form a pale yellow solution which has the smell of bleach. Blue litmus paper turns red and is then bleached. Chlorine water is acidic and has bleaching properties. The reaction can be summarized in the following equation.

Chlorine + water $\longrightarrow$ hydrogen chloride + oxygen

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

The presence of chloric (I) acid 9hypochlorous acid) gives the solution bleaching properties.

Bromine liquid is slightly soluble in water in which it forms a light brown solution which is acidic. Bromine water has less bleaching characteristics compared to chlorine water.

Iodine is almost insoluble in water.

b. Reaction of halogens with metals

Experiment 1.10

Aim

To investigate the reaction of halogens with sodium, zinc and iron metals

Apparatus and chemicals

- Thistle funnel
- Round-bottom flask
- Concentrated hydrochloric acid
- Potassium permanganate
- Side arm boiling tubes
- Boiling tubes
- Delivery tubes
- Clamp stand
- Bunsen burner
- Corks
- Bromine liquid
- Iodine
- Sodium
- Zinc

- Iron wool (steel wire)

Procedure

i. *Reaction of chlorine with sodium, zinc and iron metals*

1. Prepare chlorine gas as illustrated in the set up in fig 1.23
2. Heat sodium in the deflagrating spoon and insert the burning sodium into a jar of chlorine
3. In the case of zinc or iron, place the metal in the test-tube as in fig. 1.23
4. Pass a steady stream of chlorine gas over the heated metal
5. Heat the metal until the reaction stops and then remove the Bunsen burner
6. Write down your observations and equations for the reactions.

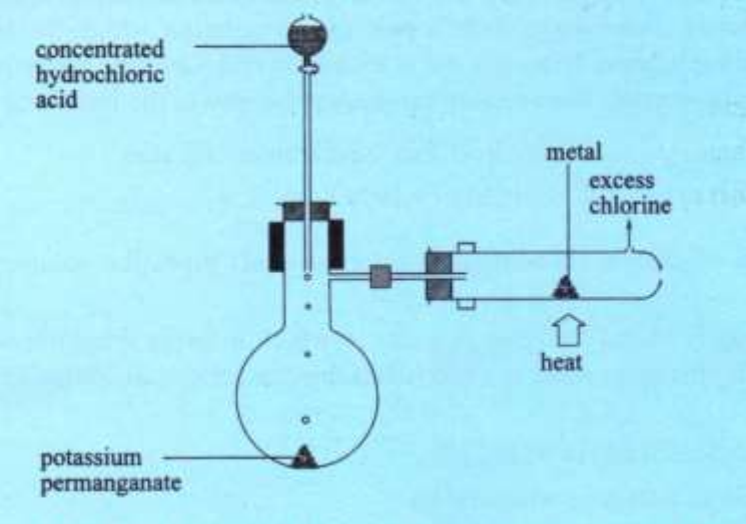


Fig. 1.23: Set-up for the reaction of chlorine gas with zinc or iron

ii. *Reaction of bromine with metals (sodium, zinc and iron)*

1. place bromine liquid in a corked conical flask with a side arm
2. Place the metal in the tube as shown.
3. Pass a steady stream of bromine gas over the heated metals as illustrated in fig 1.24
4. Remember to heat the metal until the reaction stops.

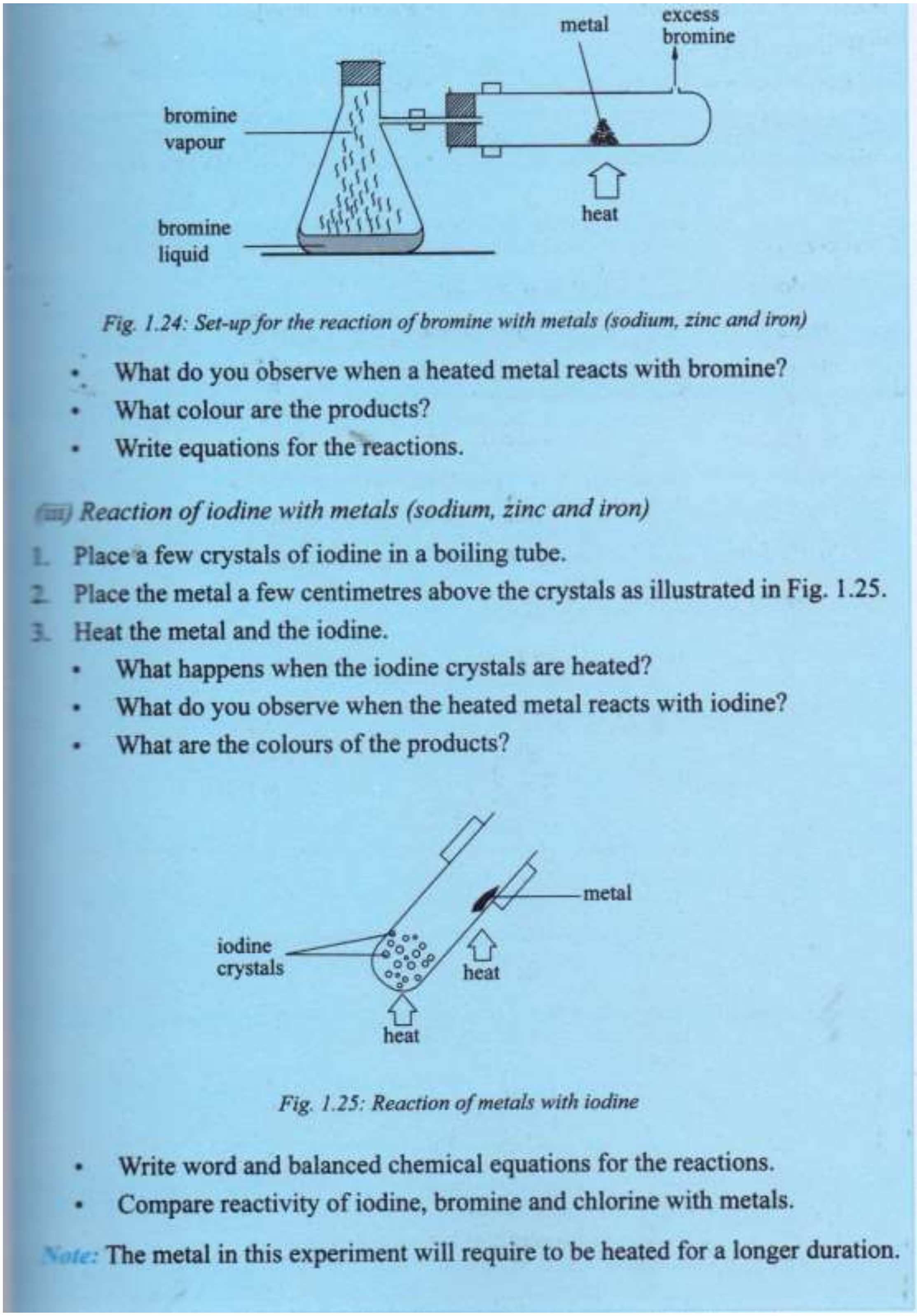


Fig. 1.24: Set-up for the reaction of bromine with metals (sodium, zinc and iron)

- What do you observe when a heated metal reacts with bromine?
- What colour are the products?
- Write equations for the reactions.

(iii) Reaction of iodine with metals (sodium, zinc and iron)

1. Place a few crystals of iodine in a boiling tube.
2. Place the metal a few centimetres above the crystals as illustrated in Fig. 1.25.
3. Heat the metal and the iodine.
   - What happens when the iodine crystals are heated?
   - What do you observe when the heated metal reacts with iodine?
   - What are the colours of the products?

Fig. 1.25: Reaction of metals with iodine

- Write word and balanced chemical equations for the reactions.
- Compare reactivity of iodine, bromine and chlorine with metals.

Note: The metal in this experiment will require to be heated for a longer duration.

Table 1.14 summarizes the observations made when various metals react with different halogens

Table 1.14: reaction of halogens with metals

Halogen	**Metal**	**Reaction**	**Reactivity**
Fluorine	Sodium, zinc and iron	Fluorine is generally a highly toxic and very reactive pale-yellow gas. Due to this reactivity, fluorine element or molecule is never found in nature. In fact no other chemical element can displace fluorine Reactions of fluorine with sodium, zinc and iron cannot be done in an ordinary laboratory	Reaction of metals with fluorine is very explosive and dangerous hence should not be done in an ordinary laboratory
Chlorine	Sodium	Heated sodium metal reacts vigorously with chlorine to form a white salt of sodium chloride. Sodium + chlorine → sodium chloride $2 Na (s) + Cl_2 (g) \longrightarrow 2NaCl (s)$ (white) The sodium chloride salt produced is soluble in water and gives a neutral solution of pH= 7 *Note:* with sodium, once the reaction is on, the source of heat can be removed and the reaction will still proceed	Reaction of metals with chlorine is vigorous
	Zinc	Zinc reacts when heated in chlorine to form a white salt of Zinc chloride. Zinc + chlorine → Zinc chloride $Zn (s) + Cl2 (g) \longrightarrow ZnCl_2 (s)$ (white) Zinc chloride salt is also soluble in water. It forms a neutral solution of pH= 9 *Note:* continued heating is not necessary for this reaction to continue.	
	Iron	When iron is heated, it glows red. Chlorine gas reacts with the red hot iron wool to form a brown solid of iron (III) chloride. Iron + chlorine → iron (III) chloride $Fe (s) + Cl2 (g) \longrightarrow FeCl_3 (s)$ (brown) *Note:* continued heating is not required for this reaction to continue	
Bromine	Sodium	Sodium when heated reacts with bromine vapour (gas) to form a white solid, sodium bromide. Sodium + bromine → sodium bromide $2Na (s) + Br_2(g) \longrightarrow 2NaBr (s)$ (white)	The reaction of bromine with metals is slower compared to the one with chlorine

		Note: heating must be continued for the reaction to proceed	
	Zinc	Heated zinc reacts with bromine vapour to form white salt, zinc bromide Sodium + bromine $\longrightarrow$ sodium bromide $2Zn\ (s) + Br_2(g) \longrightarrow 2ZnBr_2\ (s)$ (white)	
	Iron	Iron must be heated gently and continuously to react with bromine to form a brown salt, iron (III) bromide. Iron + bromine $\longrightarrow$ iron (III) bromide $2Fe\ (s) + Br_2(g) \longrightarrow 2FeBr_3\ (s)$ (brown)	
iodine	Sodium	When iodine is heated, it sublimes to form purple iodine vapour which reacts with heated sodium to form a slat, sodium iodide. Sodium + iodine $\longrightarrow$ sodium iodide $2Na\ (s) + I_2(g) \longrightarrow 2NaI\ (s)$ (white)	Iodine reacts with metals very slowly even with strong and continuous heating
	Zinc	Heated zinc reacts slowly even with continuous heating with iodine vapour to form slat, zinc iodide Znc + iodine $\longrightarrow$ sodium iodide $2Zn\ (s) + I_2(g) \longrightarrow 2ZnI_2\ (s)$ (white)	
	iron	Iron reacts very slowly, even with continuous heating, to form iron (II) iodide instead of iron (III) iodide as in the case of chlorine and bromine Iron + iodine $\longrightarrow$ iron (II) iodide $Fe\ (s) + I_2(g) \longrightarrow FeI_2\ (s)$ (white)	

Self-assessment exercise 1.3

1. Why is fluorine more reactive than chlorine?

Why is sulphur less reactive than chlorine?

2. Arrange the following elements in order of decreasing reactivity: chlorine, fluorine, bromine.

Uses of halogens

a. Fluorine

- It is used in the manufacture of plastics
- Hydrofluoric acid is extensively used in air conditioning and refrigerators
- In small amounts, fluoride in water sources prevents tooth decay. It is a constituent of toothpaste. Soluble fluoride in drinking water may damage the teeth of children growing permanent teeth

b. Chlorine

- It is used in water treatment all over the world. Even the smallest drinking water supply is currently being chlorinated
- It is extensively used in the production of paper products, dyestuffs, textiles, petroleum products, medicines, antiseptics, insecticides, foodstuffs, solvents, paints, plastics and many other products
- Most chlorine is used in the manufacture of chlorinated cleaning compounds, pulp bleaching, disinfectants and textile processing
- Chlorine is used in the manufacture of PVC pipes

c. Bromine

- Much of bromine is used in making petrol engine anti-knock compounds.
- It is used in the manufacture of fumigants, flame proof agents, water purification compounds, dyes, medicines and pesticides.
- Inorganic bromides like silver bromide are used in photography.

d. Iodine

- Iodine is used in medicine. A solution of potassium iodide and iodine alcohol is used as a disinfectant for external wounds
- Silver iodide is used in photography
- Iodine is added to table salt to prevent goiter
- It is used in testing for starch in biological sciences. A deep blue colour observed when starch reacts with iodine.
-

1.6 Group VIII elements- Noble gases

Group VIII elements are also known as *noble gases.* They were initially known as ‘inert’ gases until it was discovered that some of its members, like krypton and xenon can form compounds. Fig. 1.26 shows the common members of this group.

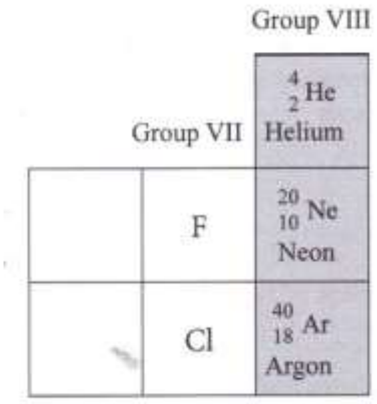


Fig. 1.26: The noble gases

Noble gases have fully filled outer energy levels as illustrated by the atomic numbers in fig. 1.26

Because of this they are very unreactive. They do not gain or lose electrons. The atomic size increases down the group because of the increase number of energy levels. Noble gases exist as monoatomic gases at room temperature. They are non-conductors of heat and electricity because they do not have delocalized electrons. The physical properties of Group VIII elements are summarized in Table 1.15 below.

Table 1.15 physical properties of noble gases

Element	**Symbol**	**Electron arrangement**	**Melting point ^0C**	**Boiling point ^0C**	**Atomic radius/ (nm)**	**Physical appearance**	**Physical state**	**Thermal conductivity**	**Electrical conductivity**
Helium	He	2	-272	-269	0.049	Colourless	Gas	Do not conduct heat	Do not conduct electricity
Neon	Ne	2.8	-249	-246	0.051	Colourless	Gas	Do not conduct heat	Do not conduct electricity
Argon	Ar	2.8.8	-189	-186	0.088	colourless	gas	Do not conduct heat	Do not conduct electricity

Helium, He 2

Neon, Ne 2.8

Argon, Ar 2.8.8

Fig. 1.27: Atomic structure and electron arrangement of noble gases

Uses of noble gases

a. Helium

- It is used in filling meteorological balloons as it is very light and safer than hydrogen. Hydrogen burns explosively in case of an accident.
- Used as a cooling medium for nuclear reactors
- A mixture of 80% helium an d20% oxygen is used as an artificial atmosphere for divers and others working under pressure
- Used as a protective gas for semiconductor materials
- Used in pressurizing rocket fuels to liquids

b. Neon

- Used in making neon advertising coloured signs. This is its largest use
- Used to make high voltage indicators
- Neon and helium are used in making gas lasers
- Liquid neon is an economical refrigerant

c. Argon

- Used in electric light bulbs and in fluorescent tubes
- Used as an inert gas shield for arc welding and cutting

Self-assessment exercise 1.3

1. State two uses of:
    a. Noble gases
    b. Alkali metals
    c. Alkaline-earth metals
2. Why is argon used in arc welding?

Hydrogen is the lightest gas but helium is preferred in filling the balloons. Give the reason why.

Summary

Alkali metals are in Group I. they include lithium (Li), sodium (Na) and potassium (K).

Alkali metals are soft and can be cut with a knife except lithium. They have low melting and boiling points

Alkali metals are very reactive compared to other metals. This is because they easily lose their outermost energy level electron.

Alkali metals form ions with one positive charge (+1). This is because their atoms lose one electron from their outermost energy levels.

The reactivity of alkali metals increase down the group. This is because the atomic radii increase down the group because of the increase in the number of energy levels.

Alkali metals form compounds with similar formulae, for example lithium sulphate (Li_2SO_4), potassium sulphate (K_2SO_4).

Alkali metals are very reactive. They tarnish easily when exposed to air due to formation of an oxide coating. That is why they are stored in paraffin or oil.

Alkaline-earth metals include beryllium, magnesium and calcium

Alkaline-earth metals have high melting and boiling points compared to alkali metals

The reactivity of elements of Group II of the Periodic table increases down the group just as in alkali metals

Alkaline-earth metals form ions with two positive charges (+2). This is because they lose two electrons from their outermost energy levels.

Alkaline-earth metals form compounds with similar formulae. For example beryllium hydroxide, $Be\ (OH)_2$, magnesium hydroxide $Mg\ (OH)_2$, and calcium hydroxide $Ca\ (OH)_2$.

Halogens are group VII elements include fluorine, chlorine bromine and iodine.

They are a group of reactive non-metals with low melting points and boiling points.

Fluorine and chlorine are gases at room temperature. Iodine sublimes on heating.

Noble gases are in Group VIII include helium (He), neon (Ne) and argon (Ar). They are the least reactive elements in the periodic table

Their atoms have the most stable arrangement of electrons in the outermost energy level. Therefore they do not gain or lose electrons.

Revision exercise 1

1. What is the difference between a metal and a metalloid? Give one example of metalloid.
2. Define the following terms:
   a. Atomic radius
   b. Ionization energy
3. Reactivity for alkali metals increases down the group while that of halogens increasing up the group. Explain this in terms of electrons loss and gain.
4. Sodium is more reactive than magnesium. Explain.
5. Name one use of the following:
   a. Sodium
   b. Neon
   c. Magnesium
6. The diagram below represents part of the Periodic table beginning with element of atomic number 3 to element L, atomic number 20.

The letters do not represent the actual symbols of the elements which occur in the positions shown, but they may be considered so for purpose of answering the questions below.

I II III IV V VI VII VIII

3					R	E		J	D
A				Q			G	M	
	^{20}L								

   a. How many electrons are there in the outermost energy level of the atom of D?
   b. Write the electron configuration of an ion of element L.
   c. What is the total number of elements in an atom of Q, M and E?
   d. Which two elements show a valence of three?
   e. Which one of the two elements in question is likely to form salts containing 3+ ions?
   f. How many neutrons are there in each atom of the principal isotopes of R?
   g. Give the letters of all elements which are gases at room temperature
   h. What would be the formulae of the compounds formed between J and A, j and L, L and G?
   i. What is the formulae of an oxide formed between element Q and oxygen?

Unit 2: Chemical bonding

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. This is because their outermost levels are completely full.

The tendency among other elements is to strive to attain the stable noble gas electronic arrangement. Some will do this by "donating" electrons; others by "accepting" electron while still others will do so sharing electrons. In this unit, we shall discuss what happens when atoms of elements donate, accept or share electrons, so as to achieve the stability of noble gases. When this happens, the atoms joining together, a chemical bond is formed. A chemical bond is an attraction force between two particles. The particles could be atoms or ions. We are going to study three main types of bonds namely:

- Ionic bonding or electrovalent bond
- Covalent bond
- Metallic bond

2.1 Ionic chemical bonding

We cannot see ionic bonds. However, ionic bonding can be best imagined as one big greedy dog stealing a bone form a little dog Fig. 2.1

Suppose the bone that is up for grabs represents an electron, and the dogs represent atoms of two different elements. When the big dog snatches the bone from the smaller dog, it I similar to one atom gaining an electron from other atom/. The bigger dog will be a bone richer, while the smaller dog will be a bone poorer.

In atoms, the atom that loses an electron becomes a positively charged ion and the one that gains the electron becomes a negatively charged ion.

(a) Two dogs with different strengths holding a bone. The huge dog is trying to snatch the bone from the smaller dog angrily.

(b) The dogs are in a big fight, there is a lot of dust.

(c) The huge dog walks away with the bone while the smaller dog is left very sad.

Fig. 2.1: Illustration of electrovalent bonding

Consider a specific case of ionic bonding between real elements such as sodium and chlorine. A sodium atom contains 11 protons and has an electron arrangement of 2.8.1. The arrangement differs from the nearest noble gas electron structure, that of neon 2.8 by the presence of one extra electron in the third energy level.

On the other hand, a free chlorine contains 17 protons and has the electron 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 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 in its outermost energy level to gain the noble gas structure. Thus the single electron from the outermost energy level of sodium atom is transferred to the outermost energy level of chlorine atom as shown in fig 2.2

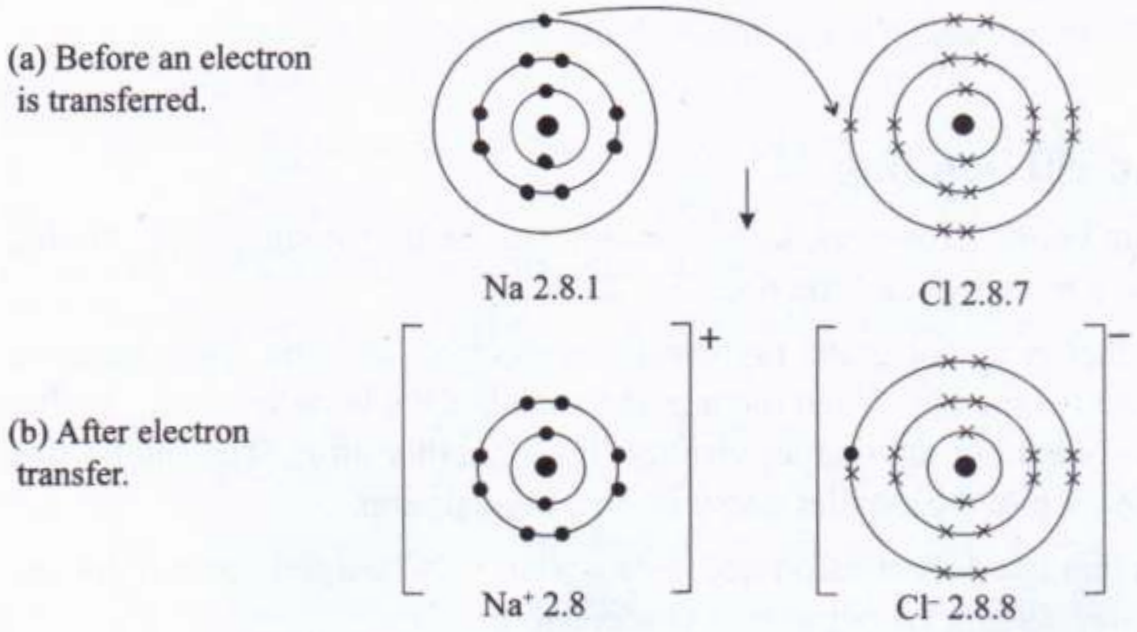


Fig. 2.2: Ionic bonding in sodium chloride (NaCl)

Sodium atoms has 11 positive charges (protons) balanced by 11 negative charges (electrons). Sodium ion has only 10 electrons. Therefore, since the positive charges in the nucleus are uncharged, there is one excess positive charge on the sodium ion. Similarly, chloride ion has one negative charge in excess of the positive charges.

A sodium ion is positively charged because of the one excess positive charge. Similarly a chloride ion is negatively charged because of the one excess negative charge. These ions have opposite charges. Remember that unlike charges attract each other.

The attraction of the oppositely charged Na+ and Cl- ions from a bond called ionic or electrovalent bond. This type of combination is called ionic bonding.

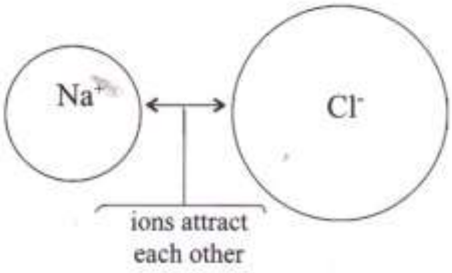


Fig. 2.3: Sodium ion combines with a chloride ion

Self-assessment exercise 2.1

The atomic number of an element X is 19 and that of another element Y I ds17.

a. What type of bond can X and Y form?
b. Draw a diagram to show bonding in X and Y.

Structure of sodium chloride

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

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.

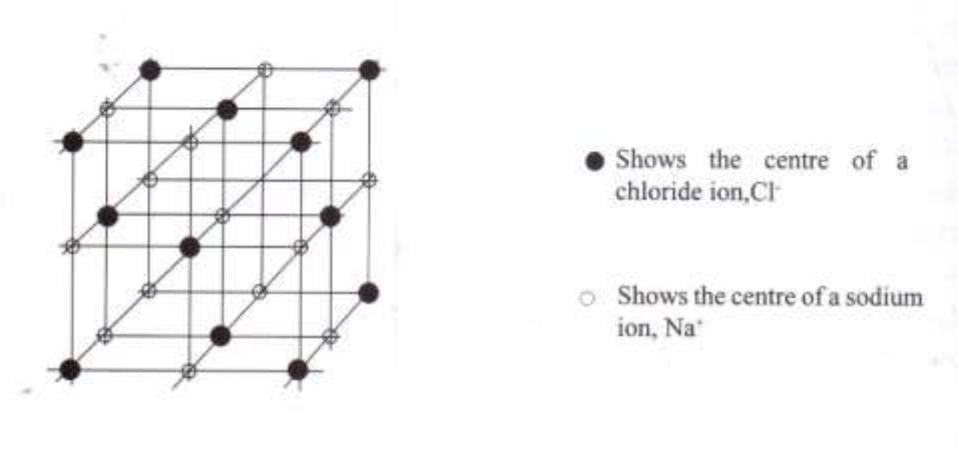


Fig. 2.4: Arrangement of sodium and chloride ions in sodium chloride crystal

Sodium and chloride ions crystallize a in a pattern (crystal lattice) forming a cube. In an end face by the cube, a Na+ ion occupies the centre with six CL+ ions spaced equally between them. 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 when sodium chloride is melted, sodium ions and chloride ions separate and thus the 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 more electrons are called cations and their positive charges are equal to the number of electrons lost. Likewise negative ions formed as a result of gaining one or more electrons are called anions and their negative charges are equal to the number of electrons gained.

The number of electrons lost form, 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 electrons involved must balance the valency requirements of elements as shown in the following examples.

a. Magnesium chloride

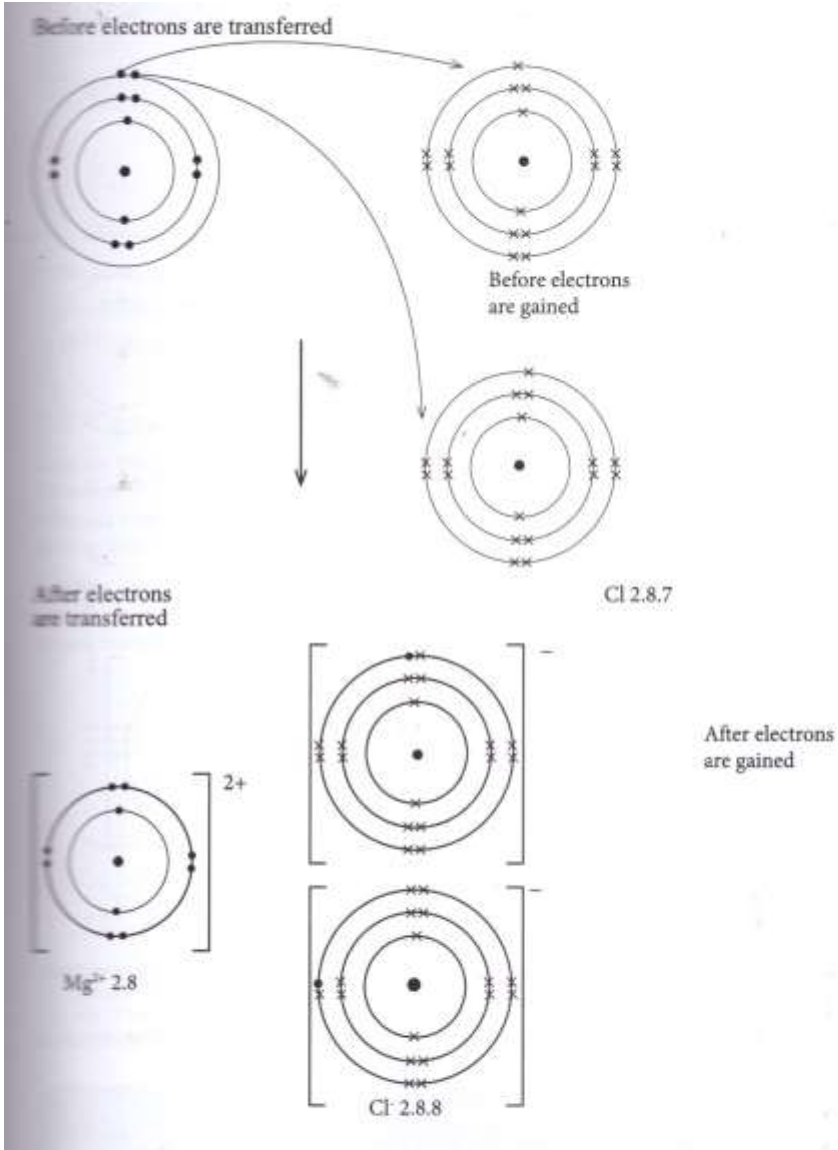


Fig. 2.5: Ionic bonding in magnesium chloride

The valency electrons (outermost energy level electrons) from the magnesium atoms are transferred to the outermost energy level of chlorine atoms as shown in Fig 2.5.

Therefore the formula of magnesium chloride is Mg^{2+}Cl$^-$ or MgCl$_2$

b. Calcium oxide

Before electrons are transferred

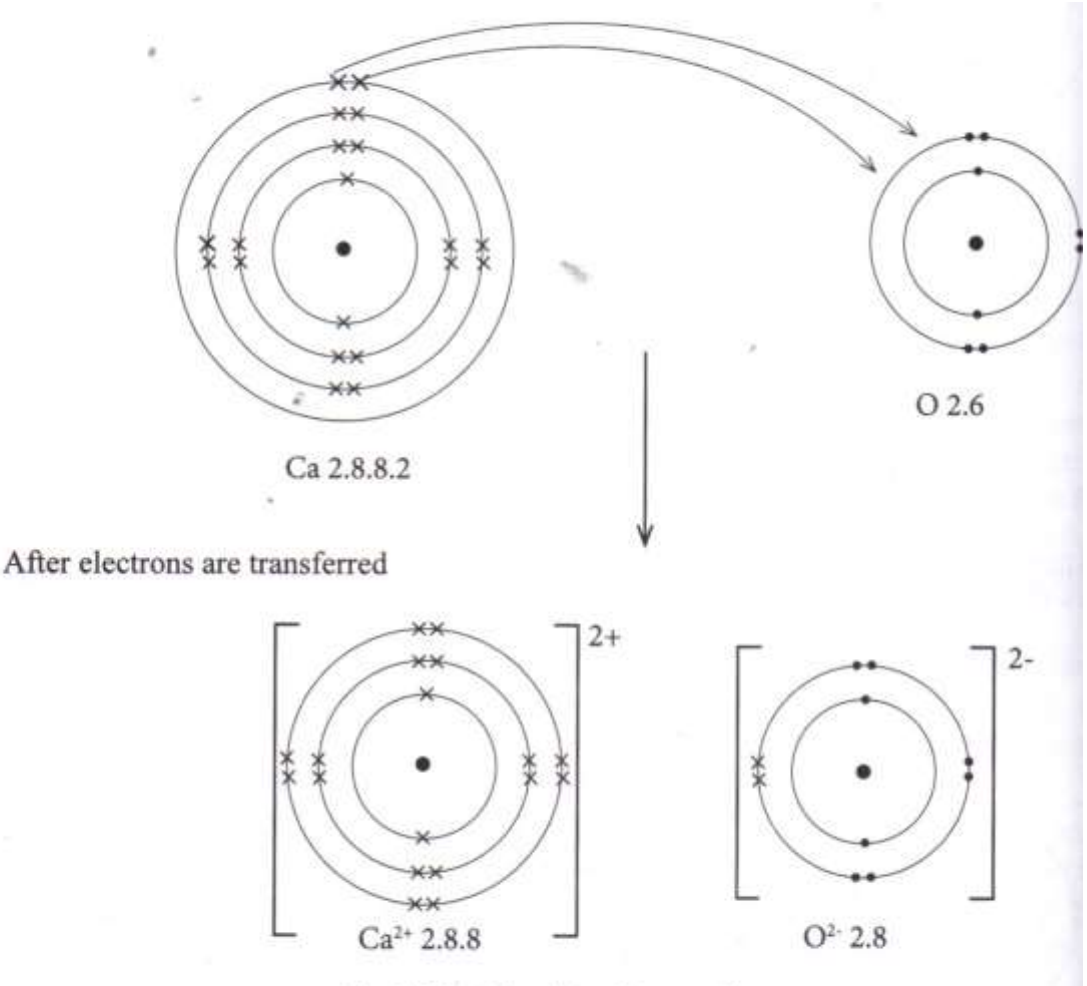


Fig. 2.6: Ionic bond in calcium oxide

The two valency electrons form one calcium atom are transferred to one oxygen atom.

Therefore, the formula of calcium oxide is Ca^{2+} O^{2-} or CaO

You may be required to write the formulae of ionic compounds from given symbols and ionic charges.

Example 1

Write the formula of potassium chloride give the charges of ions; K+ and Cl-.

$K^{+} + Cl^{-} \longrightarrow KCl$

Example 2

Write the formula of sodium oxide given the following charges of the ions; Na^+ and $O2^-$

The formula is Na_2O

Self- assessment exercise 2.2

Write the formula of the following compounds from the ions given:

a. Aluminium oxide: Al^{3+} and O^{2-}
b. Calcium bromide: Ca^{2+} and Br^-

2.2 Covalent chemical bonding

Imagine two dogs of equal strength fighting over two bones. The two dogs have equal attraction to the bones. Since they have equal strength, the dogs end up sharing the pair of bones evenly. If one dog bites the bone on one end and the other bites the other end, no dog will have more bone that the other.

Suppose the bones were electrons and the dogs were atoms. If each bone represents on electron then the two atoms will be sharing a pair of electrons. A pair of electrons shared constitutes a covalent bond.

(a) Two dogs with the same strength holding the same size of bones in their mouths gazing at each other angrily.

(b) The dogs are in a big fight.

(c) None of the dogs wants to lose the bones, hence the dogs are held together by the bones by sharing them.

Fig. 2.7: Illustration of a covalent bond

Now let us illustrate covalent bonding in a hydrogen molecule.

Formation of a hydrogen molecule (H_2)

A hydrogen molecule, (H_2), has two hydrogen atoms linked by a covalent bond.

A hydrogen atom has 1 electron in its only energy level. It is unstable. Therefore two hydrogen atoms combine by each contributing an electron. Then they share the electrons pair equally. The shared pair revolves around both toms. In effect they have the stable electron duplet arrangement, that is, first energy level with 2 electrons like helium. Fig 2.8.

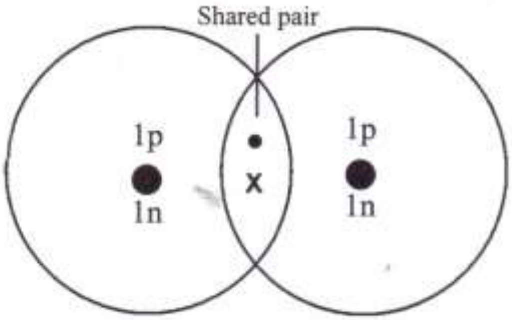


Fig. 2.8: Formation of hydrogen molecule

The shared pair is the covalent bond. It is attracted by the proton of each H atom. Sometimes the shared pair is represented by a short line (-) H-H or the pair can be shown as H∶H.

Formation of chlorine molecule (Cl_2)

A molecule of chlorine, (Cl_2), contains two chlorine atoms linked by a covalent bond.

These two chlorine 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 two atoms have a noble gas electronic structure such as argon (2.8.8(, a “shared pair” of electron is formed. Each chlorine atom contributes one electron to the shared pair, this idea of sharing can be shown in a diagram as in Fig 2.9

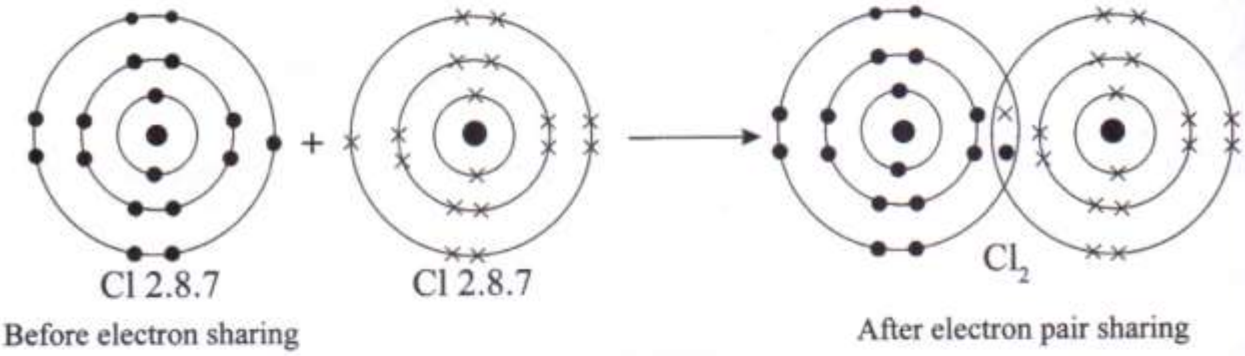


Fig. 2.9: Formation of chlorine molecule

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

Again we can see that a covalent bond is as a result of sharing two electrons, one from each atom. The nucleus of each atom attracts the shared pair strongly. Like the analogy of the two dogs (atoms) and the two bones (2 electrons) if none of the two dogs release the bo0nes, the dogs will remain together. Similarly, the two atoms will remain joined together by each atom attracting the shared pair.

Some atoms can share more than one pair of electrons. The covalence of an atom is the number of electron pairs which it shares. Let us now write simple structural formulae of some covalent compounds. The inner energy levels are omitted. If we understand how to use valency line (-) to represent a pair of electrons, we will find it very easy to draw the diagrams.

Note also that atoms of different elements can also from covalent bonds like in carbon dioxide. See Fig. 2.10. Just as we saw in ionic bonding, only outermost energy level electrons are involved in covalent bonding.

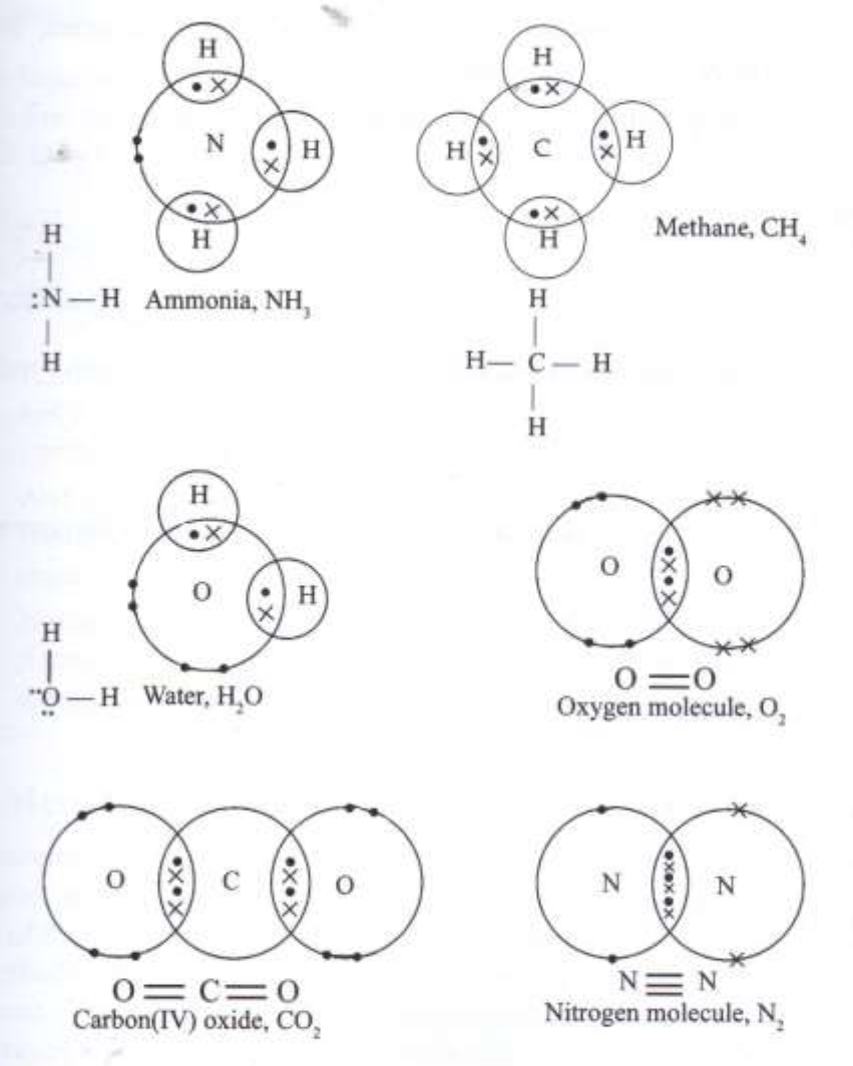


Fig. 2.10: Illustration of bonding in various molecules

Self-assessment exercise 2.3

1. Use sodium chloride and hydrogen as examples to explain what happens to electrons in the formation of ionic compounds and covalent compound respectively. (Na =11, H = 1, Cl = 17)
2. Carbon bonds with hydrogen to form methane gas, CH_4-. Silicon also bonds with hydrogen as carbon to form a compound called Silane, SiH_4-. Draw a dot (•) and cross (x) diagram showing how hydrogen atoms are bonded to S. (Si = 6, H=1)
3. Write the formula of carbon monoxide.

Writing chemical formulae using valencies

Just like we have written the formulae of ionic compounds using valencies, we can write formulae of covalent compounds using valencies as shown below.

Suppose you were asked to write the formula of potassium chloride.

Step 1

Write the symbols of the elements

K Cl

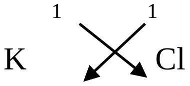
Step 2

Write the valencies of the elements above and to the right hand side of each symbol.

K^1 Cl^1

Step 3

Exchange the valencies by writing them below the symbol as shown by the arrows.



Step 4

Write the symbols close together.

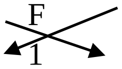
KCl

Note: one (1) is not brought down.

Other examples

a. Calcium fluoride

Symbols Ca F
Valencies 2 1



Formula Ca F

CaF_2

b. Iron (III) oxide

Symbols Fe O
Valencies 3 2
Formula Fe O
Fe_2O_3

c. Water

Symbols H O
Valencies 1 2
Formula H^1 O^2
H_2O

d. Carbon (IV) oxide

Symbols C O
Valencies 4 2
Formula C^4 O^2

Note: since valency is a common factor that can divide and form simple valencies, we divide first before exchanging the valencies.

Self-assessment exercise 2.4

1. A metallic element X has a valency of 3 and another non-metallic element Y has a valency of 1. Write the formula of the compound they form when they combine. (X and Y are not real symbols of elements.
2. An element P with atomic number 2 combines with element Q whose electronic configuration is 2.6. Write the formula of the compound they form when they react.
3. Nitrogen atom has a valency of 3 whereas hydrogen atom has a valency of 1. Ammonia gas is formed when the two atoms combine. Write the formula of ammonia.

2.3 Metallic bonding

The uttermost energy level electrons in metals are relatively few. When the atoms of the metals are closely packed, each metal loses its outer electron (s) to form a “sea” of free electrons (delocalized mobile electrons). The resulting positive metal ions are embedded in the “sea” of electrons. There is an attraction between the ions and electrons. This kind of electrostatic attraction between two positive metal ions and the delocalized electrons forms the metallic bond.

The ions arrange themselves into a giant metallic structure.

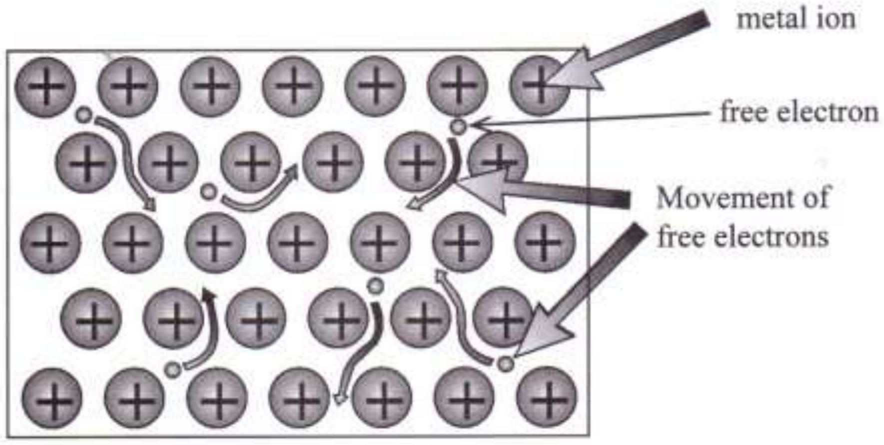


Fig. 2.11 Metallic bonding

These delocalized electrons can move on application of an electric current or heat. This explains why metals are good conductors of electricity and heat.

Self-assessment exercise 2.5

1. The electronic arrangements of Aluminium and copper are as indicated below:

Al 2.8.3
Cu 2.8.18.1
Which of the two metals is a better electricity and heat conductor? Explain your answer.

2. Describe bonding in metals

Summary

- Elements combine or bond in three main forms, namely, ionic (electrovalent) bonding, covalent bonding and metallic bonding.
- Ionic bonding occurs when a chemical bond is formed between oppositely charged ions
- Ionic bonding is as a result of donation (loss) and acceptance (gain) of one or more electrons, that is transfer of electron (s).
- The electrostatic attraction resulting from opposite charges of ions is the ionic bond.
- The charges on ions are equal to the electrons lost or gained. They are also equal to the valency of the element.

- Covalent bonding is as a result of sharing two electrons (a pair) or more than one pair of electrons by atoms.
- The greater the number of delocalized electrons (valency electrons in metal atoms) the stronger the metallic bond.
- The strength of the metallic bonds increases from left to right in a period of the Periodic Table.

Revision Exercise 2

1. Explain the following:

Although potassium is a very reactive metal and chlorine is a reactive non-metal, potassium is unreactive.

2. Draw electronic diagrams for the following species (use dot (•) and cross (x) and show only the outmost energy level electrons): water molecule (H_2O) and ammonia molecule (NH_3)
3. Two elements A and B whose atomic number are 6 and 8 respectively react to form a compound C
   a. What is the formula of C in terms of A and B?
   b. Use dots dot (•) and cross (x) to show the bonding in compound C in terms of electronic structure.
4. Explain why and how the following conduct electricity.
   a. Copper
   b. Aluminium
5. Draw electronic diagrams for the following species, use dots dot (•) and cross (x). Show only the outmost energy level electrons.
   a. Carbon dioxide (CO_2)
   b. Carbon monoxide (CO)
   c. Methane (CH_4)
6. Define a chemical bond.
7. What is the main difference between ionic bonding and covalent bonding?

Unit 3: Acids and Bases

In chemistry, substances may be described by looking at their degree of solubility pH, boiling and melting points among other factors. Classification of substances considering the pH places them into two categories; acids or bases. In this unit, we shall describe acids and bases in detail. We shall look at their properties, examples of acids and bases in daily activities, their uses and reactions among other characteristics.

3.1 Acids

The taste of oranges, lemons and citrus fruits is described as sour. The same applies to the taste of sour milk and vinegar. All these substances contain acids which give them the sharp taste of the sour taste. Some acids are found in both plant and animal materials. Such acids are referred to as organic acids. Other acids are formed from reactions of chemicals. These are called inorganic or mineral acids. Inorganic acids are common in school laboratories.

Note: the approved (IUPAC) names for these acids are at the back of the book.

A number of mineral acids are corrosive. This means that they can eat away skin, cloth and metals. It is very important that you handle these acids with great care to avoid contact with your body, clothes, books and other materials. In addition, do not taste mineral acids.

When an acid is dissolved in water the resulting solution conducts an electric current.

Tables 3.1 and 3.2 summarize examples of acids and where they are found

Table 3.1: organic acids

Name	**Where found**
Citric acid	Citrus fruits, such as oranges, lemons, *malambe*
Tartaric acid	Grapes, health salts, baking powder, *bwemba*
Lactic acid	Sour milk
Ethanoic acid	Vinegar
Methanoic acid	In ant, bee and nettle stings
Carbonic acid	Coke, lemonade, other fizzy drinks
Botanic acid	Cheese
Tannic acid	Tea

Table 3.2 Common mineral acids

Name	**Where found**
Hydrochloric acid	Found as dilute acid in the stomach. In chemicals that are used to clean metallic surfaces
Sulphuric acid	Car batteries, fertilisers, detergents
Nitric acid	Fertilisers and explosives

Uses of acids

a. Organic acids

1. Vinegar contains about 6% acetic acid. It is used in food preparation and preservation.
2. Citric acid is used in the preparation of effervescent salts and as a food preservative.
3. Tannic acid is used in the manufacture of ink and in processing leather
4. Acetylsalicylic acid which is also known as aspirin is used to relief pain, fever and reduce inflammation.
5. Ascorbic acids, also known as Vitamin C, is used an antioxidant.
6. Carbonic acid is used to make carbonated drinks i.e. soda.

b. Mineral acids

1. Sulphuric acid is used in the manufacture of fertilisers such as superphosphate and ammonia sulphate among others. It is also used in car batteries.
2. Both nitric acid and sulphuric acid are used in the manufacture of dyes, paints, drugs and explosives.
3. Hydrochloric acid is used to clean steel (a process called pickling)
4. Phosphoric acid is used to make detergents, fertilizers and soft drinks.

3.2 Bases

Drugs called anti-acids are taken to relieve acidity and heart-burn. They are examples of bases. Some bases occur naturally. A solution of plant ash in nature is a natural base. Other bases can be made in the laboratory, for example sodium hydroxide, calcium oxide among others. Bases that are soluble in water are called alkalis. For example sodium hydroxide and potassium hydroxide are alkalis. They conduct an electric current when dissolved in water. Their concentrated solutions are corrosive. Their solutions should be handled with care. However, they are some insoluble bases.

Table 3.3 summarizes the examples of bases and where they are found.

Table 3.3 Example of bases

Bases	**Where found**
Magnesium oxide	Antacid indigestion tablets
Calcium oxide	Making cement, neutralizing soil acidity
Alkalis	
Sodium hydroxide	Making soap and paper
Ammonia solution	Making fertilizer, in cleaning fluids at home

Uses of bases

Common bases include:

- sodium hydroxide
- calcium hydroxide
- magnesium hydroxide
- ammonium hydroxide
- potassium hydroxide

a. *Sodium hydroxide*
   1. used in the manufacture of soap
   2. it is used in petroleum refining
   3. used in the manufacture of medicine, paper and pulp
   4. it is used in making rayon in the textile industry

b. *Calcium hydroxide (slaked lime)*
   1. it is used in the manufacture of bleaching powder
   2. it is used to neutralize acidity in water supplies
   3. it is usually mixed with sand and water to make mortar which is used in the construction of buildings
   4. it is used to neutralize acidic soils in order to improve fertility
   5. it is used as an antidote for food poisoning
   6. it is used in the preparation of fungicides

c. *Ammonia hydroxide*
   1. used to remove ink spots from clothes
   2. used to remove grease from window-panes
   3. used in water purification

d. *potassium hydroxide*

Used in the manufacture of alkaline batteries

e. *Magnesium hydroxide*
   1. Used as a laxative (stimulation of bowel movement).
   2. Used as an antacid in treatment of heart burn.

3.3 Indicators

Look at the pictures in Fig 3.1. What can you see?

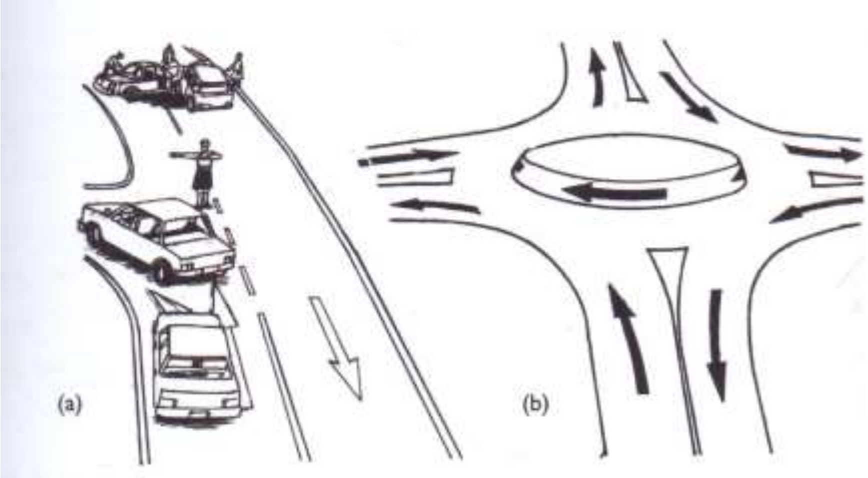


Fig. 3.1: Indicating direction

In Fig. 3.1 (a) the police woman is indicating the direction to be taken by the drivers. Also, motor vehicles are usually fitted with a lighting system that flashes to indicate the direction the driver is turning. In Fig. 3.1 (b) the arrows in the round about indicating direction a motorist should take depending on the line he or she is using.

In chemistry, indicators are equally important. We have many chemicals that appear similar to substances we come across every day. For example, we have many substances that look like water, for example, acids and bases. We use indicators to test whether they are water, acids or bases. An acid-base indicator will change to a certain colour in an acid and a different colour in a base. The acid and the base to be tested are usually in solution form.

Many plants contain dyes which can act as simple acid base indicators. They include leavers of red cabbage, petals of a flower among others.

Simple acid-base indicators

Experiment 3.1

Aim

To prepare an indicator from plant extract

Apparatus and chemicals

- Test tubes
- Pestle and mortar

- Hibiscus flower
- Leaves of acacia and tomatoes
- Propanone or ethanol;
- Water
- Aqueous solutions/suspensions of wood ash, lemon juice, soap, baking powder, anti-acid tablets and powders, toothpaste, sour milk, sodium chloride, sodium hydroxide, carbon dioxide, sulphur dioxide, sulphuric acid, hydrochloric acid.

Procedure

1. Crush some hibiscus flowers in a mortar using a pestle
2. Add a little amount of propanone or ethanol
3. Grind the petals until you get enough extract of the flower used

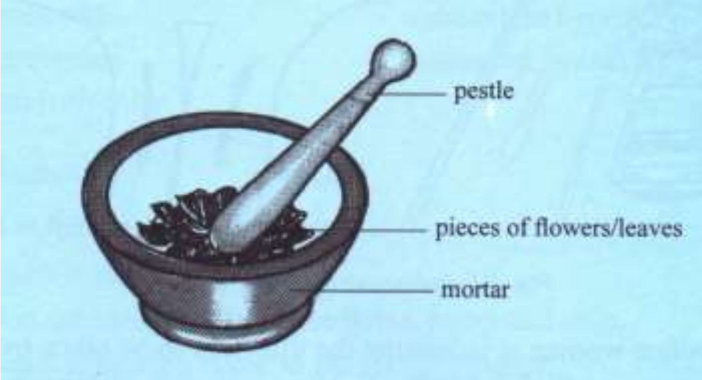


Fig. 3.2: Crushing petals

4. Filter the liquid into a clean beaker
5. Note the colour change of your filtrate. It is faint, concentrate by evaporating a little of the solvent. (Take caution to use a water bath). This I snow your laboratory made indicator
6. Half fill one test tube with dilute hydrochloric acid and another one with sodium hydroxide solution
7. Add a few drops of the indicator to each of the two test tubes and shake
8. Note and record the colour changes in a table as follows

Table 3.4 colour of indicator in acid and alkali

Colour in hydrochloric acid	**Colour in sodium hydroxide**

9. Use your indicator to test the other solutions provided.
10. Half-fill test tubes with solutions as in table 3.5. Filter the solutions where necessary to obtain clear liquids.
11. Add into each test tube a few drops of the plant extract.
12. Note the colour change of the extract. Copy Table 3.5 and record the results.

Table 3.5: colour of indicator in acid and base

Solution/suspension	**Colour of indicator**	**Acid/base**
Wood ash		
Soap		
Anti-acid tablet		
Toothpaste		
Sodium hydroxide		
Dilute sulphuric acid		
Lemon juice		
Sour milk		
Water		
Sodium chloride		

When sulphuric acid, lemon juice and sour milk were used, the colour of the plant extract changed to the same colour as when hydrochloric acid was used. This shows that these solutions are all acidic. When the plant extract was added to the following solutions; sodium hydroxide, wood ash and anti-acid tablet, the colour of the plant extract was the same as when it was added to sodium hydroxide solution. These solutions are basic.

We can therefore conclude that the plant extract indicator can be used to differentiate between acids and bases.

The substances that never changed the colour of the indicator are neither acids nor bases. Such substances are neutral. Pure water and sodium chloride are neutral.

Commercial indicators

We do not always use plant extracts whenever we want to test for acids and bases. Instead there are commercially prepared indicators that are readily available and are easy to store in a school laboratory. Commercial indicators include litmus, phenolphthalein, methyl orange, screened methyl orange and bromothymol blue.

Note: litmus is a blue vegetable compound which is extracted from plant called lichens. Litmus paper is an absorbent paper which has been dipped in litmus indicator solution and dried. It is in the form of blue or red strips of paper. You can also make indicator papers using the plant extract indicator you extracted in experiment 3.1

Experiment 3.2

Aim

To investigate colour changes of commercial indicators in acids, bases and neutral solutions

Apparatus and chemicals

- Test tubes
- Teat pipettes
- Dilute mineral acids (hydrochloric, sulphuric and nitric acid)
- Dilute base solutions (alkalis), (sodium hydroxide, ammonia solution)
- Distilled water

Procedure

1. Place seven test tubes in a rack
2. In each of these test tubes, put about 3cm^3 of dilute hydrochloric acid, dilute sulphuric acid, nitric acid, sodium hydroxide solution, calcium hydroxide solution, ammonia solution and distilled water respectively.
3. Using teat pipettes, put 2-3 drops of litmus solution or one litmus paper in each test tube.
4. Observe any colour changes
5. Repeat procedures 2 and using methyl orange, screened methyl orange and phenolphthalein indicators.
6. Copy Table 3.6 in your notebook and fill the empty spaces.

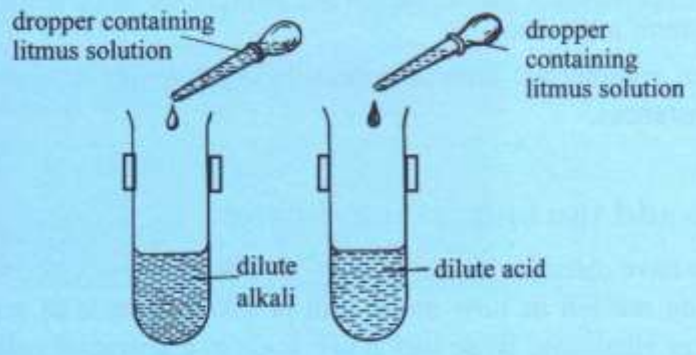


Fig. 3.3: Identifying solutions using indicators

Table. 3.6: Colour of indicator in acid and base

Table 3.6: colour of indicator in acid and base

Indicator	**Colour in acid**	**Colour in base solution (alkali)**
Methyl orange		
Litmus paper		
Screened methyl orange		
phenolphthalein		

The colour that we obtain when indicators used in this experiment are added to acids and bases are shown in Table 3.7. Compare your results of Experiment 3.2 with those in the table.

Table 3.7: colour of indicators in acids and bases

Indicator	**Colour in acid**	**Colour in base solution (alkali)**
Litmus paper	Red	Blue
Methyl-orange	Pink	Yellow
Phenolphthalein	Colourless	Pink
Screened methyl orange	Purple	Orange

Self-assessment exercise 3.1

1. Define the term indicator.
2. Name three commercial indicators and their properties in acids, bases and neutral substances

The pH scale and the universal indicator

The indicators we have discussed so far only show whether a solution is acidic or basic (alkaline). They do not tell us how much acid or alkali there is in a solution or the degree of acidity or alkalinity. To do this, a pH scale or a universal indicator is used.

Universal indicator

Universal indicator is a mixture of dyes. It gives a range of colours depending on the strength of the acid or alkali. When you use a universal indicator, you will see different types of solutions giving different colours.

The same acid of different concentrations also gives different colours with the universal indicator. Strongly acidic solutions turn universal indicator bright red for example dilute sulphuric acid. Vinegar which contains Ethanoic acid, is weakly acidic hence will turn universal indicator orange yellow. Alkaline solutions will turn universal indicator to different colours depending on their degree of alkalinity. Weak alkaline solutions turn to blue, while the strong alkaline solutions turn it violet.

The pH scale

The strength of an acid or an alkali can also be shown using a scale of numbers called the pH scale. The scale runs from 1 to 14 (Fig 3.4). When universal indicator papers are used, a pH chart with the full range of colours and numbers on is used for comparison purposes.

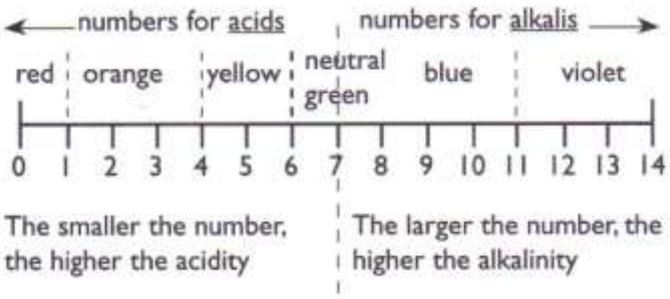


Fig. 3.4: pH scale

Note: the strongest acid has a pH of 1

The strongest alkali has a pH of 14

If a substance is neutral, it has a pH 7 (for example pure water)

Any solution with pH less than 7 is acidic and any solution with a pH more than 7 is alkaline. Therefore, when we are asked to give the acidity or alkalinity of a substance using a pH scale, we should quote the number and not the colour. Examples are given below.

Solution	**pH**
Hydrochloric acid	1.0
Rain water	6.5
Milk	6.5
Toothpaste	9.0
Pure water	7.0

Experiment 3.3

Aim

To investigate the range of pH values for acids and bases.

Apparatus and chemicals

- pH scale chart
- universal indicator solution
- dilute sulphuric acid
- lemon or orange juice
- soap solution

- test tubes
- Ethanoic acid
- sugar solution
- sodium hydroxide
- sodium hydrogen carbonate solution

Procedure

1. Put about 1 cm3 of the solutions named above in test tubes arranged in a rack.
2. Add 2-3 drops of universal indicator into each test tube
3. Compare the colour of the solution with the colours on the pH scale chart and the pH values
4. Copy table 3.8 in your notebook and record pH value of the solutions

Table 3.8: Colour of universal indicator in solution and the pH value

Solution	**Colour of the universal indicator in the solution**	**pH value**
Ethanoic acid		
Sulphuric acid		
Sugar solution		

From your results, which solutions are:

a. Acidic?
b. Alkaline?
c. Neutral?

Self-assessment exercise 3.2

1. What is a universal indicator?
2. Study the table below and answer the questions that follow.

Solution	**pH value**
A	pH7
B	pH8
C	pH4

Which solution is?

a. Acidic _____________
b. Neutral ______________
c. Basic ________________

3.4 Properties of acids

Experiment 3.4

Aim

To investigate the properties of acids.

Apparatus and chemicals

- magnesium ribbon
- zinc granules
- iron filings
- copper metal
- dilute sulphuric acid
- test tubes
- splint

a. Reaction of dilute acids with metals

Caution: never taste acids while in the laboratory. They will burn you

Procedure

1. Place magnesium, iron copper and zinc metals in four different test tubes
2. Add little dilute sulphuric acid to each metal in the test tube
3. Test the gas produced with a burning splint shown in fig 3.5. Repeat the above procedure with dilute hydrochloric acid.

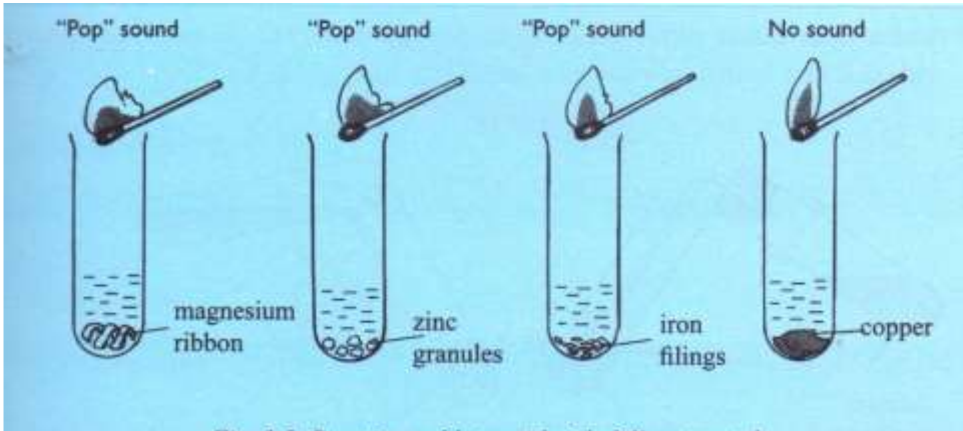


Fig. 3.5: Reacting a dilute acid with different metals

What do you observe when you add an acid to a metal?

Effervescence (bubble production) is observed when we add a dilute acid to some metals. This indicates that a gas is produced in the reaction. The gas produces a "pop" sound when burnt. This gas is hydrogen. This shows that all acids contain hydrogen. This hydrogen can be displaced (removed) by some metals. Some metals do not react with acids because they are not reactive enough. These metals include copper, silver and gold

The reaction of a metal and an acid can be summarized in a chemical equation

Metal + dilute acid ⟶ a slat + hydrogen

Note: The (+) sign before the arrow means react with and the (+) after the arrow means “and”

Examples

Magnesium + dilute sulphuric acid → magnesium sulphate + hydrogen

Zinc + dilute hydrochloric acid → zinc chloride + hydrogen

Iron + dilute sulphuric acid → iron (II) sulphate + hydrogen

b. reaction of dilute acids with metal carbonates and hydrogen carbonates

Apparatus and chemical

- calcium carbonate
- dilute nitric acid

Procedure

1. place a little carbonate or hydrogen carbonate in a test tube
2. Add a little nitric acid. What do you observe? Test the gas produced with calcium hydroxide solution as shown in Fig. 3.6.

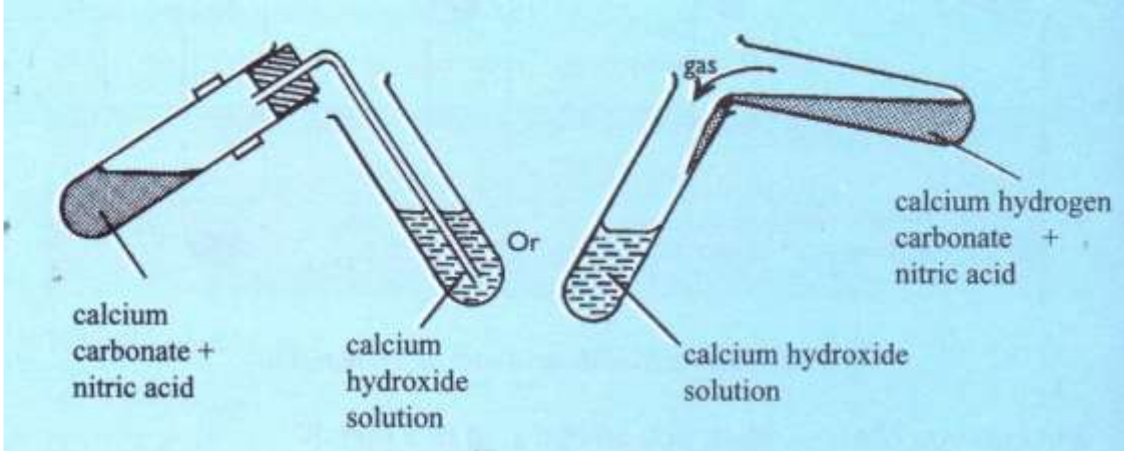


Fig 3.6: Testing for carbon dioxide

Acids effervesce (produce bubble) with metal carbonate and metal hydrogen carbonates. The general equation for this reaction is:

Metal carbonate or metal hydrogen carbonate + acid → a salt + water+ carbon dioxide

Examples include:

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

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

c. reaction of dilute acids with bases

Apparatus and chemicals

- sodium hydroxide
- universal indicator
- dilute hydrochloric acid
- test tube

Procedure

1. Put a little sodium hydroxide solution in a test tube
2. Add a few drops of universal indicator then dilute hydrochloric acid. What do you observe?

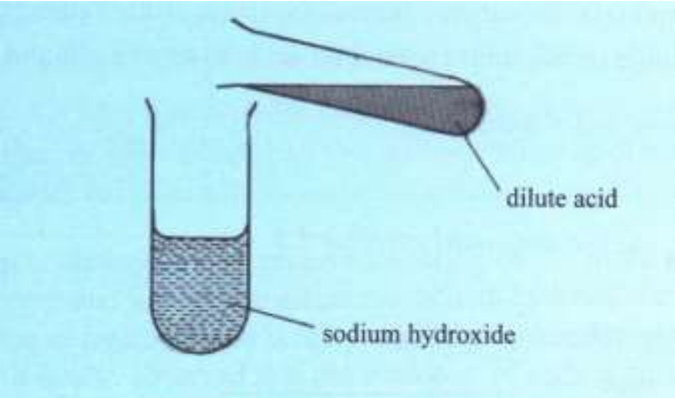


Fig. 3.7: Reacting an acid and an alkali

When bases react with acids, the strength of the acid reduces and tends to move towards the neutral pH level. The reaction of a base and an acid to produce a neutral substance is called neutralization.

Acid + base → salt + water

a. Metal hydroxide + acid → salt + water

Sodium hydroxide + hydrochloric acid → sodium chloride + water

b. Metal oxide + acid → salt + water

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

An acid has the following properties

- It is a substance with a sour taste
- It turns blue litmus paper red
- It contains hydrogen which may be placed by some metals
- It liberates carbon dioxide from carbonates and hydrogen carbonates
- It reacts with a base to give a salt and water only

3.5 Properties of bases

A base is a compound which reacts with an acid (mostly dilute) to form a salt and water only. Most metallic oxides and hydroxides are compounds which have properties of bases. Bases soluble in water are called alkalis. The common alkalis are sodium hydroxide, potassium hydroxide and ammonia solutions.

- Solution of bases (alkalis) turn red litmus blue (refer to other commercial indicators)
- All bases (soluble or insoluble) react with acids to form a salt and water only

Base + acid → salt + water

Self-assessment exercise 3.3

1. Complete the following reactions:
    a. Acid + base → ____________________
    b. Metal oxide + acid → ____________________
    c. Calcium carbonate + hydrochloric acid → ________________________________
2. Define a base

3.6 Strength of acids

Experiment 3.5

Aim

To determine the reaction of magnesium with acid of different strength

Apparatus and chemicals

- magnesium strip
- hydrochloric acid
- test tubes
- Ethanoic acid
- zinc carbonate

Procedure

1. Place the test tubes as shown in Fig 3.8
2. Put about $10cm^3$ of Ethanoic acid (acetic acid) in test tubes 1
3. Put about $10cm^3$ of hydrochloric acid or sulphuric acid in test tube 2

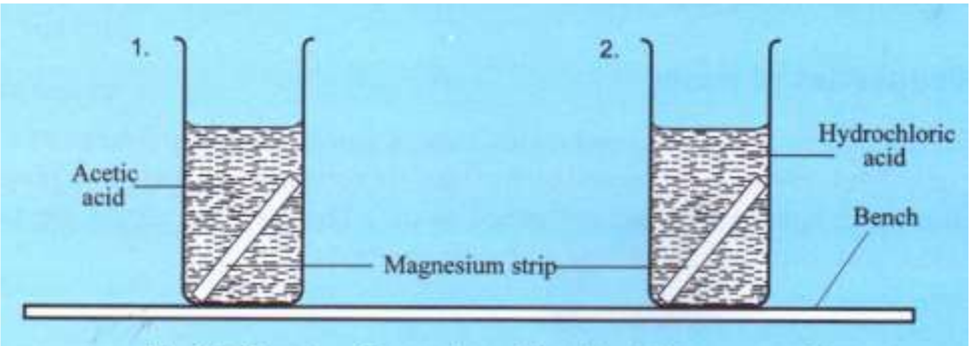


Fig. 3.8: Reaction of magnesium with acids of different strengths.

4. Clean 2 pieces of magnesium strips about 2 cm long.

Why do we clean magnesium strip first?

5. Place 2 cm long magnesium strips in each test tube at the same time
6. Repeat the experiment using 1 g of zinc carbonate instead of magnesium strips.

Magnesium strip is cleaned first to remove a coating on the metal called magnesium oxide. It is observed that when magnesium reacts with hydrochloric acid or sulphuric acid, the evolution of hydrogen gas is more rapid than when it reacts with acetic acid (Ethanoic acid). It is also observed that the evolution of carbon dioxide is more rapid than with acetic acid. This is because hydrochloric acid or sulphuric acid are strong acids while Ethanoic acid is a weak acid.

Note: the strength of an acid should not be confused with the concentration of its solution. The exact meaning of the term strength will be explained later in the course.

Examples of strong and weak acids are given in the following table.

Table 3.9: Examples of strong and weak acids

Strong acid	**Weak acid**
• Hydrochloric acid • Nitric acid • Sulphuric acid • Phosphoric acid	• Ethanoic acid (acetic acid) • Methanoic acid (formic acid) • Carbonic acid • Citric acid • Lactic acid

If the proportion of an acid is very high and that of water is very low in an acidic solution, it is said to be a concentrated acid.

If the proportion of water is very high and that of acid is very low in an acidic solution, it is said to be a dilute acid.

Self-assessment exercise 3.4

1. Name three applications of neutralization reaction.
2. Magnesium is more reactive in hydrochloric acid than in acetic acid (Ethanoic acid). Which acid is a:
   a. Weak acid?
   b. Strong acid?

3.7 Neutralization reaction

Experiment 3.6

Aim

To investigate the reaction of a dilute acid with an alkali

Apparatus and chemicals

- burette
- pipette
- conical flask
- beaker
- funnel
- evaporating dish
- Bunsen 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 3.9a. 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 3.9 b.
3. When all the alkali has been used up, the indicator suddenly turns colourless. This shows that the solution is neutral. This is the end point. There is no need of adding more acid. Look at Fig 3.9c.

4. You can tell how much acid was added, using the scale on the burette. Look at Fig 3.9d. 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 25 cm^3 of alkali is put in the flask, and the correct amount of acid added. See Fig 3.9e
6. The solution from the flask is heated to allow the water evaporate. See fig 3.9 f. you will find that dry crystals of sodium chloride are left behind.
7. Record observations in your notebook. What do you conclude?
    - What do you think would happen if we evaporated the solution obtained in step 3 to dryness?
    - Why the whole process was repeated i.e. steps 1 to 4?

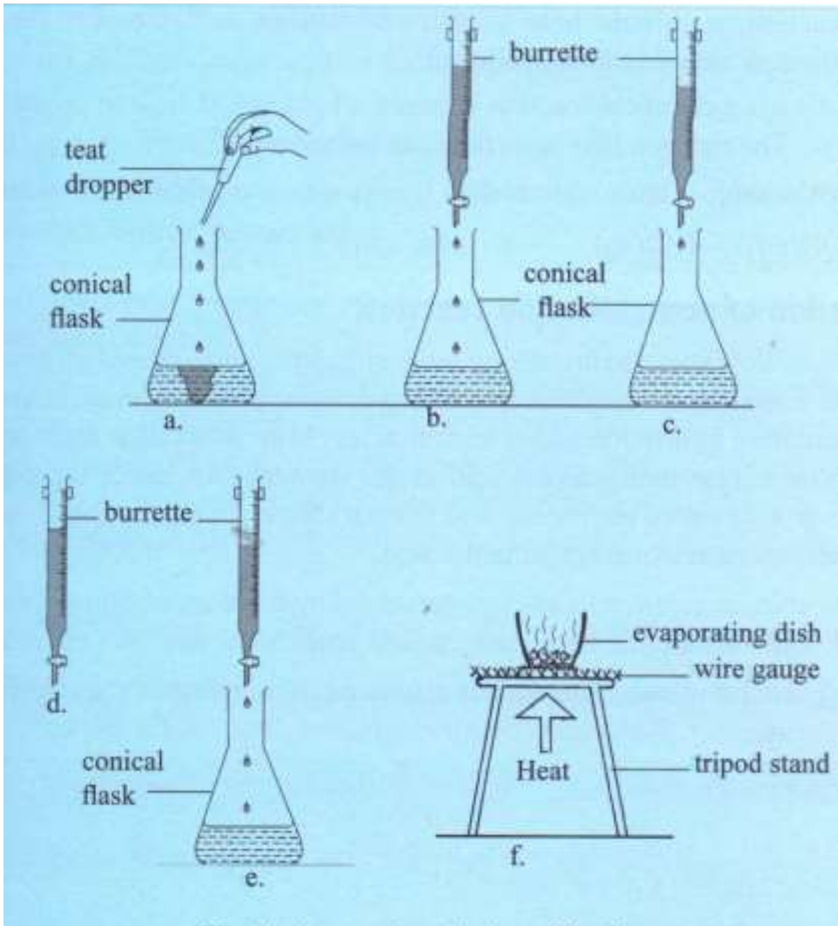


Fig. 3.9: Preparation of sodium chloride

In this experiment, the three main apparatus used are burette, pipette and conical flask. The pipette is used to transfer the alkali into a conical flask, whereas the burette is used to transfer the acid, little by little into the conical flask.

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

The experiment is repeated without an indicator to prove that the indicator does not take part in the reaction but only aids in colour change.

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. The filtrate is then evaporated to get a colourless salt.

In this reaction, a soluble base is reacted with an acid through the process of neutralization. A soluble salt is produced.

Neutralization 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

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

Application of neutralization reaction

1. Health salts: excess hydrochloric acid in gastric juice causes stomach ulcers. It can be neutralized by excess sodium hydrogen carbonate, magnesium hydroxide or Aluminium hydroxide added in anti-acids. Milk is alkaline and can be used to neutralize excess hydrochloric acid in the stomach. An insect for example injects an acid called Methanoic acid when it stings. An alkali rubbed on the insect's bite relieves pain by neutralizing the acid.
2. Excess acids in acidic soils can be neutralized by frequent addition of calcium oxide (quick lime) or calcium hydroxide (slaked lime) in the soil.
3. Tooth paste for mouth cleaning has alkalis which neutralizes the acid which causes teeth decay.

Summary

- An acid has a sour taste, turns blue litmus paper red, phenolphthalein turns colourless and has a pH of below 7
- Acids react with bases to form a salt and water only
- Acids reacts with some metals to form a salt and a hydrogen gas
- Acids present in plant and animal materials are called organic acids
- Bases are metal oxides and hydroxides
- Soluble bases form solutions called alkalis. These alkalis react with acids to form a salt and water only
- An alkali turns litmus paper blue, phenolphthalein pink and has a pH of above 7
- Some plant extracts can be used as indicators

- A solution which is neither acidic nor basic is described as neutral, for example pure water
- Universal indicator gives a range of colours depending on the strength of an acid or an alkali
- The pH scale ranges from 1 to 14. The lower the pH the higher the acidity. The higher the pH the higher the alkalinity of the liquid or solution.
- Strong acids have a low pH while weak acids have a high pH. Weak bases have a low pH while strong bases has a high pH.
- Strong acids include sulphuric acid, hydrochloric acid and nitric acid
- Weak acids include acetic acid, and all organic acids
- Strong bases include sodium hydroxide potassium hydroxide and their metal hydroxides.
- The reaction between acids and bases is called neutralization
- Weak bases include ammonium hydroxide, calcium hydroxide solution (lime water) and magnesium hydroxide solution

Revision exercise 3

1. Define the following terms:
    a. An acid
    b. Base
    c. Neutralization
2. What is an indicator as used in chemistry? Name two common indicators found in the laboratory and give their colours in an acidic solution.
3. What is an alkali? Name two common alkalis. Write a word equation to show how the named alkali reacts with dilute hydrochloric acid
4. Name two metals that react with dilute sulphuric acid. Write a word equation only to show how they react with dilute sulphuric acid.
5. Name the acids found in:
    a. Sour milk
    b. Orange juice
    c. Vinegar
    d. Anti-stings
    e. Fizzy drinks
6. Five solutions were tested with universal indicator and their pH values recorded in the table below.

Solution	**pH value**
A	11.0
B	2.0
C	6.0
D	7.0
E	12

Which solution is?

a. Most acidic?
b. Least acidic
c. Neutral
d. Most basic?
e. Least basic?

7. Copy and complete the following word equations.
   a. Acid + metal oxide →
   b. Dilute sulphuric acid + magnesium →
   c. Acid + carbonate →
   d. Dilute hydrochloric acid + sodium carbonate →
8. Solutions can be classified as strong acids or weak acids and strong bases or weak bases. The list below shows solutions and their pH values.

Solution	**pH values**
S	14.0
T	2.5
U	8.5
V	6.5
W	7.0

   a. Classify solutions S to was either:
      I. Strong or weak acid
      II. Strong or weak base
   b. Which solution is likely to be:
      I. Sodium hydroxide?
      II. Lemon juice
      III. Rain water?
   c. Select any pair that react to form a neutral solution.
9. A bee sting contains Methanoic acid. This is why it is so painful. How can you treat your sister or brother if he/she is stung by a bee?
10. State the colours of the following indicators in acids and base (solutions)

Indicator	**Acids**	**Bases (solutions)**
Litmus		
Methyl orange		
phenolphthalein		

11. Magnesium reacts with an acid usually found in the stomach to form salt A and gas B the:
      I. Acid
      II. Salt A
      III. Gas B
   c. Write a word equation for the reaction between this acid and a base, (copper (II) oxide).

UNIT 4:HYDORCARBONS

4.1 Organic Chemistry and Hydrocarbons

Carbon has the ability to bon with other carbon atoms to form stable compounds with long chains. This property of carbon is known as catenation. Carbon also bond with other atoms of different elements such as hydrogen, oxygen, among others. When carbon bonds with hydrogen only, compounds formed are called hydrocarbons. Hydrocarbons are organic compounds with a general formula –C_xH_y.

When carbon bonds with oxygen only, other compounds called oxycarbons are formed. Oxycarbons are inorganic compounds with general formula –C_xH_y. Examples of oxycarbons are carbon dioxide (CO_2) and carbon monoxide (CO). Oxycarbons are mainly gases at room temperature.

Carbon-containing compounds, such as bottled liquid petroleum gas (LPG), kerosene, petrol among others, are studied under a branch of Chemistry called organic chemistry. Such compounds are known as hydrocarbons, and were formed over one million over years by slow decomposition of plant and animal remains, that were trapped under layers of rock. In the absence of oxygen and under high pressure, crude oils was formed. Why these compounds are called hydrocarbons? As the name suggest, hydrocarbons are compounds which contain hydrogen and carbon only. A hydrocarbon has the molecular formula, C_xH_y where x and y represent whole numbers. This means carbon can form molecules that contain any number of carbons atoms linked by strong covalent bonds.

Crude oil is a mixture of hydrocarbons. The hydrocarbons in crude oil are important fuels and raw materials for many important products such as plastics and fabrics.

Apart from hydrocarbons and oxycarbons, there are other organic compounds such as those form:

- Vegetable sources for example carbohydrates, oils, textiles and rubber
- Animal sources for example proteins and fats

All food substances, except water, are composed of organic compounds. Manufactured substances for example drugs, dyes, pharmaceuticals, paper, ink and pain also contain organic compounds.

Homologous series

Research has shown that certain organic compounds have similar properties and may sometimes behave similar ways. Such compounds are grouped together to form a homologous series. A homologous series is a group of compounds which have the following characteristics:

- The compounds can be represented by a certain general formula
- The compounds differ from one member to the next by a constant atom or group of atoms e.g. CH_{-2}.
- The compounds can be prepared using similar methods
- The compounds show gradual change in physical properties
- They exhibit similar chemical properties.

Examples of homologous series in organic chemistry include:

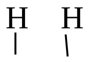
- Alkanes
- Alkenes
- Alkynes
- Alkanols
- Alkanoic acids

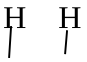
At this level, we shall study only two hydrocarbons groups include:

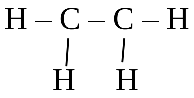
- Alkanes – carbon atoms are linked by single covalent bonds
- Alkenes- carbon atoms are linked by at least one double covalent bond.

Saturated and unsaturated hydrocarbons

As mentioned earlier, carbon has the ability to bond with four other atoms and form a long chain. When carbon bonds with three hydrogen atoms, in a chain, the hydrocarbon that results is said to be saturated i.e. it has a maximum number of hydrogen atom attached to it. When a carbon atom is bonded to less than three hydrogen atoms, it is said to be unsaturated. Look at the structures below.

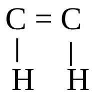






$$\begin{matrix} H - & C & - & C & - H \\ & | & & | & \\ & H & & H & \end{matrix}$$

A.



$$\begin{matrix} C & = & C \\ | & & | \\ H & & H \end{matrix}$$

B.

Each carbon in A is bonded to a maximum number of hydrogen atoms i.e. three. In B each carbon atom is bonded to two hydrogen atoms. A is therefore a saturated hydrogen while B is an unsaturated hydrocarbon.

Experiment 4.1

Aim

To test for unsaturation

Apparatus and chemicals

- Test tubes
- Acidified potassium permanganate solution
- Teat pipette
- Bromine water

Procedure

1. Bubble ethane through acidified potassium permanganate solution
    - What do you observe?
    - Record your observation in your notebook.
2. Bubble ethane through bromine water
    - What do you observe
    - What can you conclude from the above observations?
3. Record your observations and conclusions in your notebook
4. What would happen if we bubbled ethane through bromine water?

Ethane decolourises potassium permanganate solution as well as bromine water. The decolouration of these solutions is the test for the presence of a multiple bond; may either be a double or triple bond.

Note: ethane cannot decoulorise potassium manganite (VI) solution or bromine water because it lacks either double or triple bonds.

Therefore, alkanes are saturated hydrocarbons whereas alkynes are unsaturated. Alkynes which are known to have a least one triple bond between two carbon atoms have at least one triple bond between two carbon atoms are also unsaturated hydrocarbons.

A homologous series is a group of compounds in which all the members:

- Differ from each other by – CH_2 – group
- They have similar chemical properties
- Can be prepared using a general method
- Exhibit a gradual along the series in physical properties such as melting point, boiling point and solubility

4.2Alkanes

There is a variety of hydrocarbons in crude oil. Most of them are alkanes. Examples of alkanes which are found in crude oil are methane, ethane, propane and butane. These four alkanes are gases. Crude oil also contains petrol and kerosene.

From activity 4.1 on counting the total number of carbon and hydrogen atoms, you realize that the previous compound differs from the next by a – CH_2- group. This shows that the compounds form a homologous series. The compound in the series are known as alkanes.

Alkanes are hydrocarbons with the general formula CnH2n where n denotes the number of carbon atoms in each molecule. The number, n= 1, 2, 3…etc.

The simplest alkane has only one carbon atom in its molecule. From the general formula shown above, we can get the formula and structure of the alkane. For example, where number of carbon atoms is one, i.e. n=1 the formula is obtained by substituting the value of 'n' in the formula, i.e. $C1H_{2x1\ +2'}$ which gives CH4. This is the formula of methane; the first member of alkane group.

Similarly, we can drive the second alkane, where n=2, i.e. $C_2\,H_{(2x2)\ +2} = C_2H_{4+2} = C_{12}H_6$ hence we get C_2h_6. This alkane is called ethane.

Likewise for n=3

$C_3H_{(3x2)+2} = C_3H_{6+2}$

= C3H8, we get C3H8 which is the third member of alkane group known as propane etc.

Self-assessment exercise 4.1

Use the general formula of alkanes to derive the molecular formula of alkanes with:

a. Four carbon atoms
b. Five carbon atoms
c. Six carbon atoms

Naming of alkanes

We have derived the molecular formula of the first three straight chain alkanes which are:

- Methane –CH4
- Ethane – C2H6
- Propane – C3H8

Probably, you still remember from your Form 1 Mathematics, names given to figures with different numbers of sides. For example:

- Five ____________ pentagon
- Six ____________ hexagon
- Seven ____________ heptagon
- Eight ____________ octagon

From these examples, you note that the prefixes used i.e. pent,-hex, hept and oct correspond to the number of sizes. The same applies to naming of alkanes. ‘Met’ means ‘one’, ‘eth’ means two, etc. the following table gives a summary of the prefixes used for the various alkanes with different numbers of carbon atoms.

Table 4.1: summary of prefixes of alkanes

Prefix	**Number of carbon atoms**
Meth	For 1 carbon atoms
Eth	For 2 carbon atoms
Prop	For 3 carbon atoms
But	For 4 carbon atoms
Pent	For 5 carbon atoms
Hex	For 6 carbon atoms
Hept	For 7 carbon atoms
Oct	For 8 carbon atoms
Non	For 9 carbon atoms
Dec	For ten carbon atoms

The names of alkanes are derived by first writing a prefix and then adding the ending “ane”. For example an alkane with:

- One carbon atom is methane
- Two carbon atoms is ethane
- Three carbon atoms is propane
- Four carbon atoms is butane

Following the same procedure, write the names of straight chain alkanes with five, six, seven, eight, nine and ten carbon atoms. Refer your work to Table 4.2

You may have noted that the last three letters, ane, form the last part of the groups/class names listed. This indicates that the compounds belong to the class of hydrocarbons called alkanes.

Formulae of alkanes

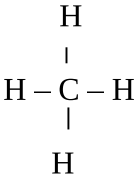
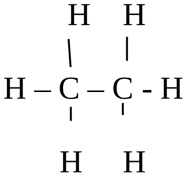
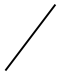
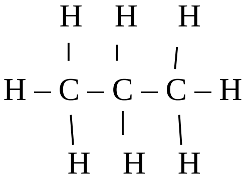
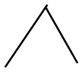
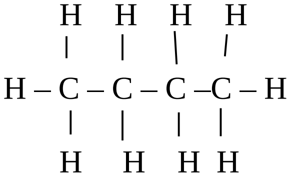
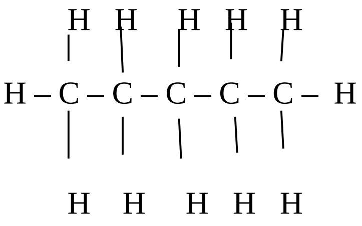
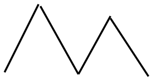
The four types of formulae are:

- Molecular formula

- Structural formula
- Condensed formula
- Skeletal formula

Molecular formula simply shows the number of atoms involved in bonding while structural formula shows how the atoms bond in a compound in addition to giving the atoms involved in bonding. Condensed formula on the other hand categorizes atoms in groups depending on how they occur in the structural formula. The skeletal formula is a zigzag line representation of the bonding in carbon atoms. The start and end of each line represen6ts a carbon atom. Table 4.2 gives a summary of the four types of formulae for the first three alkanes.

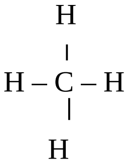
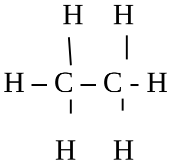
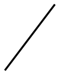
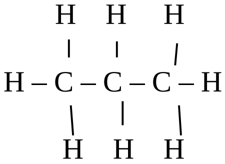
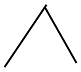
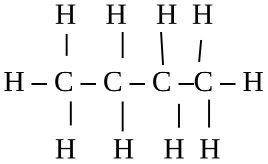
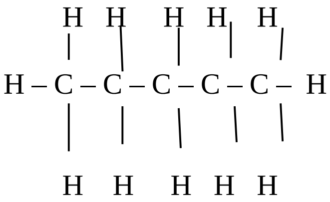
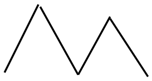
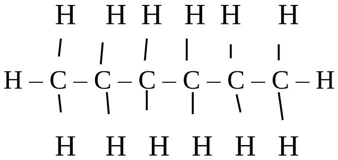
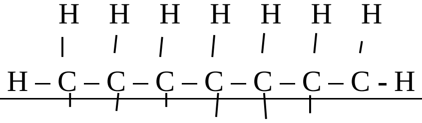
Table 4.2 formula for first five alkanes.

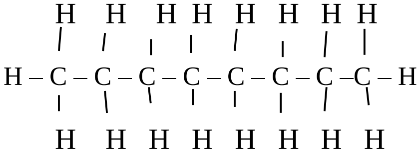
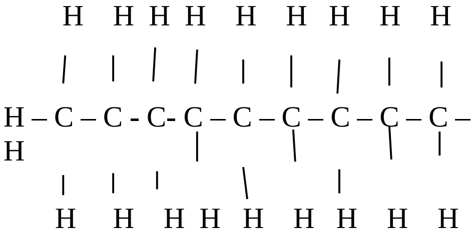
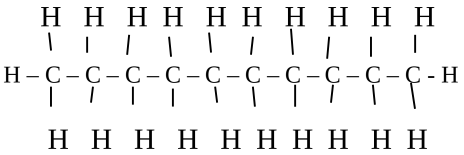
Alkane	**Molecular formula**	**Structural formula**	**Condensed formula**	**Skeletal formula**
Methane	CH_4	 $H-\overset{H}{\underset{H}{C}}-H$	CH_4	
Ethane	C_2H_6	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$	CH_3 CH_3	
Propane	C_3H_8	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$	CH_3 CH_2 CH_3	
Butane	C_4 H_{10}	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$	CH_3 $(CH_2)_2$ CH_3	
Pentane	C_5H_{12}	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$	CH_3 $(CH_2)_4$ CH_3	

Self-assessment exercise 4.2

Write the molecular formula and draw the structural formula of straight chain alkanes with six, seven, eight, nine and ten carbon atoms.

Table 4.3 formula of the first ten straight chain alkanes

Number of carbon atoms	**Alkane**	**Molecular formula**	**Structural formula**	**Condensed formula**	**Skeletal formula**
1	Methane	CH_4	 $H-\underset{H}{\overset{H}{C}}-H$	CH_4	
2	Ethane	C_2H_6	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ CH_3$	
3	Propane	C_3H_8	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ CH_2\ CH_3$	
4	Butane	$C_4\ H_{10}$	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ (CH_2)_2\ CH_3$	
5	Pentane	C_5H_{12}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ (CH_2)_3\ CH_3$	
6	Hexane	C_6H_{14}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ (CH_2)_4\ CH_3$	
7	Heptane	C_7H_{16}	 $H-\overset{H}{C}-\overset{H}{C}-\overset{H}{C}-\overset{H}{C}-\overset{H}{C}-\overset{H}{C}-\overset{H}{C}-H$	$CH_3\ (CH_2)_5\ CH_3$	

			H H H H H H H		
8	Octane	C_8H_{18}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ (CH_2)_6\ CH_3$	
9	Nonane	C_9H_{20}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ (CH_2)_7\ CH_3$	
10	Decane	$C_{10}H_{22}$	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	$CH_3\ (CH_2)_8\ CH_3$	

Self-assessment exercise 4.3

a. Write down the structural formula of pentane,
b. Draw the skeletal formula (a) above.

Sources of alkanes

Alkanes are mainly obtained through fractional distillation of crude oil.

Crude oil is a complex mixture of hydrocarbons. It is a thick black liquid with a strong smell. It is formed from the decay of remains of small animals and plants under layers of rocks over millions of years. Unless its components are separated, crude oil does not have much use. This is dine as a refinery. During fractional distillation at are refinery the mixture of hydrocarbons is sorted out into groups or individual hydrocarbons called fractions. A fractionating column is used to separate the mixture into various components as shown in Fig 4.1

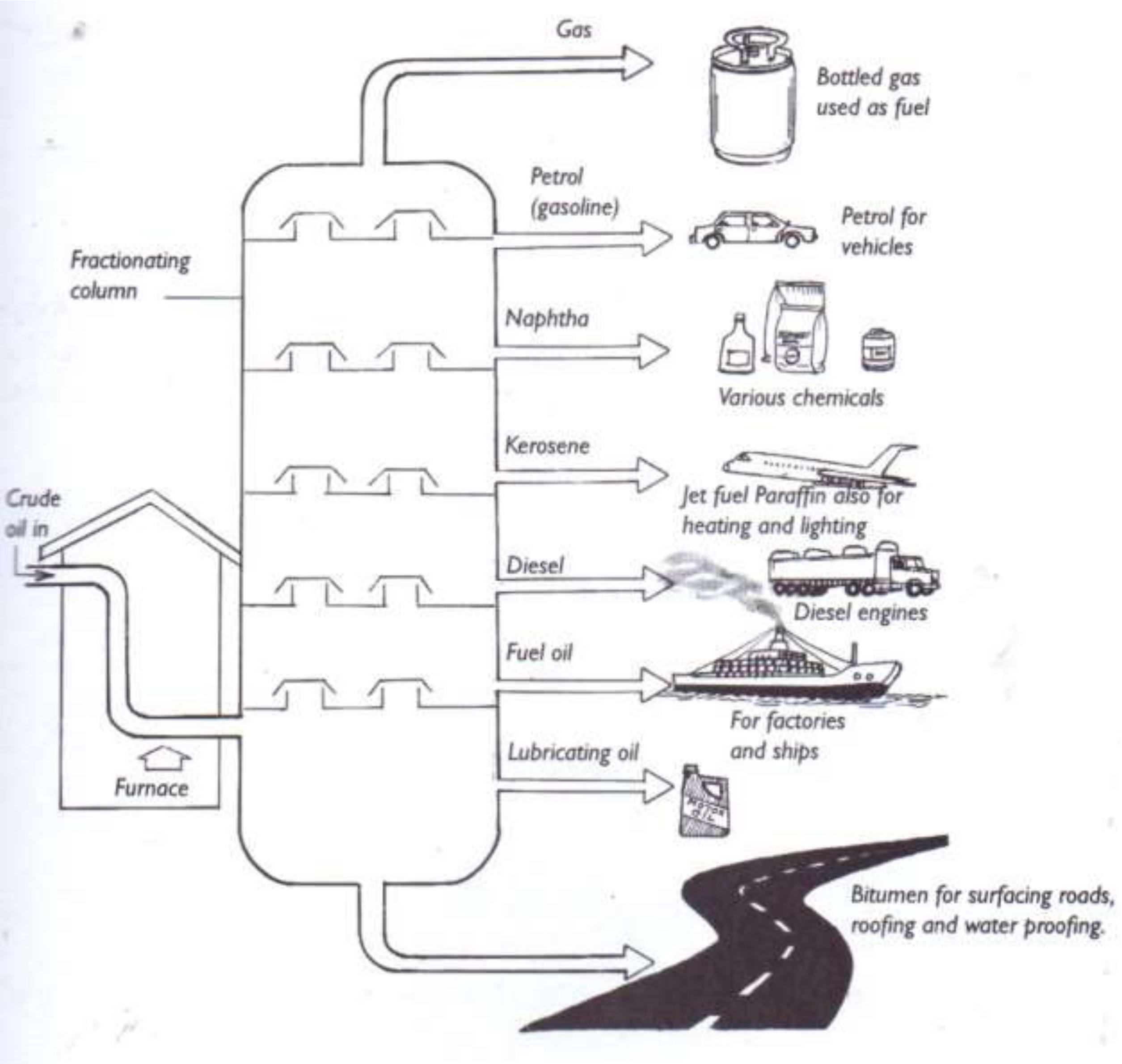


Fig. 4.1: Fractional distillation of crude oil in a refinery

The crude oil is first heated up in a furnace. As it is heated, the small molecules boil off first. They enter the column as a gas. The fractionating column is hot at the bottom and cooler at the top. The gas molecules then condense. The larger hydrocarbons have higher boiling points. This means that the larger hydrocarbons, with the high boiling points, turn back to liquids easily nearer the bottom.at high temperatures, the hydrocarbons are in form of gases. They hence rise up at the column. The different fractions condense and are connected at different levels as shown in Fig 4.1. The last fraction obtained is called bitumen. It is used to surface roads.

Self-assessment exercise 4.4

1. Write down the name, molecular formula and structural formula of an example of a hydrocarbon.
2. In **purification** of crude oil, the oil is **fractionally** distilled.
    a. Explain the meanings of the words in bold
    b. Write down the names of the first and last fractions during fractional distillation of crude oil.

iii. What are the uses of the first and last fractions?

Preparation and properties of alkanes

Earlier, we learnt that members of the same homologous series have similar physical and chemical properties. They can also be prepared using similar methods. In this section, we shall learn about preparation and properties of methane as a representative of alkanes.

Experiment 4.2

Aim

To prepare and investigate chemical properties of methane gas.

Apparatus and chemicals

- Hard glass test tube/ round bottomed flask
- Delivery tube
- Beehive shelf
- Gas jar
- Source of heat
- Trough
- Mortar and pestle
- Sodium ethanoate
- Soda lime (a mixture of sodium hydroxide and calcium oxide0

Note: it is easier to handle soda lime than the deliquescent sodium hydroxide. It does not dissolve easily.

Procedure

1. Place about 4-5 g of sodium ethanoate and an equal amount of soda lime in a mortar and grind well with a pestle
2. Transfer the mixture into a hard glass test tube
3. Set up apparatus as shown in Fig 4.2
4. Heat the test tube carefully. Make sure that the water is not sucked back by removing the delivery tube from water immediately after heating is stopped.

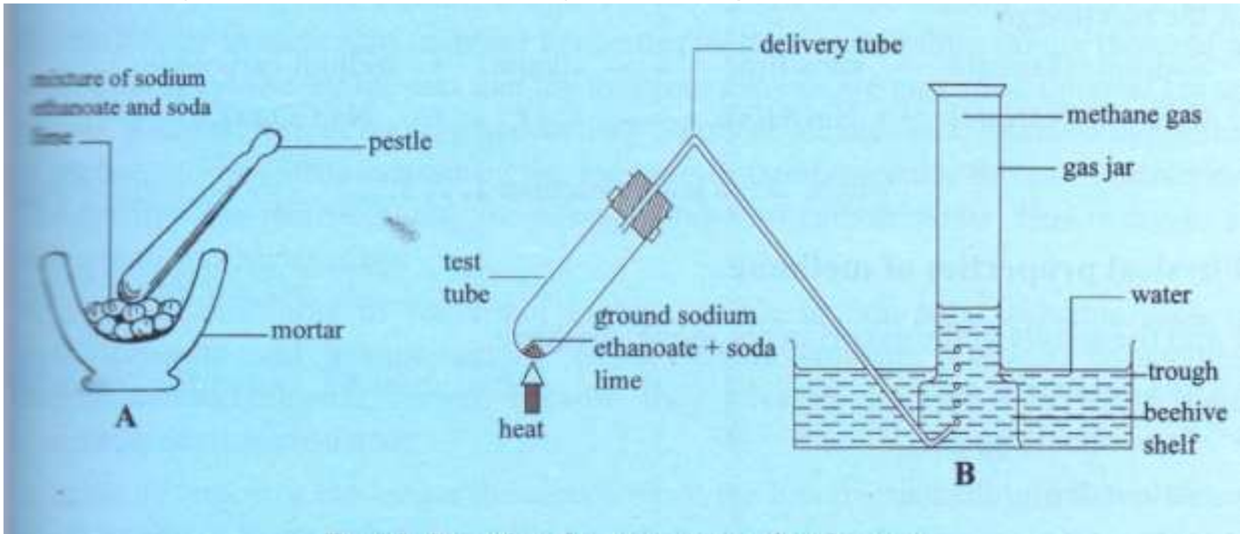


Fig.4.2: Procedure for preparation of methane gas

- What do you observe?
- What is the smell of the gas collected?
- What physical property of the gas is tested by the method of collection used?
- What can you conclude from your observations?
- Record your observations and conclusions in your notebook.

5. Collect three test tubes full of this gas for carrying out the tests shown in Table 4.4. Copy the table 4.4. And record your observations in the spaces provided.

Table 4.4: observations on chemical reactions of methane

Test-tube	**Test**	**Observations**
1	Introduce a lighted splint at the mouth of the test tube containing the gas	
2	Invert a test tube full of methane in a beaker containing methylbenzene	
3	Add about 4-5 drops of bromine water	

- What can you conclude about methane?
- Write equations for the reactions that take place.

Methane is prepared in the laboratory by heating a mixture of sodium ethanoate and soda lime. The sodium hydroxide in the mixture reacts with sodium ethanoate to form methane gas and sodium carbonate. The equation for the reaction is:

Sodium ethanoate + soda lime $\longrightarrow$ methane + sodium carbonate

$CH_3COONA(S) + NaOH\ (S) \longrightarrow CH_4\ (g) + Na_2CO_3\ (S)$

Note: Generally any alkane can be prepared by similar reactions. The general equation of the reaction is:

Sodium alkanoate + soda lime $\longrightarrow$ alkane + sodium carbonate

$C_nH_{2n+1}COONA(S) + NaOH\ (S) \longrightarrow C_nH_{2n+2} + Na_2CO_3\ (S)$

Where n = a whole number 1, 2, 3…

Physical properties of methane

Form the results of the experiments above, methane:

- Is a colourless gas
- Is an odourless gas
- Is less denser than air
- Is insoluble in water but soluble in organic solvents such as benzene, ether among others

Note: the first four alkanes are gases. The next six are liquids.

Table 4.5 below shows the trends in physical properties of alkanes

Table 4.5 trends in physical properties of alkanes.

No. of carbon atoms	**Name of alkane**	**Melting point (^0C)**	**Boiling point (^0C)**	**Density g/cm3**	**Physical state of alkanes at room temperature**	**Solubility in water**	**Solubility in organic solvent**
1	Methane	-182	-161	0.424	Gas	insoluble	soluble
2	Ethane	-183	-88	0.546	Gas	Insoluble	Soluble
3	Propane	-188	-42	0.501	Gas	Insoluble	Soluble
4	Butane	-138	0	0.579	Gas	Insoluble	Soluble
5	Pentane	-130	36	0.626	liquid	Insoluble	Soluble
6	Hexane	-95	69	0.657	Liquid	Insoluble	Soluble
7	Heptane	-90	99	0.684	Liquid	Insoluble	Soluble
8	Octane	-57	126	0.703	Liquid	Insoluble	Soluble
9	Nonane	-53	151	0.718	Liquid	Insoluble	Soluble
10	Decane	-29	174	0.730	liquid	insoluble	soluble

You may have noted that the trends in physical properties of alkanes are mainly determined by the carbon chain length. As length of the carbon chain increases the molecular masses also increase hence the melting and boiling points increase are well. This is evident by the fact that the first four alkanes are gases and the next six alkanes are liquids. The increase in melting and boiling points of alkanes with increase in number of carbon atoms is also caused by the increase in intermolecular forces of attraction. The density also increase with increasing number of carbon atoms. This is due to an increase in molecular mass.

Alkanes are insoluble in water but highly soluble in non-polar solvents such as methylbenzene and hexane among other organic solvents. Organic compounds do not conduct electric current because they have no ions in their liquid state. The same applies to alkanes.

In terms of viscosity, the longer the carbon chain the less freely the liquid flows. Hence the viscosity of hydrocarbons increase with increase in molecular mass and tendency of the long molecules to become tangled up with one another like a thread does. Kerosene has a low viscosity, followed by petro; then diesel, oil is fairly viscous whereas grease is highly viscous.

Note: Viscosity refers to how freely a liquid flows. If a liquid has a high viscosity it implies that the liquid does not flow easily and vice versa.

Chemical properties of alkanes

In general, alkanes are less reactive compared to other hydrocarbons. The strong carbon-carbon and carbon hydrogen covalent bonds make the alkanes relatively more stable hence less reactive.

Alkanes burn in air or oxygen. The first four which are gases burn easily with a yellow flame when ignited. Some liquid form like petrol and kerosene also burn with a yellow sooty flame when ignited.as the carbon atom chains increase in length, the harder it becomes to ignite them and the more sooty the flame becomes. For example, it is more difficult to ignite wax and bitumen than paraffin.

When alkanes burn in air or sufficient oxygen, they produce carbon dioxide and water. However, if they burn in insufficient oxygen, they produce carbon monoxide and water. When oxygen is very little carbon and water are produced.

Let us now study two important chemical reactions, using methane as a representative of alkanes.

a. Combustion of methane

Methane burns in air or sufficient oxygen to produce carbon dioxide gas and water.

$$\text{Methane + oxygen} \longrightarrow \text{carbon dioxide + water}$$

$$CH_4\,(g) + 2O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(l)$$

This reaction produces a lot of heat hence methane is a good source of fuel. If there is insufficient oxygen for complete combustion, methane produces carbon monoxide and water.

Methane + insufficient oxygen ⟶ carbon dioxide + water

$$2CH_4\,(g) + 3O_2\,(g) \longrightarrow 2CO\,(g) + 4H_2O\,(l)$$

Methane + limited oxygen ⟶ carbon + water

$$CH_4\,(g) + O_2\,(g) \longrightarrow C\,(s) + 2H_2O\,(l)$$

If there is very limited oxygen supply, soot is produced. Soot is essentially the element –Carbon (C).

b. Substitution reaction of methane

Alkanes also react with halogens in the presence of sunlight.

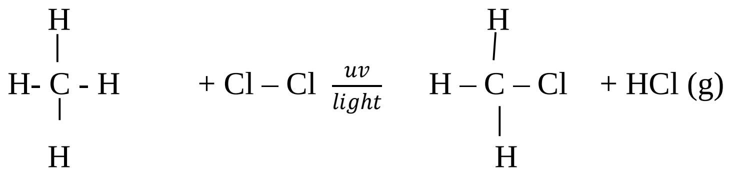
Note: Alkanes do not react with halogens in the d ark. This is because there is no enough energy which is necessary to start the reaction.

Energy is needed to break the halogen bond in the reactants first before a reaction can occur. This initial energy comes from sunlight. Consider the reaction below between chlorine and methane.

$$\text{Methane + chlorine} \xrightarrow[light]{uv} \text{chloromethane + hydrogen chloride}$$

$$CH_4\,(g) + Cl2\,(g) \xrightarrow[light]{uv} CH3\ Cl\,(g) + HCl\,(g)$$

Or



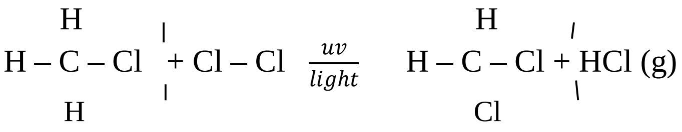
$$H\text{-}\overset{\displaystyle H}{\underset{\displaystyle H}{C}}\text{-}H + Cl-Cl \xrightarrow[light]{uv} H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-Cl + HCl\,(g)$$

In this reaction, one of the chlorine atoms reacts with methane and replaces (substitutes) a hydrogen atom in the methane molecule. This kind of reaction is called a substitution reaction. Hence we say methane has undergone substitution reaction.

In the presence of excess chlorine, further substitution of the remaining hydrogens can occur. Finally all the hydrogen atoms are substituted by chlorine atoms to form tetra chloromethane.

- Chloromethane + Chlorine $\frac{uv}{light}$ dichloromethane + hydrogen chloride

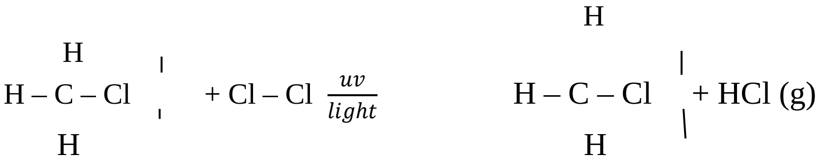
CH_3 Cl (g) + Cl2 (g) $\frac{uv}{light}$ CH_2Cl_2 (g) + HCl (g) or



$$\mathrm{H-\underset{H}{\overset{H}{C}}-Cl\ +\ Cl-Cl\ \frac{uv}{light}\ H-\underset{Cl}{\overset{H}{C}}-Cl\ +\ HCl\ (g)}$$

- Dichloromethane + Chlorine $\frac{uv}{light}$ trichloromethane + hydrogen chloride

CH_2 Cl_2 (g) + Cl_2 (g) $\frac{uv}{light}$ $CHCl_3$ (g) + HCl (g) or

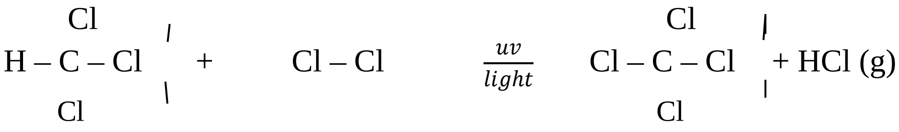


$$\mathrm{H-\underset{H}{\overset{H}{C}}-Cl\ +\ Cl-Cl\ \frac{uv}{light}\ H-\underset{H}{\overset{H}{C}}-Cl\ +\ HCl\ (g)}$$

- Trichloromethane + chlorine $\frac{uv}{light}$ tetrachloromethane + hydrogen chloride

$CHCl_3$ (g) + Cl_2 (g) $\frac{uv}{light}$ CCl_4 (l) + HCl (g)

Or



$$\mathrm{H-\underset{Cl}{\overset{Cl}{C}}-Cl\ +\ Cl-Cl\ \frac{uv}{light}\ Cl-\underset{Cl}{\overset{Cl}{C}}-Cl\ +\ HCl\ (g)}$$

Note: in tetrachloremethane; all the four hydrogen atoms in methane (CH_4) have been replaced with chlorine atoms hence the formula CCl_4. Bromine can substitute the hydrogen atoms in methane in a similar way in the presence of sunlight as shown below.

Other alkanes can undergo similar substitution reactions. Follow the procedure in the examples above and write equations for the reactions between ethane and chlorine.

Self-assessment exercise 4.5

1. What do you understand by the term hydrocarbon?
2. Name the third alkane in homologous series.
3. Alkanes do not conduct electricity. Explain why.
4. Write a chemical equation for the combustion of methane.

Uses of alkanes

Fig. 4.3 gives a summary of the uses of various alkanes.

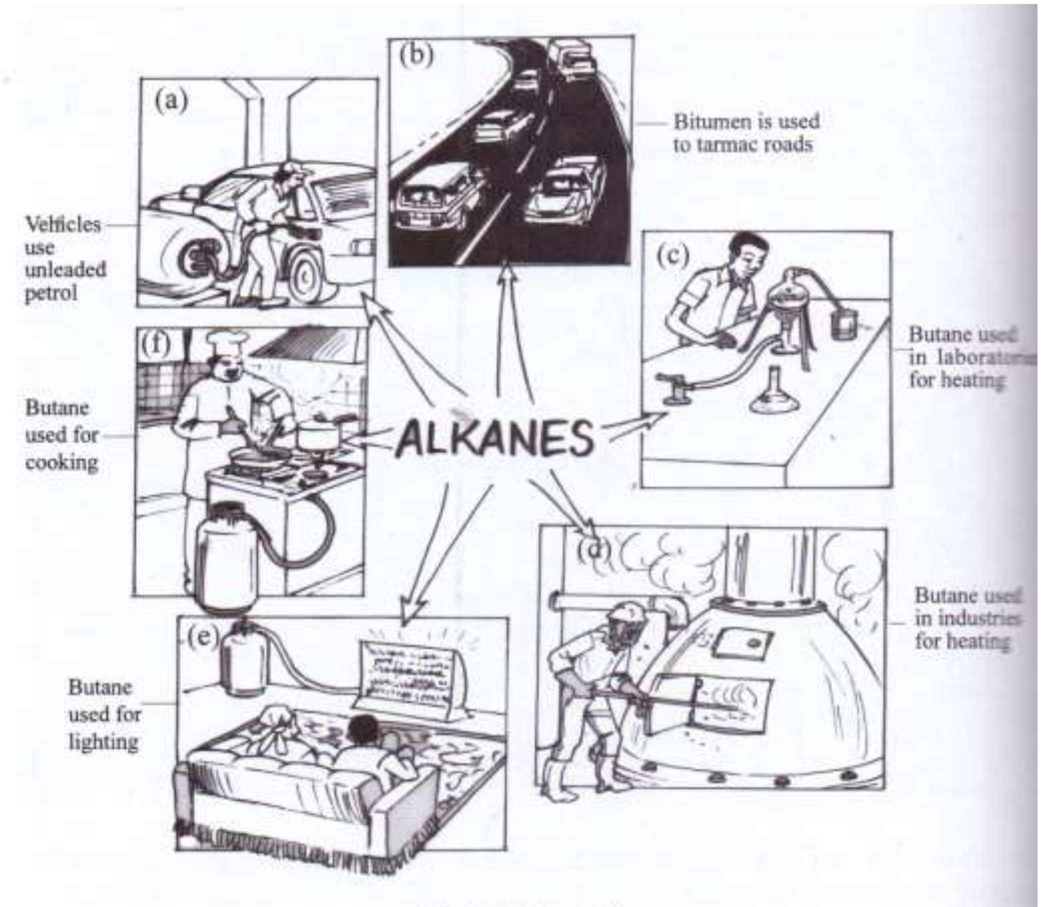


Fig. 4.3: Uses of alkanes

1. Methane (CH_4), also known as biogas, is used as a source of fuel for cooking. It is also used as fuel in vehicles in some countries as a substitute for petrol.
2. Propane (C_3H_8) and butane (C_4H_{10}) which are liquid components of liquidified petroleum gas (LPG) are used in school laboratories and in industries for heating purposes. Fig 4.3 (a,b,c and d)
3. Butane (C_4H_{10}) mixed with small proportions of propane (C_3H_8) are compressed in gas cylinders and used in household lighting, water heating and cooking. See Fig 4.3 e and f.
4. Petrol is used in automobiles all over the world. Fig 4.3a Petrol is mainly heptane (C_7H_{16}) but other compounds like tetraethyllead (IV), Pb (CH_3CH_2)$_4$ are added to create more smoothly burning fuel. Another additive is 1, 2 –dibromoethane (CH_2BrCH_2Br). This prevents lead from accumulating in the engine. Unfortunately this results in formation of lead bromide which is emitted from car exhaust systems thereby adding lead to the atmosphere, causing pollution. Lead is very poisonous. It can cause damage to the brain and nervous system especially in young children. This is the reason why it is nowadays recommended that motorists uses unleaded petrol.
5. Alkanes with higher numbers of carbons are used as solvents in the manufacture of industrial chemicals.
6. Solid alkanes have other uses. For example, bitumen is used to tarmac roads, Fig. 4.3(b). Petroleum jelly for example (Vaseline) is a mixture of paraffin wax 9a solid alkane) and oil etc.
7. Candle burns to give light. Candle wax is a hydrocarbon. The wick in a candle gives a surface over which the molten wax vaporizes and is able to be ignited.

Self- assessment exercise 4.6

1. What is the meaning of the term homologous series?
2. State three uses of alkanes.
3. (a) State the conditions necessary for substitution reaction of methane with bromine.

(b) Write an equation to show how hydrogen atoms are substituted.

4.3Alkenes

Alkenes differ from alkanes in that they have at least one double bond between two adjacent carbon atoms. They are examples of unsaturated hydrocarbons. Their general formula is (C_nH_{2n}) where n is an integer. Unsaturated means they have double covalent bonds between any two carbon atoms. The first member of alkene group is ethane (C_2H_4) with n = 2.

Note: there is no alkene where n = 1.

Table 4.4 shows the first alkenes and their formula

Naming of alkenes

We have already learnt how Alkanes are named. A similar method can be used to name alkenes. However, the easiest method of naming alkenes is by replacing the ‘a’ in the corresponding alkane with an “e”. For example,

- Ethane gives ethene
- Propane gives propene
- Butane gives butane
- Pentane gives pentene
- Hexane gives hexene etc.

The names of alkenes end with –ene.

Table 4.6: the first ten alkenes.

Check on the formula of alkenes. What difference do you notice from one formula to the next? They differ by a $-CH_2$ group, hence from a homologous series.

Self-assessment exercise 4.7

1. Explain briefly why methane does not exist.
2. Differentiate between alkanes and alkenes.
3. Name the following alkenes:
   a. CH3 CH2 CH = CH2
   b. CH3 CH = CHCH3
4. Write the structural formula for the following alkenes:
   a. Hexane
   b. Pentene
5. Name the following compounds.
   a.
   b. CH3Ch = ChCh3
   c. CH3 (CH2)2 CH = Ch2
6. Draw the structural formula of the following compounds
   a. Butane
   b. Propene
   c. Butane
7. What is the difference between saturated and unsaturated hydrocarbon?

Sources of alkenes

Alkenes can be obtained in two different ways:

- Dehydration of alcohols
- Cracking hydrocarbons

a. Dehydration of alcohols

Dehydration refers to the process of removing water from a compound. In most cases alcohols are dehydrated using concentrated acids such as sulphuric acid or phosphoric acid. The alcohol is heated using excess concentrated acid which converts it to an alkene and water. The overall equation of the reaction is:

$$\underset{\textit{Alcohol}}{C_n H_{2n+1} OH} \xrightarrow[\textit{heat}]{\text{conc acid}} \underset{\textit{alkene}}{C_n H_n} + \underset{\textit{water}}{H_2O}$$

Experiment 4.3

Aim

To prepare ethene by dehydration of ethanol.

Apparatus and chemicals

- Thermometer
- Sand bath
- Cork rubber stopper
- Delivery tubes
- Beehive shelf
- Concentrated sulphuric acid
- Bromine water
- Acidified potassium manganite (vii)
- Round bottom flask
- Boiling tube
- Bunsen burner
- Trough
- Gas jar
- Ethanol
- Broken porcelain/dry sand

Caution: *concentrated sulphuric acid is corrosive. Ethanol and ethene are flammable. There is substantial production of sulphur dioxide in this reaction which can cause irritation and even trigger an asthmatic attack. This experiment should therefore be done in a fume chamber or in an open space.*

Procedure

1. Place about 20cm3 of ethanol in a round bottomed flask and slowly add about 40cm3 concentrated sulphuric acid while cooling and shaking the flask.
2. To the round bottomed flask, add 2-3g of clean broken pieces of porcelain or sand and assemble the apparatus as shown in Fig 4.4
3. Heat the flask gently over a water bath. What do you observe?
4. Allow the gas produced to escape from the delivery tube for a few minutes. Then collect samples of this gas in boiling tubes.
5. Close the tube with a cork, then remove the cork and light the gas.

If the mixture contains ethane, it burns with a blue flame: if ethane is mixed with air, the mixture in the test tube ignites with an explosion; so take care! Allow more gas to escape from the delivery tube until you prove that it is pure ethane then collect several gas jars of ethane.

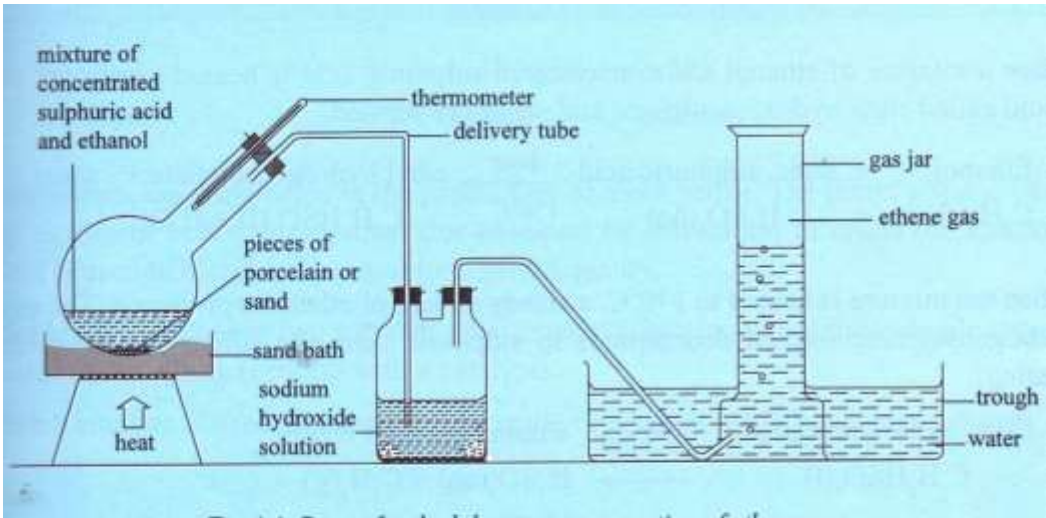


Fig.4.4: Set-up for the laboratory preparation of ethene gas

- The first potion of the gas evolved is simply displaced hot air from apparatus.
- Ethene is flammable and it is important not to ignite the gas from the delivery tube

6. Disconnect the delivery tube connecting the round bottomed flask to the trough before you stop heating.
    - Why must the tube be disconnected before you stop heating?
7. Bubble the gas through the test tubes containing:
    a. Bromine water
    b. Acidified potassium manganite (VII). Record your observations.
    - What do you conclude?
8. The residue in the flask should be disposed of by first allowing it to cool, and then pouring it into a large volume of cold water

Questions

1. Why is sulphuric acid added into the flask while cooling and shaking the glass?
2. Hwy should clean, dry broken porcelain or sand be added into the mixture of concentrated sulphuric acid and ethanol?
3. Why is it advisable to heat the acid gently?
4. Why does the gas ignite with an explosion if it is not pure ethane?
5. Why must the delivery tube be disconnected from the round bottomed flask before the heat is turned off?
6. Why should the residue in the flask be first poured into a large volume of cold water when disposing it?

When the mixture of ethanol and concentrated sulphuric acid is heated gently; an oily liquid called ethyl hydrogensulphate and water are formed.

Ethanol + conc. Sulphuric acid $\xrightarrow{\text{heat}}$ ethyl hydrogensulphate + water

$$C_2H_5OH\ (l) + H_2SO_4\ (aq) \xrightarrow{170} C_2H_5HSO_4\ (l) + H_2O\ (l)$$

When the mixture is heated to 170^0C, a steady stream of ethane is produced. The ethyl hydrogensulphate formed decomposes to sulphuric acid and ethene gas on further heating.

Ethyl hydrogensulphate $\longrightarrow$ sulphuric acid + ethane

$$C_2H_5HSO_4\ (l) \longrightarrow H_2SO_4\ (aq) + C_2H_4\ (g)$$

Ethanol contains elements that make up water i.e. hydrogen and oxygen hence water is removed in the above reaction. When concentrated sulphuric acid reacts with this water, a lot of heat is produced that is, the reaction is extremely exothermic. It is therefore necessary to cool the flask while this reaction is being carried out. Ethanol and concentrated sulphuric acid are miscible although they have different densities. Shaking the flask helps to mix the two liquids properly. Dry broken porcelain increase the surface area on which gas bubbles can form. They also ensure smooth boiling of the mixture.

The mixture of ethanol and the acid is heated gently to prevent a violent reaction which can cause large amounts of sulphuric acid to jump off the flask into the delivery tube. Pure ethane burns with a blue flame. Otherwise, if mixed with air, it explodes. That is why the first potion o the gas collected is allowed to escape.

The delivery tube should first be disconnected from the round-bottomed flask before we stop heating to prevent water from being sucked back into the flask. The sulphuric acid remaining in the flask as a residue should be diluted by pouring it into a large volume of cold water. This is to dilute the acid so as to make it less corrosive.

Note: by dehydrating different alcohols, different alkenes are produced. The general equation for the production of alkenes through dehydration of alcohol is:

$$\text{Alcohol} \xrightarrow[-\ \text{Water}]{\text{Conc. Sulphuric acid}} \text{Alkene}$$

$$C_nH_{2n+1}OH \xrightarrow[-\ H_2O]{\text{Conc. } H_2SO_4} C_nH_{2n}$$

b. Cracking of hydrocarbons

Cracking involves breaking long-chain alkane molecules into shorter alkanes and an alkene. For example, when hexane is cracked, butane and ethane are obtained, i.e.

$$\text{Long alkane} \xrightarrow{700^0C} \text{shorter alkane + alkene}$$

$$\underset{\text{Hexane}}{C_6H_{14}\,_{(1)}} \longrightarrow \underset{\text{butane}}{C_4H_{10}\,(g)} + \underset{\text{ethene}}{C_2H_4\,(g)}$$

In industries, cracking helps in the production of more petrol. The petrol obtained this way is also of better quality than that obtained by distillation of crude oil. Cracked petrol is used to blend other petrol to improve equality.

Cracking can be done in two ways, **thermal cracking** (heating like in the example above) or **catalytic cracking** (heating with a catalyst).

When a catalyst is used, cracking can be made to occur at fairly low temperatures.

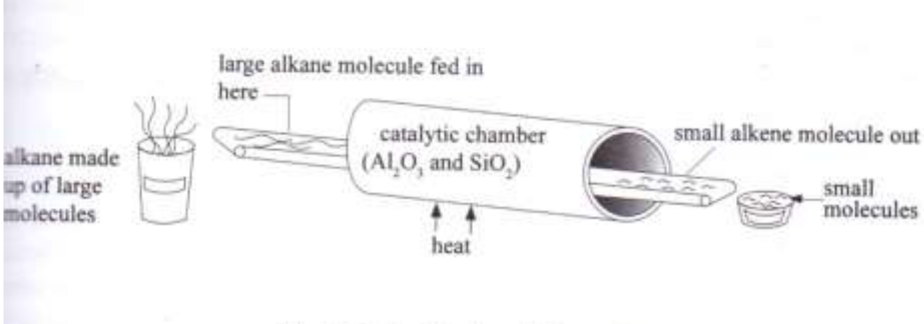


Fig. 4.5: Cracking large hydrocarbons

Physical properties of alkenes

Trends in physical properties of alkenes

Physical properties of ethane

- It is colourless gas

- It has no smell
- It is insoluble in water

Table 4.7 shows a summary of physical properties of the first five alkenes. What can you observe about the trends in the physical properties of alkenes?

Table 4.7: trends in physical properties of alkenes

Name of alkene	**formula**	**Melting point ^{0}C**	**Boiling point ^{0}C**	**Density (g/cm3)**	**Physical state at room temperature**	**Solubility in water**	**Solubility in organic solvent**
Ethane	C_2H_4	-169	-104	-	gas	insoluble	soluble
Propene	C_3H_6	-185.2	-47.7	-	gas	insoluble	soluble
Butane	C_4H_8	-185.3	-6.2	-	gas	insoluble	soluble
Pentene	C_5H_{10}	-138	30.0	0.640	liquid	insoluble	soluble
Hexane	C_6H_{12}	-98	63.9	0.674	liquid	insoluble	soluble

- There is increased melting and boiling points as the number of carbon atoms increase. This is because with increasing carbon atoms the molecular mass increases. This also causes an increase in the intermolecular forces of attraction. To break these intermolecular forces, more energy is required hence the increase in melting and boiling points.
- The first three alkenes i.e. ethene, propene and butene are gases at room temperature while pentene and hexane are liquids. This also has to do with increase in intermolecular forces of attractions.
- Alkenes are organic compounds hence they are insoluble in water but soluble in organic solvents.

Chemical properties of ethane

A double bond between atoms is a stronger link than a single bond. However the second bond is weaker than a single bond; in terms of energy needed to break bonds. As such, the second bond is easier to break. This bond easily opens up and takes additional atoms. The bond is therefore largely responsible for the chemical reactions undergone by ethane.

Let us now consider the individual reactions of ethane.

a. Combustion of ethane

Ethene is highly flammable and burns readily in air to form carbon dioxide and water.

$$\text{Ethane} + \text{oxygen} \longrightarrow \text{carbon dioxide} + \text{water}$$

$$C_2H_4\ (g) + 3O_2\ (g) \longrightarrow 2CO_2\ (g) + H_2O\ (l)$$

b. Addition reaction of ethane

I. *Addition of hydrogen (catalytic hydrogenation)*

Ethane reacts with hydrogen in the presence of nickel catalyst to form ethane. A temperature of 150^0C is required.

Ethene + hydrogen $\xrightarrow[150^0C]{\text{nickel catalyst}}$ ethane

C2H4 (g) + h2 (g) → C2H6 (g)

OR

This process is also called hydrogenation and because a catalyst is used, it is catalytic hydrogenation.

II. *Hydration*

When ethane is react with concentrated sulphuric acid, it forms a compound called ethylhydrogensulphate. When added to water and warmed an alcohol called ethanol is formed.

We can summarize the reaction as follows:

- Reaction of ethane with sulphuric acid

Ethane + sulphuric acid → ethyl hydrogen sulphate + ethanol + water

- Reaction of ethane with steam

Ethane reacts with steam to form an alcohol called ethanol which is a hydrocarbon with an oxygen atom. This is the method of production of industrial alcohol.

$CH_2 = CH_2$ (g) + H_2O (g) → CH_3 CH_2OH (l)

c. Addition of halogens to ethane

Halogens readily add across a double bond of alkenes to form compounds called dihalides. This process is called halogenation. We may represent a halogen molecule with X2 or X-x, as shown in the example below.

General reaction of ethane with a halogen.

Note: when naming the product formed, in which a hydrogen atom is substituted with a halogen atom, the prefix is borrowed from the name of the halogen where –ine part is replaced with “o” as follows:

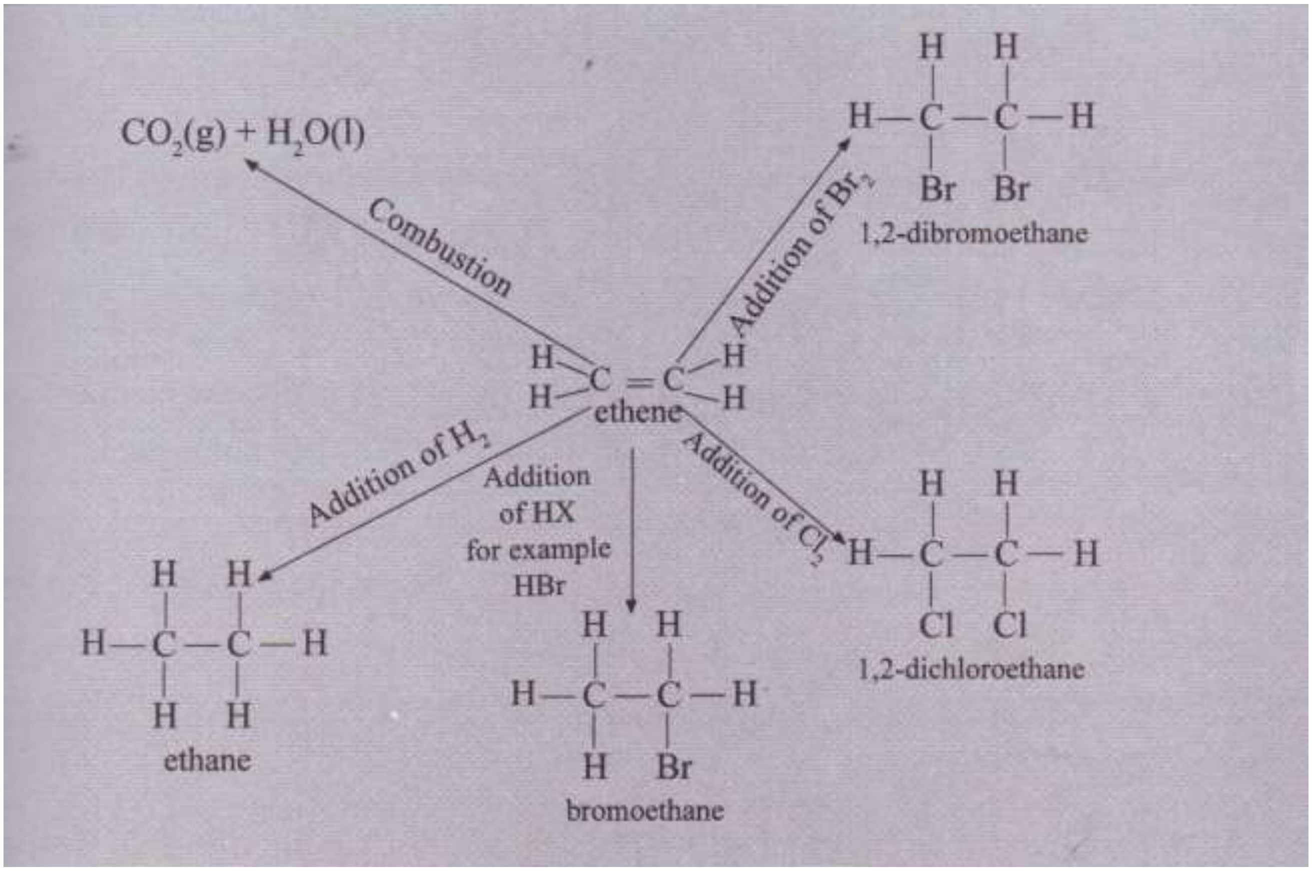
- From chlorine to chloro
- Bromine to bromo
- Iodine to iodo
- Therefore chlorine reacts with ethane to form 1, 2-dichloroethane.

Bromine reacts with ethane to form 1, 2-dibromoethane.

Bothe 1, 2-bromoethane and 1, 2-dibromoethane are colourless. Ethane is instantly decolourises both bromine and chlorine. The decolourisation of bromine is a test for unsaturated hydrocarbons.

Follow the above procedure and write the equation and structures of the reactants and products formed in the reaction between ethane and iodine.

Summary of reactions of ethane



Uses of alkenes

Some uses of alkenes are given in Fig 4.6

1. Ethane and propene are used in the manufacture of plastics. These plastics are made through addition polymerization. Polythene is a plastic used to make dustbins, bags, electrical insulators and clothing. Fig 4.6a, b c, d, e, f and g.
2. Ethane is used in artificial ripening of fruits such as mangoes, bananas, among others fig 4.6d.

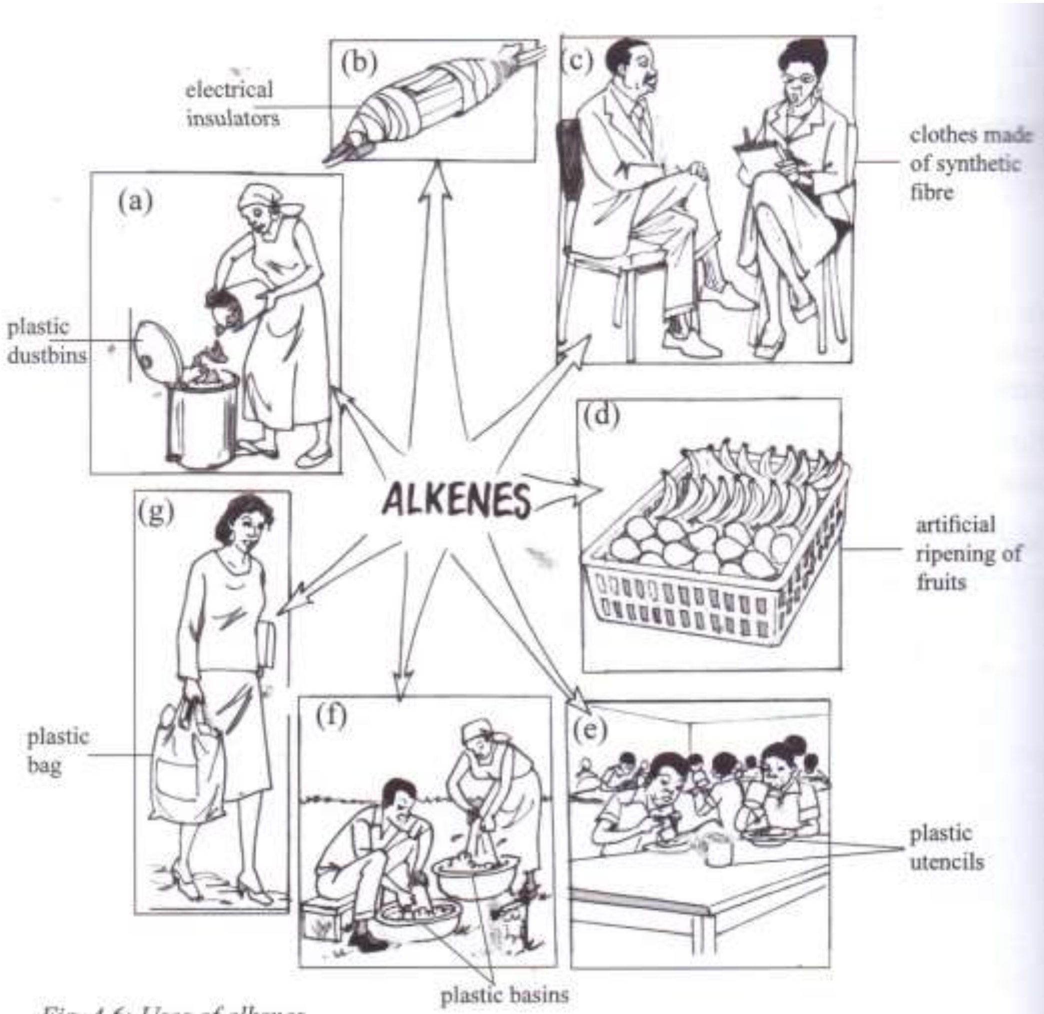


Fig. 4.6: Uses of alkenes

Self-assessment exercise 4.8

1. An alkene has a melting point of -1400C and boiling point of 640C. Would you expect this alkene to be a solid, liquid or gas at room temperature (250C)? Explain.
2. Draw and name the structural formula of pentene.
3. Write an equation to show the complete combustion of ethane in oxygen.
4. Draw a structural formula to show how ethane reacts with chlorine to produce 1, 2-dichloroethane.
5. Write an equation to show how ethane reacts with hydrogen to produce ethane.
6. Name three uses of alkene.

Summary

- Organic chemistry is the study of carbon compounds
- The carbon atoms join together to form a great variety of carbon compounds
- A homologous series is made up of organic compounds within the same family ascending with a-CH2 group.
- Members of a homologous series:

  - ➢ Can be represented by a general formula, for example; alkane CnH2n+2, alkenes Cnh2n
  - ➢ Can be prepared by similar methods
  - ➢ Exhibit similar properties
  - ➢ Have similar chemical properties
  - ➢ Exhibit a gradual change in physical properties as melting point, boiling point and solubility.
- Cracking is the heating or catalytic breaking of large alkane molecules to produce shorter molecules
- Alkanes are saturated hydrocarbons which contain single carbon to carbon (C-C) bonds
- Alkenes are unsaturated hydrocarbons containing double carbon to carbon (C=C) bonds
- Alkanes undergo substitution reactions
- Alkenes undergo addition reactions.

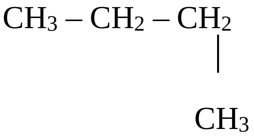
Revision exercise 4

1. What is the difference between a structural formula and a molecular formula? Show this difference using methane and ethane.
2. Study the following formulae:
   A. C_4H_8
   B. $CH_3\ CH_2CH_2CH_3$
   C. C_3H_6
   D. C_6H_{20}
   E. $C_{10}H_{22}$

a. Which of these compounds are:
   I. Alkanes?
   II. Alkenes?

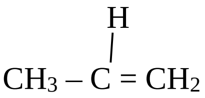
b. Name each compound listed above.

3. Draw a structural formula for CH_3Br. From which alkane is this derived from?
4. What is the difference between saturated and unsaturated hydrocarbons? Give an example of each.
5. Name the following compounds.
   a. $CH_3 – CH_2 – CH_2 – CH_2 – CH_3$
   b.



$$\underset{\displaystyle CH_3}{CH_3 – CH_2 – \underset{|}{CH_2}}$$

   c. $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
   d.



$$CH_3 – \overset{\displaystyle H}{\overset{|}{C}} = CH_2$$

   e. $CH_3CH=CHCH_3$

6. Write a balanced equation for each of the following reactions.
   A. Burning of butane
   B. The reaction between hex-2-ene and hydrogen
   C. The addition of bromine to but-1-ene
7. What is the purpose of the fractional distillation of crude oil?

B.What is basic principle that is used in fractional distillation?

c. how can the process be demonstrated in the laboratory?

8. Nature gas consists mainly of methane and ethane. Name the homologous series to which these components belong.
9. Name two sources of:
   a. Alkanes
   b. Alkenes
10. Explain the difference between physical properties of methane and propane.
11. How are the substitution reactions in alkanes different from the addition reactions in alkenes? Give relevant examples.

UNIT 5: AIR

Air is the most common substance known to human beings. 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.

5.1 Percentage composition of dry air by volume

The table below shows the composition of air.

Table 5.1: percentage composition of air

Substance	**Percentage (%)**
Nitrogen	78
Oxygen	21
Carbon dioxide	0.03
Noble gases	About 1
Smoke/dust particles	Variable
Polluting gases	Variable

Table 5.1 shows that air is a mixture of several gases. Oxygen constitutes 21% of air.

5.2 Separation of components of air by fractional distillation of liquid air

The best way to obtain oxygen, nitrogen and noble gases on large scale is to separate them from liquid air. This is done through the process called fractional distillation of liquid air. Remember that air is a mixture of gases which include nitrogen, oxygen, carbon dioxide, noble gases and water vapour. The substances to be removed first are water vapour, dust particles and carbon dioxide. If water vapour and carbon dioxide are not removed, they would form solids at low temperatures and block the pipes.

Next, the air is compressed at about 200 atmospheres. This makes the air hot (just like air gets hot when we pump it into a bicycle tube). It is then allowed to expand through a jet which make sit get very cold and some return into liquid. Compression and expansion are repeated several times and each time air gets colder. When the temperature reaches -200^{0}C, the gases, nitrogen, oxygen and argon become liquid, neon and helium gases remain and are removed.

Since nitrogen and oxygen have different boiling points, they are separated by fractional distillation. When liquid air is warmed slowly, liquid nitrogen which has a lower boiling point of -196^{0}C distills first and can be stored under pressure in steel cylinders. The remaining liquid is very rich in oxygen. On further heating, argon whose boiling point is -186^{0}C distills leaving oxygen which has higher boiling point of -183^{0}C. Once separated, the two gases are also stored and sold commercially in steel cylinders.

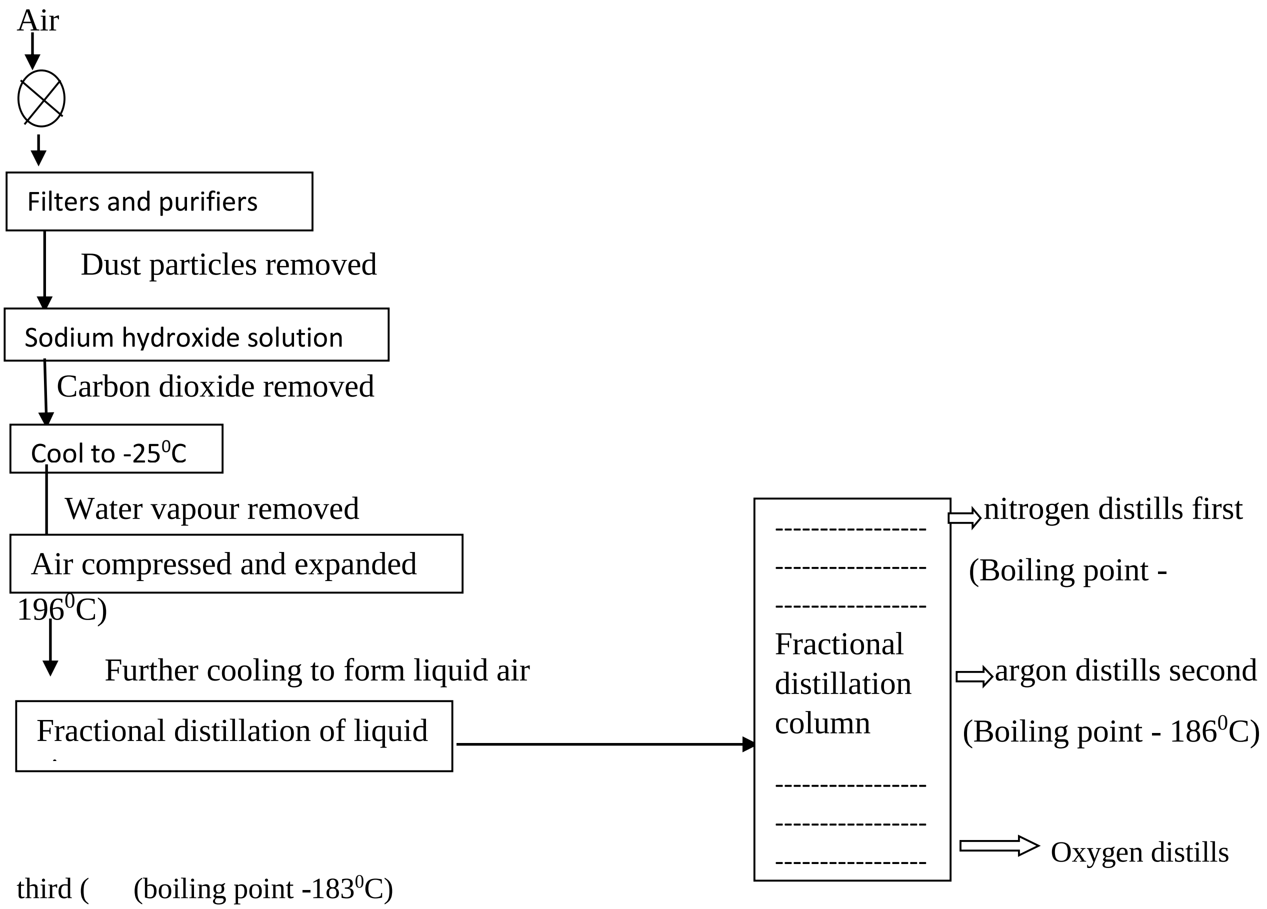


Fig. 5.1: Fractional distillation of liquid air

Uses of gases found in air

Nitrogen

Nitrogen is important in many ways: it dilutes the oxygen in air and therefore weakens its power, otherwise processes of burning and rusting in air would have been very rapid and uncontrollable. Leguminous plants for example beans, peas among others have bacteria in their roots which convert nitrogen in the air into compounds which make them richer in nitrogen compounds.

Oxygen

As we have seen, oxygen is the active fraction of air. It is necessary for breathing, burning, rusting and decay. It must be stressed that without it, life would be impossible.

Carbon dioxide

Ordinary air contains 0.03% of carbon dioxide whereas exhaled air contains about 3%.the burning of wood, coal, petrol, oils and other carbon compounds ass carbon dioxide to the atmosphere. Green plants use it to make their food. We also use it for preserving soft drinks and beer.

Noble gases

These are argon, neon, helium, krypton and xenon. They do not react with any other substances under ordinary conditions.

Argon is used in electric bulbs, helium is used in weather balloons and neon in coloured investments signs.

Self-assessment exercise 5.1

1. Which two major gases are found in air?
2. Why is it possible to separate the components of air using physical methods? How do we call this method?

5.3 Quantitative determination of the fraction of oxygen in air

Experiment 5.1

Aim

To determine the percentage of air used when a candle burns.

Apparatus and chemicals

- trough
- candle
- 30 cm ruler
- Beehive shelf
- Gas jar
- Sodium hydroxide solution

Procedure

1. Invert an empty gas jar over the candle before lighting it
2. Measure the height A in cm as shown in Figure 5.2 a.
3. Remove the gas jar

4. Light the candle and immediately cover it with the empty gas jar
5. When the candle flame goes off, measure height B in cm.

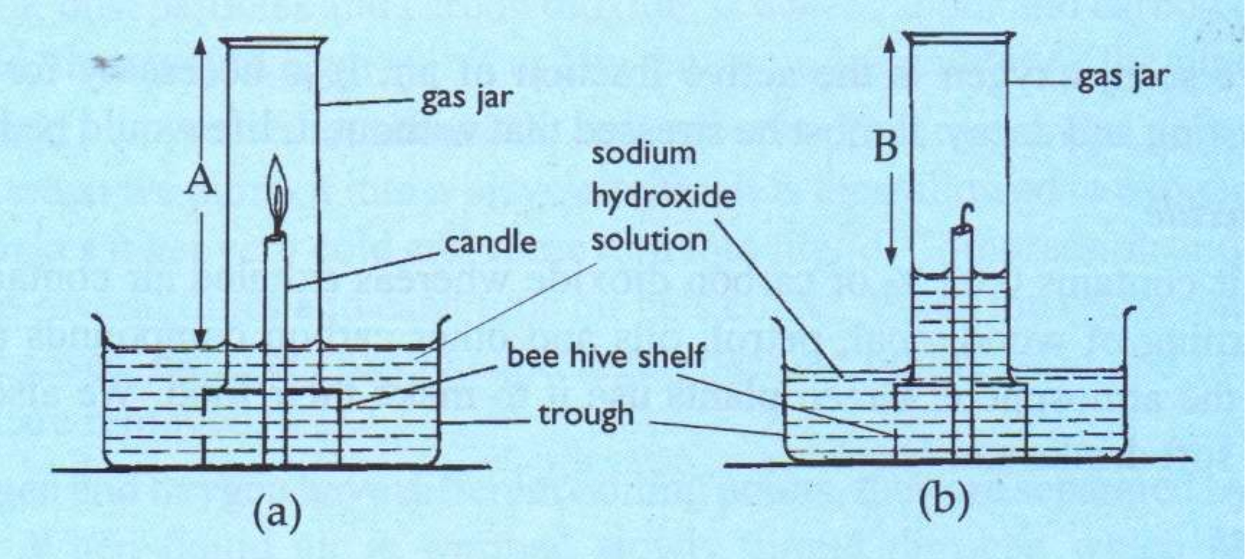


Fig. 5.2: Determining the percentage of air used when a candle burns

When the candle is lit, it burns for some time then it goes off, the water level rises inside the gas jar and goes down in the trough

Amount of air in the gas jar = A

Amount of air used is given by subtracting height B from A = A – B.

Therefore, the percentage of oxygen is given by:

$$\frac{(\text{Height A- height B}) \times 100}{\text{Height A}} = \frac{(A-B) \times 100}{A} = C\ \%$$

Substituting values of A and B in the equation give the value of C as 21%. Hence about 21% of air is used during the burning of the candle. This is the approximate percentage of oxygen in air.

5.4 Laboratory preparation of oxygen

Experiment 5.2

Aim

To prepare and determine some properties of oxygen.

Apparatus and chemicals

- Flat –bottomed flask
- Trough
- Gas jars/boiling tubes
- Wooden splints
- Manganese dioxide
- Sulphur
- Sodium
- Iron
- Thistle funnel or dropping funnel
- Bee hive shelf
- Deflagrating spoon
- 20 volume hydrogen peroxide
- Carbon (wood charcoal)
- Phosphorus
- Magnesium
- Copper turnings

Procedure

1. Place a spatula full of manganese dioxide into the flask
2. Arrange the apparatus as shown in Fig. 5.3. Note that the thistle funnel is dipped into the flask contents. If a dropping funnel is used, dipping is not necessary
3. Add some hydrogen peroxide to the manganese dioxide. What happens when hydrogen peroxide comes into contact with manganese dioxide?
4. Collect 7-8 jars of oxygen. Cover the jars while still under water. You may discard the first gas collected as this is largely air that was in the apparatus

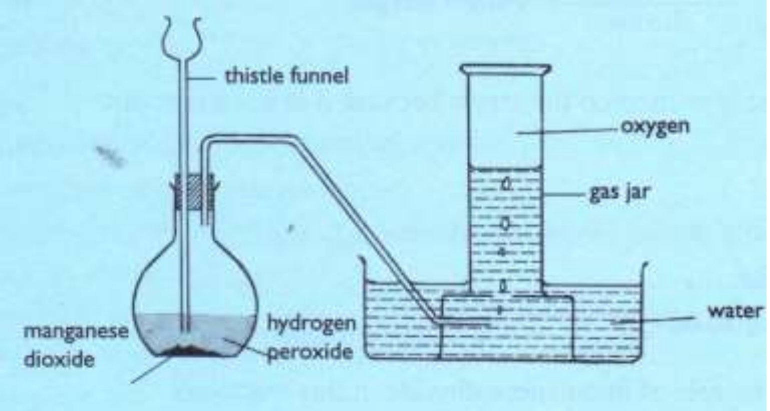


Fig. 5.3: Set-up for laboratory preparation of oxygen

- What is the colour of the gas collected?
- Smell the gas by wafting some to your nose.
- What does it smell like?
- What does the method of collection of the gas tell you about solubility of oxygen?

5. Test the gas in one jar with a glowing splint (see Fig. 5.4)

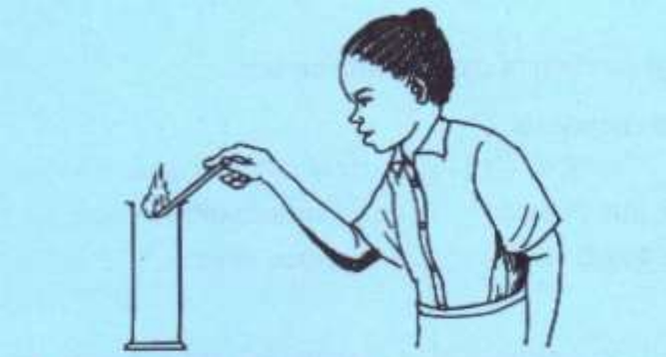


Fig. 5.4: Testing the presence of gas with a glowing splint

- What happens to the glowing splint?

Oxygen is prepared in the laboratory by decomposition of hydrogen peroxide which is a colourless liquid. Hydrogen peroxide decomposes to form water and oxygen at room temperature. The decomposition reaction is represented in the equation below.

$$\text{Hydrogen peroxide} \xrightarrow{\qquad\qquad} \text{water + oxygen}$$

This reaction is very slow, but the manganese dioxide speeds it up. It acts as a catalyst. A catalyst is a substance which speeds up the rate of a chemical reaction but remains unchanged at the end of the reaction.

$$\text{Hydrogen Peroxide} \xrightarrow[\text{dioxide}]{\text{manganese}} \text{+ water + oxygen}$$

Note: A catalyst is written out on the arrow because it is not a reactant.

Self-assessment exercise 5.2

Hydrogen peroxide reacts with manganese dioxide to produce water and oxygen

a. What is the role of manganese dioxide in this reaction?
b. Write the chemical equation for the above reaction.
c. What is the test for the gas produced in the equation above?

5.5 Physical and chemical properties of oxygen

Reaction of oxygen with non- metals

Experiment 5.3a

Aim

To determine the reaction of oxygen with carbon

Apparatus and chemicals

- charcoal
- red and blue litmus paper
- deflagrating spoon
- gas jar of oxygen
- distilled water
- source of heat

Procedure

1. Place a small amount of powder of charcoal in a deflagrating spoon. Heat it until red hot, then quickly plunge the spoon into a gas jar of oxygen as in Fig. 5.5. Do not uncover the gas jar until the burning is complete. Test the product with moist red and blue litmus papers.
2. Remove the cover from the gas jar, and quickly add a little water, shake the jar and test the resulting solution with litmus paper or litmus solution.
3. Copy Table 5.2 in your book and record your observations.

Experiment 5.3 b

Aim

To determine the reaction of oxygen with sulphur

Apparatus and chemicals

- Sulphur
- Deflagrating spoon
- Source of heat
- Gas jar of oxygen

Procedure

1. Place some sulphur in a deflagrating spoon
2. Heat until the sulphur starts burning
3. Note the colour of the flame and quickly plunge the spoon into a gas jar of oxygen, as shown in Fig. 5.5
4. Record your observation in table 5.2 copied in your notebook.

Experiment 5.3 c

Aim

To determine the reaction of oxygen with phosphorous

Apparatus and chemicals

- Phosphorus
- Gas jar of oxygen
- Source of heat
- Deflagrating spoon

Procedure

1. Place a piece of white phosphorous in a deflagrating spoon
2. Heat until the phosphorus starts burning
3. Note the colour of the flame and quickly plunge the spoon into a gas jar of oxygen
4. Record your observation in table 5.2 drawn in your notebook

Caution: Take care when using phosphorus; it burns spontaneously

Table 5.2: Combustion of non-metals in oxygen

Element (non-metal)	How does it burn? (colour of flame)	Colour changes of litmus papers, if any	Is the solution acidic or alkaline?

Carbon Sulphur Phosphorus			

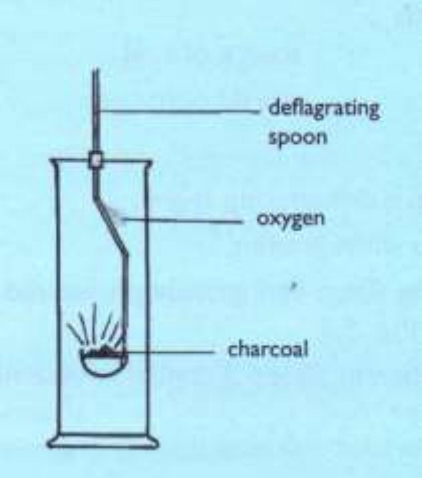


Fig. 5.5: Burning carbon in oxygen

Sulphur burns in oxygen with a blue flame producing sulphur dioxide.

Sulphur + oxygen → sulphur dioxide

Sulphur dioxide is an acidic gas. It turns wet blue litmus paper red. It dissolves in water to form sulphorous acid, which also turns blue litmus paper red.

Sulphur dioxide + water → sulphurous acid

Carbon (wood charcoal) burns brightly with a red glow to form carbon dioxide which dissolves in water to produce carbonic acid. The acid turns blue litmus red.

Carbon + oxygen → carbon dioxide

Carbon dioxide + water → carbonic acid

Phosphorus burns more brightly to produce white fumes of two oxides namely phosphorous (III) oxide and phosphorous (V) oxide. These oxides are also acidic

Phosphorous + oxygen → phosphorous trioxide

Phosphorous + oxygen (excess) → phosphorous pentoxide

When shaken with water, the oxides of phosphorous dissolve o form two acids:

Phosphoric (III) and phosphoric (V) acids.

Phosphorous trioxide + water → phosphoric (III) acid

Phosphorous pentoxide + water $\xrightarrow{\qquad\qquad}$ phosphoric (V) aid

Reaction of oxygen with metals

Experiment 5.4a

Aim

To determine the reaction of oxygen with magnesium.

Apparatus and chemicals

- Magnesium ribbon
- Bunsen burner
- Blue and red litmus papers
- A pair of tongs
- Gas jar of oxygen

Procedure

1. Hold a piece of magnesium with a pair of tongs
2. Heat in a Bunsen burner flame till it ignites and quickly transfer it into a gas jar of oxygen. What do you observe?

Caution: avoid looking directly at the flame.

3. Add a little water into the gas jar and shake it. Dip a blue and red litmus paper into the solution. Observe what happens to the litmus papers.

Experiment 5.4b

Aim

To determine the reaction of oxygen with sodium

Apparatus and chemicals

- Sodium
- Bunsen burner
- Water
- Deflagrating spoon
- Gas jar of oxygen
- Blue and red litmus paper

Procedure

1. Place the piece of sodium in a deflagrating spoon. Heat over the Bunsen burner flame until it ignites, while it is still burning, lower it into the gas jar full of oxygen, see Fig. 5.6. note the colour change of the flame
2. When sodium has stopped burning, add a little water to the jar and shake well. Add red and blue litmus papers into the solution and observe any colour changes.

Experiment 5.4c

Aim

To determine the reaction of oxygen with iron

Apparatus and chemicals

- Iron wool
- Bunsen burner
- Water
- Deflagrating spoon
- Gas jar of oxygen
- Blue and red litmus papers

Procedure

1. Wrap some iron wool on the deflagrating spoon
2. Heat the iron wool over the Bunsen burner flame until it is red hot. Put into the jar of oxygen
3. When the reaction is complete, add some water. Shake and add red and blue litmus papers into the gas jar. Observe nay colour changes.

Experiment 5.4d

Aim

To determine the reaction of oxygen with copper

Apparatus and chemicals

- Copper turnings
- Gas jar of oxygen
- Deflagrating spoon
- Blue and red litmus papers

Procedure

1. Put copper turnings in a deflagrating spoon
2. Heat strongly and then dip the spoon into a gas jar of oxygen.
3. Test the products formed when water is added to the jar with red and blue litmus papers as described earlier

Experiment 5.4e

Aim

To determine the reaction of oxygen with calcium

Apparatus and chemicals

- Calcium pieces
- Gas jar of oxygen
- Deflagrating spoon
- Blue and red litmus papers

Procedure

1. Put calcium pieces in a deflagrating spoon
2. Heat strongly until it burns and dip into the gas jar of oxygen
3. After the reaction is complete, add some water and shake
4. Test the products formed with red and blue litmus papers

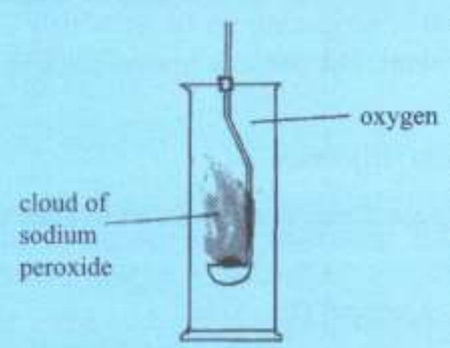


Fig 5.6: Burning of sodium

Element	How does it burn? (colour of flame)	Colour changes of litmus papers, if any	Is the product soluble in water?	Is the solution acidic or alkaline?
Carbon Sulphur Phosphorus				

Discussion

From the results, we can arrange the metals according to how reactive they are with oxygen, staring with the most reactive.

Magnesium burns with a very bright flame in oxygen to produce a white powder of magnesium oxide.

Magnesium + oxygen ⟶ magnesium oxide

Magnesium oxide dissolves in water to form magnesium hydroxide.

Magnesium oxide + water ⟶ magnesium hydroxide

Sodium burns with a bright yellow flame to produce sodium peroxide.

Sodium + oxygen ⟶ sodium peroxide

This oxide dissolves readily in water to produce sodium hydroxide solution and oxygen gas. The solution turns red litmus paper blue.

Sodium peroxide + water ⟶ sodium hydroxide + oxygen

Although sodium peroxide produces an alkali when reacted with water, it is not classified as a base. We shall learn more about this in a later unit.

Iron burns with yellow sparks producing an oxide which is insoluble in water and therefore does not form an alkali. The oxide is brown-black in colour and is called iron (III) oxide.

Iron + oxygen ⟶ iron (III) oxide

Copper burns with a blue-green flame in a gas jar of oxygen to form black copper (II) oxide.

Copper + oxygen ⟶ copper (II) oxide

This oxide is not soluble in water. Therefore, it does not form an alkali

Calcium burns vigorously in oxygen producing a bright red flame. It forms white calcium oxide.

Calcium + oxygen ⟶ calcium oxide

Calcium oxide dissolves in water to form calcium hydroxide solution (lime water)

Calcium oxide + water ⟶ calcium hydroxide

Experiment 5.5a

Aim

To determine the action of oxygen on litmus

Apparatus and chemicals

- Blue and red litmus papers
- Gas jar of oxygen

Procedure

Insert wet blue and red litmus papers into a gas jar of oxygen. What do you observe? (Fig. 5.7)

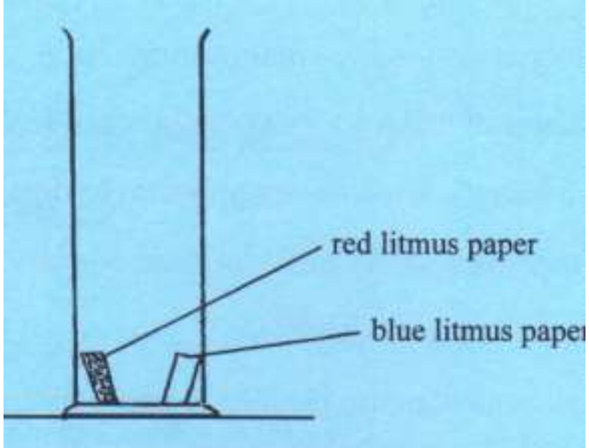


Fig. 5.7 Litmus papers in oxygen

Experiment 5.5b

Aim

To investigate the solubility of oxygen

Apparatus and chemicals

- Gas jar of oxygen
- Distilled water
- Trough

Procedure

1. Invert a gas jar full of oxygen in a trough with distilled water and note the level of water inside the gas jar
2. Leave for some time and note any change in the water level. (Fig. 5.8)

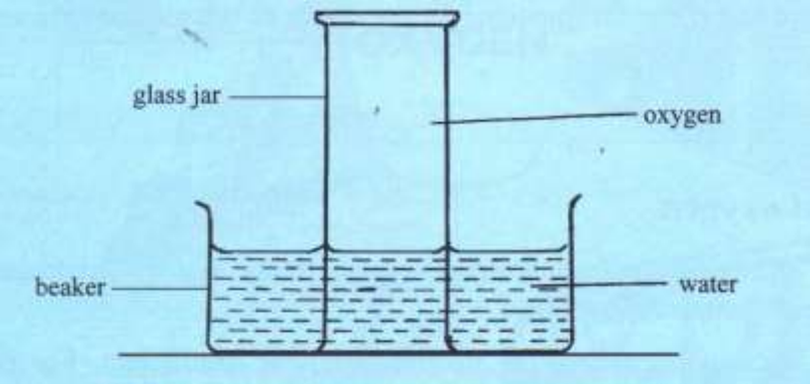


Fig. 5.8 Set-up for testing the solubility of oxygen in distilled water

- What changes do you observe? Explain your observation
- From the results of the experiments described above, what are the properties of oxygen?

From the experimental results, the following are physical properties of oxygen:

- Oxygen is a colourless gas and has no smell
- It is slightly soluble in water
- It is slightly denser than air
- Oxygen is a neutral gas

Chemically, oxygen is very reactive. It supports burning better than air. It allows elements to burn in it more brightly than in air.

When elements burn in oxygen, they combine with it to produce oxides. Therefore, an oxide is a compound of oxygen and another element. The reactions are called oxidation reactions. Elements are oxidized when they react with oxygen to form oxides.

Non- metallic elements form oxides which in solution from turn moist blue litmus paper red. Such oxides are said to be acidic. Solutions of these oxides in water are therefore acidic. However there are some oxides which are neutral. This means they have no effect on litmus. Examples are carbon monoxide and hydrogen oxide (water).

Metals burn in oxygen to give oxides called bases or basic oxides. Some metal oxides. For example iron monoxide, are insoluble in water. Solutions of some basic oxides such as sodium oxide are known as alkalis.

Self-assessment exercise 5.3

1. State two physical properties of oxygen.

2. Write word and chemical equations for the reactions between oxygen and
   a. Calcium
   b. Sodium

Uses of oxygen

1. In welding and cutting of metals (oxy-acetylene flame and oxy-hydrogen flame offer very high temperatures) Fig 5.9d
2. As an aid to breathing where the natural supply is insufficient. For example high-attitude climbing, driving and also in hospitals (Fig 5.9b, c f)
3. Oxygen is used to remove impurities in the process of making steel (Fig 5.9a)
4. Liquid oxygen is used to burn the fuel in some space rockets. For example, oxygen-hydrogen liquid mixture is used as a fuel in rocket ships and other space craft. (Fig 5.9e.
5. A mixture of charcoal, petrol and liquid oxygen is used as an explosive.

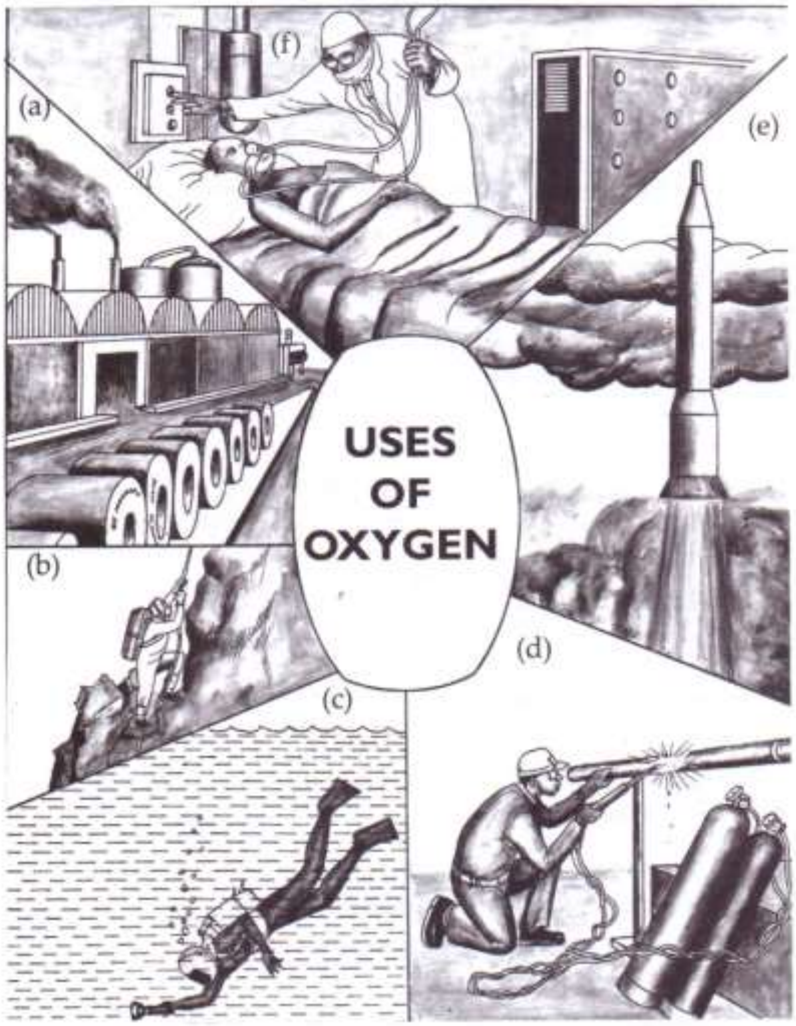


Fig 5.9 Some uses of oxygen

5.7 Atmospheric pollutants and their sources

Atmospheric pollutants refers to the introduction of substances to the environment which have harmful effects on the ability of the environment to support life. It may also mean making the environment ‘dirty’. The air over and near large centres or industrial areas is very different from pure air as it contains a high number of undesirable substances. Such air is said to be polluted.

When man and other living things breathe, they take in air with the pollutants that are mixed in it. Pollution therefore endangers health and well-being. For example, it causes respiratory diseases and can even cause cancer.

Atmospheric pollutants can be divided into two categories: poisonous and solid particles.

1. Poisonous gases and their effects on health and the environment

These arise because of complete or incomplete combustion of fuels such as firewood and petrol among others. The poisonous gases include:

a. Sulphur dioxide (SO2)

Most fuels contain sulphur compounds. Whenever such raw fuels are burned, the sulphur forms sulphur dioxide and hydrogen sulphide. These gases are serious pollutants. They are toxic and cause respiratory diseases. They also form ‘acid’ rain which attacks metals.

Sulphur + oxygen → sulphur dioxide

b. Nitrogen dioxide (NO2)

This can be formed from nitrogen in the air combining with oxygen during lightning or in internal combustion of engines where the temperatures are very high. This gas is toxic and promotes respiratory dis4eases. It also forms ‘acid’ rain which attacks metals.

Nitrogen + oxygen → nitrogen dioxide

c. Carbon dioxide *CO2)

It prevents layers of warm air from escaping into the upper atmosphere causing greenhouse effect that leads to global warming. On a global scale, this brings about melting of the polar ice caps thus raising the level of oceans, hence flooding of lowland areas.

d. Carbon monoxide (CO)

If a fuel, such as charcoal is burned in an insufficient supply of air, the carbon in the fuel is not completely burnt to carbon dioxide. Instead, carbon monoxide is formed.

Carbon + oxygen → carbon monoxide

Carbon monoxide is poisonous because it combines with haemoglobin in the red blood cells to form a stable compound called carboxyhaemoglobin. Carboxyhaemoglobin does not give up oxygen to body cells like oxyhaemoglobin (combination of haemoglobin and oxygen). This causes suffocation and may lead to death. Carbon monoxide is dangerous because it has no odour or colour and thus cannot be easily detected.

e. Unburnt hydrocarbons (C_xH_y), such as petrol fumes, can be carcinogenic.

f. Chlorine

This gas is formed when insecticides and perfume sprays are decomposed in the atmosphere. Chlorine destroys ozone gas which forms the ozone layer in the upper atmosphere.

g. Ozone

Ozone (O30 is a form of oxygen. It is generally a light blue gas. It is known to cause headaches among people and regarded poisonous to human. However, its presence at a height of 12-50km (upper atmosphere) acts as a barrier to direct ultra-violet radiation from the sun. This ultra-violet radiation is quite harmful to living things. Below 12km (lower atmosphere), the presence of ozone is harmful to crops, clothes, human lungs, tyres of bicycles and cars.

It is produced naturally when radiation from the sun strikes oxygen in the earth's upper atmosphere, converting some of it into ozone. It is also produced during lightning storms. Presence of electricity might convert oxygen into ozone.

Recent testing of the ozone layer has demonstrated that waste gases from industries and from aerosol sprays are damaging it. This layer is of vital importance in protecting life on earth from too much exposure to the sun's ultraviolet radiation, which can cause skin cancer. About 95 to 99 percent of this radiation is blocked by the ozone layer.

Due to the high degree of destruction, efforts are being directed towards:

(i). Maintaining the ozone layer which prevents ultra-violet radiation reaching the earth's surface.

(ii). Preventing formation of ozone within the lower atmosphere.

2. Solid particles

Large quantities of solid particles suspended in the air are also pollutants. Their source could be cement, flour, sugar industries, mining, quarrying and construction activities.

Examples of solid particle pollutants are carbon and dust. Carbon particles are added to the air from inefficient combustion of fuels. Dust particles are also found in the air. They are fine, dry powder consisting of tiny particles of waste matter or dust.

Smoke from tobacco or any other source also consists of carbon and other solid particles. Petrol contains a compound called tetra ethyl lead that is added to it to make it burn better. Compounds of lead are produced when petrol burns in vehicle engines. These compounds are dangerous air pollutants because they accumulate in the body and damage the brain. Thus lead compound is not added any more in petrol.

All these solid particles are very small, they can enter the lungs and be retained there. They increase the risk of respiratory problems such as lung cancer and other diseases.

Sources of pollutants

a. Smoke from industries
b. Smoke form vehicles
c. Charcoal burning in a blaze
d. Flooding
e. Smoke from burning tyres

Figure 5.10 a, b and c below shows how industries, automobiles and burning substances add gases like carbon dioxide and solid particles, in the atmosphere

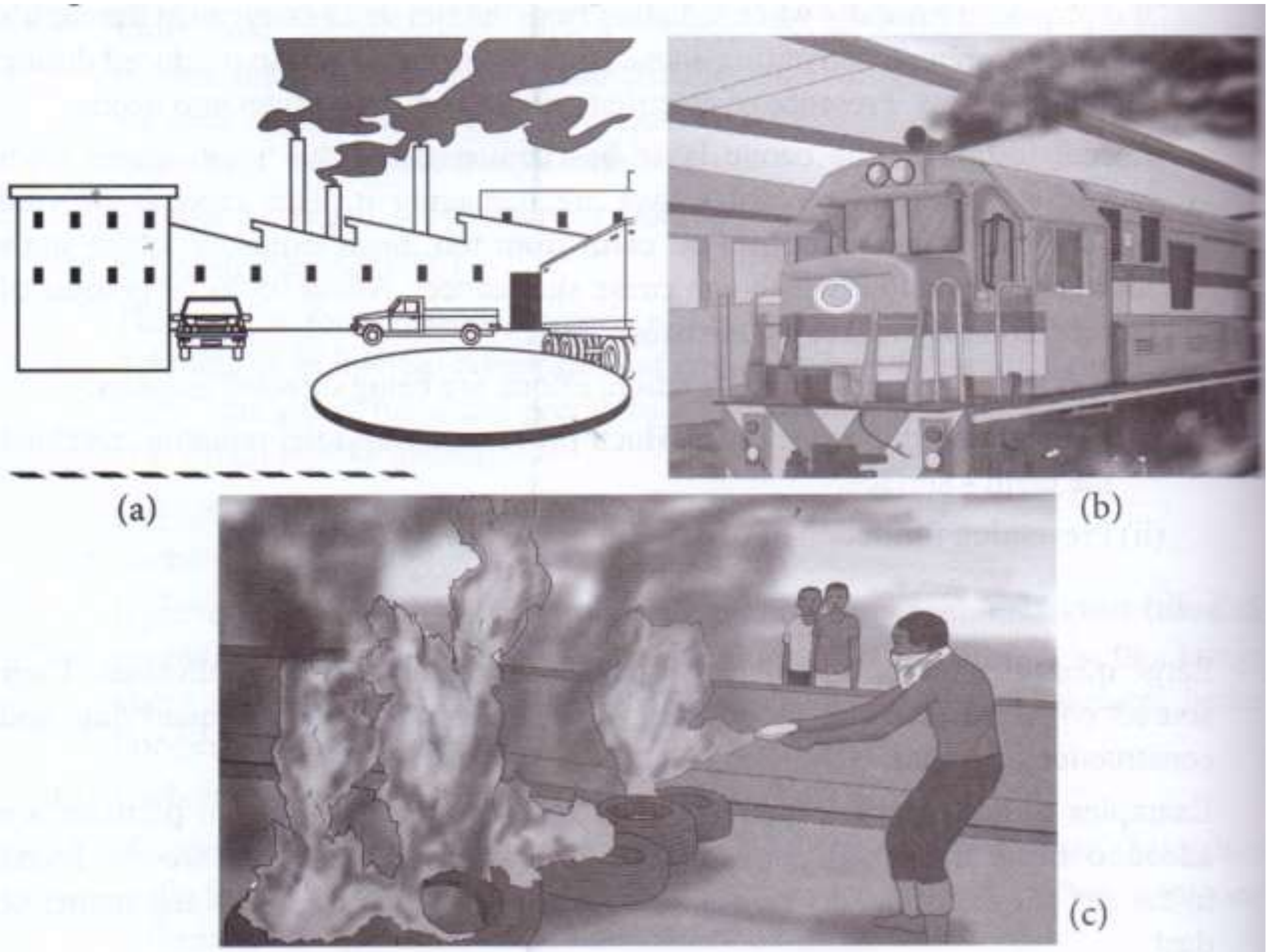


Fig 5.10: Forms of pollution

Carbon pollutes the air. Some gases like carbon dioxide among other gases cause global warming which leads to flooding in some areas.

Fig. 5.11: Flooding - an effect of global warming

Efforts being made to reduce atmospheric pollution

The atmosphere is vital to us, we need its oxygen from respiration to help supply most of our energy requirements by the combustion of fuels, and for many of the chemical reactions which are important in industries. The atmosphere is one of the most important natural resources therefore, it should be preserved in a clean and acceptable state at all costs.

Here are some of the steps that should be taken:

- Improved combustion of fuel in petrol and diesel engines
- Introduction of better processing of fuels to make them free from sulphur compounds
- Introduction of better and more efficient filter systems in industries
- Introduction of smokeless solid fuels or fuels like hydrogen.

Self-assessment exercise 5.4

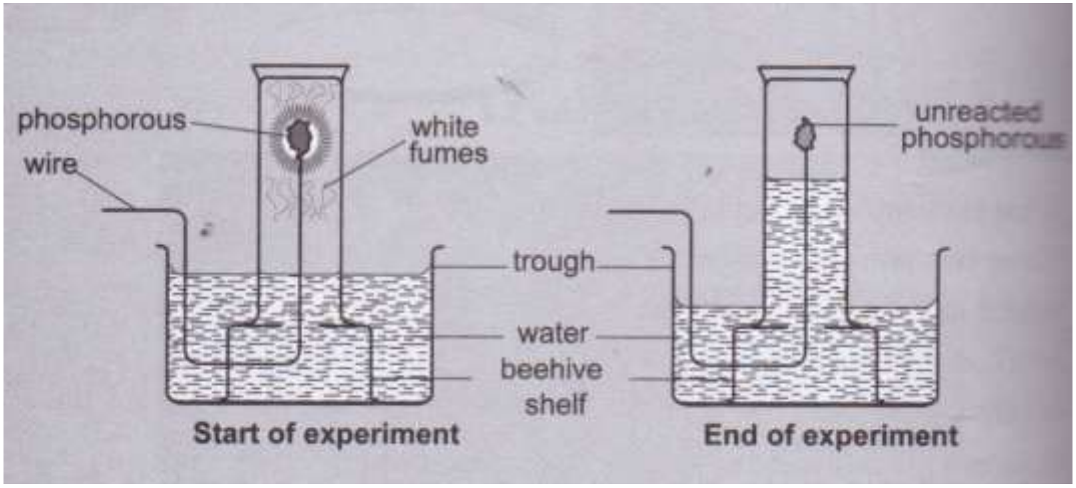
1. What is chemical pollution?
2. Name two poisonous gases.
3. Which gas causes:
   a. Global warming?
   b. Ozone damage?
4. How can global warming and ozone damage be prevented?

Summary

- Air is a mixture of gases.
- The most predominant component of air is nitrogen (78%) which is followed by oxygen (21%)
- The active part of air, that is, the one which supports combustion (burning) is oxygen
- Nitrogen, oxygen and noble gases are manufactured by fractional distillation of liquid air.
- Non-metals combine with oxygen during burning to form acidic oxides
- Metals combine with oxygen during burning to form basic oxides
- Sodium reacts with air in low temperatures to form sodium oxide (Na_2O). During burning, sodium combines with oxygen in the air or pure oxygen to form sodium peroxide (Na_2O_2).
- Oxygen is prepared in the laboratory using hydrogen peroxide. Manganese dioxide is used as a catalyst.
- The poisonous gases in the air include sulphur dioxide, nitrogen dioxide, carbon monoxide and chlorine among others.
- Ozone layer protects life on earth from too much exposure to the sun’s ultraviolet radiation
- Carbon dioxide in the upper atmosphere causes global warming.

Revision exercise 5

1. A student set up the experiment below to determine the percentage of oxygen in air.

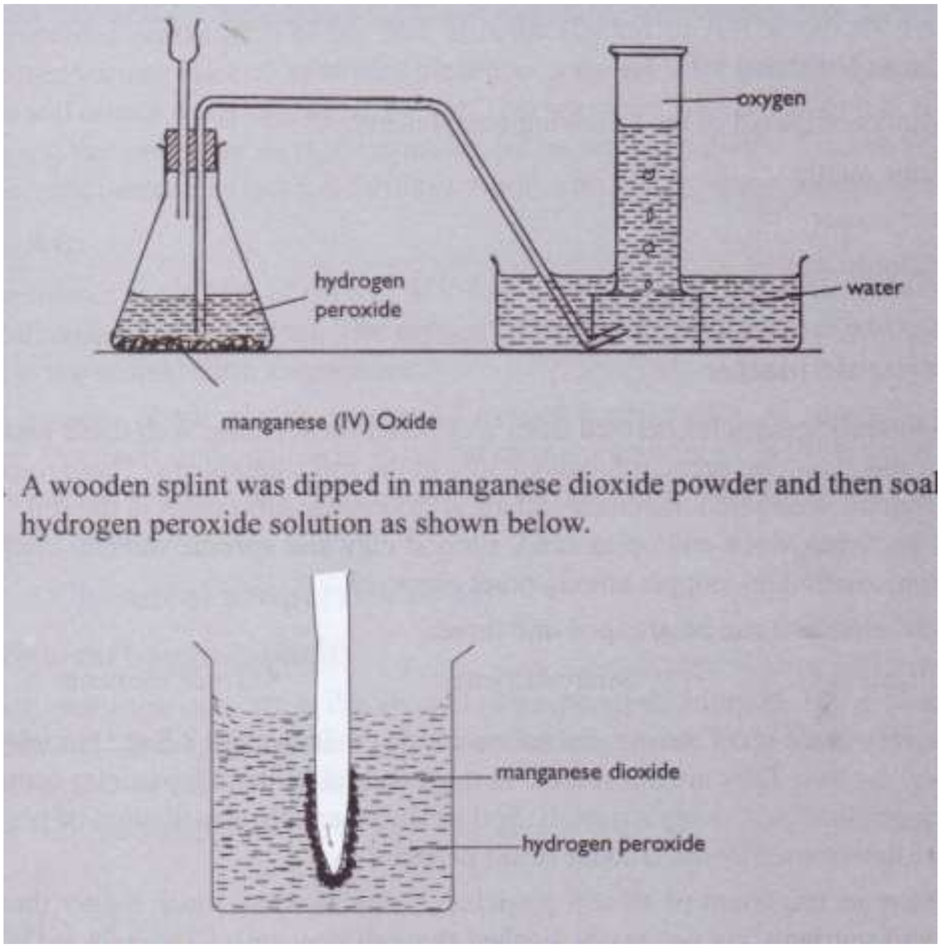


   a. What were the white fumes formed in the gas jar at the start of the experiment?
   b. Where did the white fumes go at the end of the experiment?
   c. The volume of the air inside the gas jar at the start of the experiment was 100 cm3. At the end of the experiment, only 80 cm3 was remaining. Calculate the percentage of oxygen in the air from these results.
   d. Write down a word equation from the reaction between phosphorous and oxygen.
   e. At the end of the experiment, red and blue litmus papers were dipped into the solution in the trough. State and explain what was observed
2. Name two pollutants produced when petrol is burnt in car engines.
3. Manganese (IV) oxide acts as a ____ in the decomposition of hydrogen peroxide to produce ____.
4. Oxygen is usually collected over ______.
5. Many metals burn in oxygen (or in air) to give ____ oxides. These oxides react with an acid to give ____ and _____ only.
6. The product of burning magnesium weighs _____ than the magnesium. When candle burns, there is an apparent ______ in mass.
7. Air contains approximately 21% of ______ and about 0.03% of ____.
8. Oxygen is prepared in the laboratory using hydrogen peroxide solution and manganese dioxide

Hydrogen peroxide → water + oxygen

Why is the manganese dioxide not present in the above word equation?

9. The diagram below shows preparation of oxygen gas. Identify two mistakes with the set up.



A wooden splint was dipped in manganese dioxide powder and then soal hydrogen peroxide solution as shown below.

a. Effervescence is observed because of the production of ____ gas.
b. Write the word equation for the reaction.

10. Mention two steps which should be taken to maintain the atmosphere in a clean and acceptable state.

UNIT 6: SOIL

6.1 Composition of soil

Soil is mainly composed of the following components:

- Inorganic matter
- Organic matter
- Water (moisture)
- Air

Inorganic matter

These include rock particles derived from weathered rock. Along with these particles there are minerals which account for about 50% of the soil components. These minerals derived from the weathered materials include all inorganic substances in the soil. They form soil particles which end up as sand, silt and clay and specific mineral elements such as iron, aluminum, copper among other elements.

The mineral elements can be grouped into three

- Silicates
- Micronutrients
- Trace elements

Sand is mainly made up of the mineral called quartz or silica (SiO_2). Sand particles are chemically inactive. They are not soluble in most natural acids. Silt particles contain a larger proportion of secondary minerals. Soil structure and the distribution of pores in the soil are determined by the amount of silt particles in it.

Clay is the most important of all soil particles. It retains water much longer than silt or sand. Soil nutrients are not easily leached through clay soil. Clay soils are heavy, plastic and sticky when wet but they harden and crack on drying.

c. Organic matter

It constitutes about 5% of the soil. It consists of animal waste and humus. Humus is derived from decaying parts of plants, animal waste and remain. Bacteria are involved in the decomposition of these materials to form humus. The amount of humus in the soil depends on the amount of vegetable matter that has decomposed.

Humus plays a role in determining the acidity or alkalinity of the soil. The higher the amount of humus in the soil, the higher the number of bacteria. These bacteria help in the decomposition of organic matter into simple substance which help in increasing soil fertility. The higher the amount of humus the higher the soil pH.

Increase in humus content →

Low pH increase in soil pH high pH

Increase in soil fertility

d. Water (Moisture)

It constitutes nearly 25% of the soil. I is usually called soil water. As a solvent, it dissolves various mineral salts which are necessary for plant growth. Too much water in the soil causes water logging in the soil. This interferes with circulation of air, thereby reducing the amount of oxygen necessary for bacteria and plants. Less bacteria means less humus formation (less soil fertility). A soil with less oxygen is said to be anaerobic.

e. Air

It constitutes about 15% of the soil. It is found in soil pores. It is sometimes referred to as soil atmosphere or soil air. The oxygen content in the soil is lower and more varied than in the normal earth atmosphere.

This gaseous phase of soil consist of oxygen, carbon dioxide, nitrogen and noble gases. The gases in the soil exist in the following proportions: 78.6% nitrogen, 0.2% to 0.6% carbon dioxide, and 20.6% to 21.0% oxygen. The other percentage is occupied by noble gases.

6.2 Chemical properties of soil

Cation exchange capacity

Cation exchange capacity is the amount of exchangeable ions in 100 g of soil. These cations include3 calcium, potassium, sodium, among others. Fertile soils have very high cation exchange capacity (C.E.C.)

Soils with very high amount of exchangeable sodium are toxic to most crops except rice.

Salinity of soils

The amount of slats in the soils per unit mass is known as salinity. Saline soils have excess salts especially sodium chloride which can be toxic to most crops.

Soil pH

This is the most important chemical property of the soil. It determines the soil fertility in terms of mineral nutrients.

Soil pH is the acidity or alkalinity of a soil solution. The degree of acidity depends on the concentration of hydrogen ions in the soil solution. A high concentration of hydrogen ions makes the soil acidic. There are alkalis in the soil solution which neutralize the acids in the soil.

Self-assessment exercise 6.1

1.

a. Name four main components of soil.

b. Which components of air determine the acidity or alkalinity of air?

2. Define the following terms:

a. Action exchange capacity

b. Soil salinity

Experiment 6.0

Aim

To determine soil pH in the laboratory

Apparatus and chemicals

- g of soil
- Barium sulphate powder
- universal indicator
- test tube
- distilled water
- pH chart

Procedure

1. put 2 g of soil in a test tube
2. Add barium sulphate powder and 5cm3 of water. Shake the mixture for about 2 minutes. Barium sulphate helps the solution to clear. Let the mixture settle until there is a layer of clear water at the top.
3. Add about 1 cm^3 of universal indicator
4. Compare the colour of solution with a corresponding colour on the chart of the universal indicator bottle. Read off the pH.

Soil pH is a chemical characteristic of soil. It is used to determine soil acidity or alkalinity. All soils of pH more than 7 are alkaline. A soil pH equal to 7 is neutral. Soil pH of about 4.5 or less is strongly acidic.

The acidity of the soil can be lowered (increase of pH) by application of lime ($Ca (OH)_2$) or basic fertilizer.

In raising acidity or lowering soil pH, application of sulphur or acidic fertilizer such as sulphate of ammonium (NH4)2 SO4) is recommended.

6.3 Soil pollution

This can be defined as releases of substances in the soil by human activities in such quantities whose effects are either harmful or unpleasant to human, animals or plants.

A pollutant is a substance mainly waste that contaminates air, water or soil.

Sources of soil pollutants and their effects

Human beings deliberately or accidentally discharge chemicals into the soil which accumulate to undesirable levels to cause harm to soil organisms. These include the following:

Oxide of sulphur

These include sulphur dioxide (SO_2) and sulphur trioxide (SO_3). They enter the soil as 'acid' 'rain'. "Acid" rain effects the pH of the soil to the detriment of some plants and animals. It also causes leaching of mineral nutrients leading to decrease of soil fertility

Aerosols

They are used to control pests and diseases in plants and animals. They usually contain heavy metals, for example copper and mercury, which when present in the soil, plants takes them. When animals eat these plants their toxicity increases leading to their death.

Chemicals like aerosols kill some microorganisms. These microorganisms are very important in the formation of humus in the soil therefore soil fertility decreases.

Petroleum products

Oil spills, for example, make soil micro- organisms get insufficient oxygen from the soil.

Agro-chemicals

These include chemical fertilizer, insecticides, pesticides and fungicides

Solid waste

This comes from households and industries. Some are biodegradable while others are non-biodegradable. Solid wastes are a nuisance and may also be injurious. For example, glass bottles and scrap metals. They destroy the beauty of the environment and offer good breeding grounds for mosquitoes. Non-biodegradable solid wastes, for example rubber, plastic containers, scrap metals and glass bottles, affect soil aeration thus growth and activity of affecting microorganisms.

6.4 Prevention of soil pollution

The following steps can be taken:

- Recycling all solid wastes, for example polyethene paper, plastic container, glass bottles and scrap metals.
- Biodegradable wastes which are combustible can be burnt in the incinetors
- Organic farming should be encouraged. This will discourage excessive use of chemical fertilisers
- Biological control of pests and diseases in plants should be encouraged. This will discourage excessive use of pesticides and insecticides.
- Pipeline transportation of petrol and petroleum products should be encouraged to minimize the risk of spillage.

Self-assessment exercise 6.2

a. State two methods of preventing soil pollution
b. How does ‘acid’ rain affect the soil?

Summary

- Soil is mainly composed of the following: inorganic matter, organic matter, water (n=moisture) and air
- Humus which is part of organic matter plays a role in determining the acidity or alkalinity of the soil
- The lower the amount of humus the lower the pH of the soil
- Waterlogged soils have insufficient oxygen which reduces the amount of microorganisms and therefore less organic matter because of less humus formation.
- Increasing soil pH means lowering soil acidity. Lime ($Ca\ (OH)_2$) can be used.
- Soils can be polluted by human activities, for example addition of fertilizers to the soil
- Organic farming should be encouraged instead of excessive use of chemical fertilizer to prevent soil pollution.

Revision exercise 6

1. State two main components of soil
2. Which components of the soil determines the pH of the soil?
3. Low pH of the soil means high ………… Os the soil.
4. State
    a. One method of raising the soil pH
    b. One method of lowering soil pH
5. Mention one method of preventing soil pollution.
6. State two sources of soil pollution.

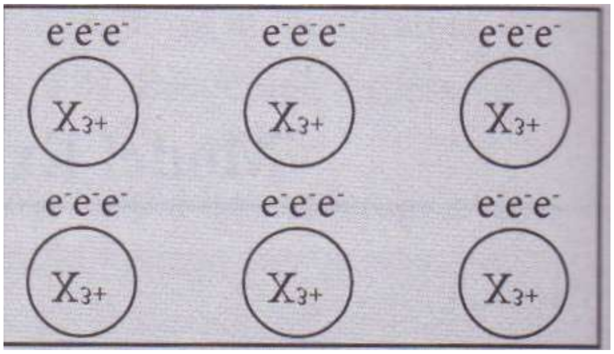
Model examination paper

Section A

1. Group I elements are
    A. Metalloid

   B. Non-metals
   C. Metals
   D. Gases
2. Which of the given elements is a metalloid?
   A. Calcium
   B. Silicon
   C. Argon
   D. Potassium
3. Electron affinity
   A. Increase down the group of the Periodic Table
   B. Increases from right to left of the Periodic Table
   C. Increases up the group of the Periodic Table
   D. Increases across the period of the Periodic Table
4. Electronegativity
   A. Is the ability to gain electrons
   B. Is the ability of an atom to attract the shared pair of electrons
   C. Is the ability of an atom to lose electrons
   D. Is the ability to form negative ions
5. Ionization energy
   A. Is the energy required to gain electrons by an atom
   B. Is the energy released when ions form
   C. Is the energy required to lose an electron from an atom in gaseous state
   D. Is the energy required to separate ions
6. Identify a metal that will react vigorously with water.
   A. Magnesium
   B. Zinc
   C. Potassium
   D. Copper
7. Group II elements
   A. Are less reactive than Group I
   B. Do not react with cold water
   C. Are not stored in normal school laboratories
   D. Form ions with two negative charge
8. Reactivity of Group VII elements increases
   A. Down the group
   B. Up the group
   C. As the number of energy levels increases
   D. As the number of protons increases
9. Group I elements
   A. Are also called alkaline-earth metals
   B. Form ions with one negative charge
   C. Form ions with one positive charge
   D. Have two electrons in the outermost energy level.

10. Electrovalent bonding involves
    A. Gain and loss of protons
    B. Sharing of a pair or pairs of electrons
    C. Gain and loss of all electrons in the second energy level
    D. Gain and loss of one or more electrons
11. Covalent bonding involves
    A. Sharing of a pair or pairs of protons
    B. Gain or loss of one or more electrons
    C. Sharing of a pair or pairs of electrons
    D. Sharing of all electrons in the outermost energy level.
12. When writing balanced chemical equations
    A. The reactants and products must be in solid form
    B. The number of atoms of the elements involved must be equal on both sides of the arrow
    C. The reactants must appear on the right hand side of the arrow
    D. The products must be written in words
13. The balanced chemical equation for the reaction between sodium and cold water is:
    A. Na (s) H2O (l) $\longrightarrow$ NaOH (aq) + H2 (g)
    B. Na (s) + H2O (l) $\longrightarrow$ na (OH) (aq) + H2 (g)
    C. 2Na (s) + 2H2o (l) $\longrightarrow$ 2NaOH (aq) + H2 (g)
    D. Na (s) + 2H2O (l) $\longrightarrow$ NaOH (aq) + H2 (g)
14. The structure below shows a types of bonding. What type of bonding is it?



    A. Covalent bonding
    B. Ionic bonding
    C. Metallic bonding
    D. Electrovalent bonding
15. Identify the types of bonding found in a carbon dioxide molecule
    A. Covalent bonding
    B. Metallic bonding

    C. Ionic bonding
    D. Electrovalent
16. Name the type of bonding present in magnesium chloride
    A. Metallic bonding
    B. Covalent bonding
    C. Ionic bonding
    D. Non-metallic bonding
17. What happens when red litmus paper is placed in an acid?
    A. Pink colour is formed
    B. There is no observable colour change
    C. A blue colour is formed
    D. Litmus paper turns colourless
18. An alkali turns phenolphthalein indicator
    A. Yellow
    B. Colourless
    C. Purple
    D. Pink
19. An acid and a base react to form
    A. Salt only
    B. Water and carbon dioxide
    C. Salt and water only
    D. Salt and hydrogen gas only
20. The reaction between an acid and a base is called
    A. Evaporation
    B. Displacement
    C. Neutral
    D. Neutralization
21. Which of the following substances is basic?
    A. Sour milk
    B. Lemon juice
    C. Pure water
    D. Sodium hydroxide
22. Identify the acid injected to the body by a bee sting?
    A. Citric acid
    B. Ethanoic acid
    C. Methanoic acid
    D. Tannic acid
23. Identify a substance added to the soil to neutralize acidity
    A. Clay
    B. Lime
    C. Sand
    D. Water
24. Which of the following pH can be that of rain water?

A. 1.0
B. 6.5
C. 9.0
D. 7.0

25. Alkanes names end with
A. –ene
B. –ane
C. –yne
D. –ol

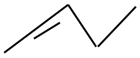
26. Alkenes contain
A. Single bonds only between the carbon atoms in the molecules
B. Single and double bonds between the carbon atoms in the molecules
C. Double between the two carbon atoms in the molecules
D. Ionic bonds between the two carbon atoms in the molecule

27. Name the following compound.

$CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_2\,CH_3$

A. Pentane
B. 2, methyl butane
C. 2-methylbutane
D. 3-methylbutane

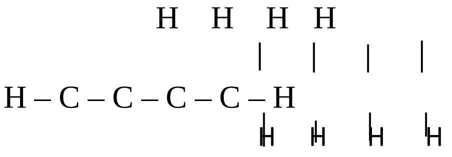
28. Name the skeletal formula



A. Butane
B. But- 1 – ene\
C. Pentene
D. Butyne

29. Which of the given homologous series is not a hydrocarbon?
A. Alkanes
B. Oxycarbons
C. Alkenes
D. Alkynes

30. Which is the condensed formula of the structure below?



$$H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-H$$

A. C_4H_9
B. $CH_3\ CH_2CH_2CH_3$
C. $CH_3\ (CH_2)_2\ CH_3$
D. C_2H_{10}

31. Which of the following compounds would undergo substitution reactions with chlorine gas in the presence of U.V light?

A. Alkenes
B. Alkane
C. Alkynes
D. Alcohols

32. Which of the following are the products of burning hydrocarbons in air?
    A. Steam
    B. Carbon monoxide an hydrogen
    C. Water and hydrogen
    D. Carbon dioxide and water
33. Which of the reactions below is an addition reaction?
    A. Reacting alkanes with halogens
    B. Reacting alkenes with halogens
    C. Burning of alkanes in air
    D. Reacting alkenes with hydrogen in absence of a catalyst
34. Which of the following statement is false about air?
    A. It is a mixture of elements only
    B. It is a mixture of elements and compounds
    C. It is a mixture of nitrogen, oxygen, carbon dioxide, water vapour and dust particles.
    D. It is a mixture of elements compounds and dust particles
35. The fraction of oxygen in air is about
    A. ¼
    B. ¾
    C. $\frac{1}{5}$
    D. $\frac{1}{3}$
36. Which one in the following list does not belong to the group?
    A. Sodium oxide
    B. Calcium oxide
    C. Sulphur dioxide
    D. Magnesium oxide

37. Burning sulphur is lowered into a gas jar full of oxygen. What visible change would you observe?

A. brilliant white flame
B. A blue flame
C. colourless fumes of sulphur dioxide
D. a few fumes of sulphur trioxide

38. Sodium burns in air with a

A. a blue flame

B. a golden yellow flame

C. greenish yellow flame

D. a lilac flame

39. Ionic bonding has a chemical bond formed
    A. between two atoms

    B. between two positively charged ions

    C. as a result of loss or gain of electrons

    D. between two negatively charged ions

40. covalent bonding is a chemical bond formed
    A. as a result of sharing electrons between atoms

    B. as a result of sharing a pair of electrons and protons

    C. as a result of loss or gain of electrons

    D. by outermost energy level electrons only

41. Bonding only involves
    A. all electrons of the atoms

    B. all the nuclei of the atoms

    C. energy levels which have no electrons

    D. only outermost energy level electrons

42. Identify a gas used in preparing sodas.
    A. Nitrogen

    B. oxygen

    C. carbon dioxide

    D. Ammonia

43. Which of the following gases used in fire extinguishers.
    A. Nitrogen

    B. chlorine

    C. Oxygen

    D. Carbon dioxide

44. Identify a gas used in filling air balloons.
    A. argon

B. oxygen

C. helium

D. neon

45. From the list below, which is not a source of alkanes?
A. Fossil fuels

B. Natural gas

C. Cracking of saturated hydrocarbons

D. reacting Sodium alkanoate with water

46. Identify a source of alkenes
    A. Fossils fuels
    B. Natural gases
    C. Dehydration of alcohols
    D. Fermentation of alcohols
47. Which of the substances below is not a common component of soil?
    A. Organic matter
    B. Water
    C. Iron ore
    D. Inorganic matter
48. The higher the organic matter in soil the
    A. Lower the humus content
    B. Higher the soil pH
    C. Higher the soil acidity
    D. Lower the soil fertility
49. The higher the action exchange capacity of the soil the
    A. Lower the soil fertility
    B. Higher the soil fertility
    C. Higher the water content of the soil
    D. Higher the production of carbon dioxide when the soil is heated
50. ‘Acid’ rain enters the soil and causes the pH of the soil to
    A. Decreases
    B. Increase
    C. Decrease and then increase
    D. Increase and then decrease

Section B

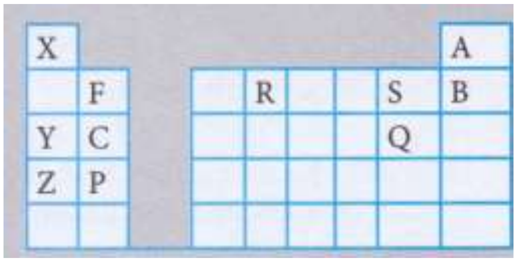
51. Three elements P, Q and R have the following electronic configurations:

P (2:2) Q (2:8:2) R (2:8:8:2)

a. Are the elements metals or non-metals? Explain. (2 marks)

b. Comment on the trend in the atomic radii of these elements. Explain. (2 marks)
c. Which element is the most reactive? Explain. (2 marks)
d. Comment on the trend of ionization energy of the elements. (1 mark)
e. Give the general name assigned to the given elements. (1 mark)
f. Chlorine consists of two isotopes 37Cl and 35Cl. If the relative atomic mass of chlorine is 35.5, determine the percentage relative abundance of each isotope. (2 marks)

52. The grid below shows part of the periodic table. (The letters do not represent the actual symbols of the elements.)



a. Which of them elements is the most reactive metal? (1mark)
b. Which of the elements is the most reactive non-metal? (1 mark)
c. What name is given to the group of elements to which C and F belong? (1 mark)
d. What name is given to the group of elements including A and B? (1 mark)
e. Element Q forms a compound with p. write the formula of the compound. (1 mark)
f. Identify the:
    a. Most electronegative element (1 mark)
    b. Most electropositive element (1 mark)
g. What elements are free in nature? Explain. (2marks)
h. What is the general name of the group in which element z belongs? (1 mark)

53.

a. What do you understand by the term ionic bonding? (1 mark)
b. How do ionic bonds form? (1 mark)
c. Use dots (.) and cross (x) diagrams to show how the following compounds form bonds
    I. sodium chloride (Na =11 Cl = 17) (2 marks)
    II. Calcium oxide (Ca = 20 O = 8) (2 marks)
d. How do covalent bonds form? (1 mark)
e. Using dots(.) and (x) diagrams, show how the following compounds form.
    I. Oxygen molecule (O=8) (1 mark)
    II. Carbon dioxide (C=6 O=8) (2 marks)

55. Use the following table to answers the questions that follow.

Solution	A	B	C	D	E	F	G	H
pH	1.0	4.0	5.0	6.5	7.0	7.5	8.0	11.0

a) Select the acidic solution

b) Select the most alkaline solution

c) Which solution is destilled water?

d) Which solution is likely to be?

    a. Lemon juice?
    b. Common salt?
    c. Rains water
    d. Soap solution
    e. Ash solution
    f. Lime water

e) Which solution will have no effect on litmus paper? (1 mark)