Comprehensive

Chemistry

for Rwanda Schools

Student's Book

Secondary 2

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PREFACE

Our aim was to write a book that should be sensitive to the needs of learners and help them build relevant skills and competences that prepare them to be well integrated in the society and exploit employment opportunities. Learners in the 21st century have to overcome the challenges brought about by the rapid developments in science, technology and society. This book provides them with learning experiences that enable them to attain knowledge and develop a global perspective and life-long learning skills, so that they can contribute to today's knowledge-based economy.

This book, ***Comprehensive* Chemistry**, has been developed on the basis of the **Ordinary Level Chemistry** syllabus. It provides skills to the learners that guide the construction of theories and laws that help to explain natural phenomena and manage people and the environment. The units covered in this book help learners to have a better understanding of the impact and influence chemistry has in a modern scientific world. The units include chemical bonding; trends in properties of elements in the periodic table; water pollution; effective ways of waste management; categories of chemical reactions; preparation of salts and identification of ions; the mole concept and gas laws; preparation and classification of oxides; electrolytes and non-electrolytes and properties of organic compounds and uses of alkanes.

Illustrations and activities given in the text help learners to obtain practical learning. Through experimentation, observation and presentation of information during the learning process, the learners develop not only deductive and inductive skills but also communication, critical thinking and problem solving skills as they try to make inferences and conclusions. Assessment at the end of each unit will help them evaluate their skills. Glossary given at the end of each unit will help the learners know the meaning of difficult words.

This book develops a sense of national identity and teaches learners to be committed to contribute to the economic development of the nation and the world. It enhances curiosity and interest in the natural and technological world as well as understanding the impact of science and technology on society. An honest attempt has been made to develop care and concern for the environment.

Any suggestions for the improvement of this book will be gratefully appreciated and acknowledged.

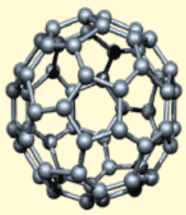
Unit 1

Chemical Bonding

LEARNING OBJECTIVES

After reading this unit, you will be able to:
- explain the nature of ionic, covalent and metallic bonding.
- state the typical physical properties of ionic compounds, and of covalent compounds.
- explain the physical properties of metals in terms of their structure.

KNOWLEDGE GAIN



In 1985, a new allotrope of carbon Buckminsterfullerene was discovered. It has a cage-like ring structure which resembles a football. It is made of twenty hexagons and twelve pentagons.

1.1 STABILITY OF ATOMS

ACTIVITY 1.1: Showing Stability of Atoms

- Take a glass full of water. Try adding water into it. Are you able to add?
- Now take another glass of water but a quarter (one-fourth) filled.
- Try adding water into it. Now, are you able to add or not?

Perform the two activities in classroom and then discuss your answers among your classmates.

In the above activity, you will observe that when the glass was already filled, there was no space to add more water into it. Thus, the water in the glass remained stable. A noble gas has a fully filled outermost shell just like the glass full of water. It has eight electrons in the outermost shells except helium (2 electrons).

When atoms or the elements combine to form molecules, a force of attraction is developed between the atoms (or ions) which holds them together. **The force which links the atoms (or ions) in a compound is called a chemical bond** (or just "bond"). A bond is formed so that each atom acquires a stable electronic configuration similar to that of a noble gas.

Table 1.1: Electronic Configurations of Noble Gases (or Inert Gases)									
Noble gas (Inert gas)	Symbol	Atomic number	Electronic configuration						Number of electrons in outermost shell
			K	L	M	N	O	P	
Helium	He	2	2						2
Neon	Ne	10	2,	8					8
Argon	Ar	18	2,	8,	8				8
Krypton	Kr	36	2,	8,	18,	8			8
Xenon	Xe	54	2,	8,	18,	18,	8		8
Radon	Rn	86	2,	8,	18,	32,	18,	8	8

The atoms combine with one another to achieve the inert gas electron arrangement and become more stable. So, when atoms combine to form compounds, they do so in such a way that each atom gets 8 electrons in its outermost shell or 2 electrons in the outermost shell.

An atom can achieve the inert gas electron arrangement in three ways:

- *By losing one or more electrons* (to another atom).

  Atoms with 1, 2 or 3 electrons in the outermost shell lose electrons to achieve stability.
- *By gaining one or more electrons* (from another atom).

  Atoms with five, six or seven electrons in the outermost shell gain three, two or one electron respectively to achieve stability.
- *By sharing one or more electrons* (with another atom).

  Atoms with four to seven electrons in outermost shell may achieve stability by sharing them with each other.

Table 1.2: Electronic Configuration of Some Metals and Non-metals						
Type of element	Element	Atomic number	Number of electrons in shells			
			K	L	M	N
Metals	Sodium (Na)	11	2	8	1	
	Magnesium (Mg)	12	2	8	2	
	Aluminium (Al)	13	2	8	3	
	Potassium (K)	19	2	8	8	1
	Calcium (Ca)	20	2	8	8	2
Non-metals	Nitrogen (N)	7	2	5		
	Oxygen (O)	8	2	6		
	Fluorine (F)	9	2	7		
	Phosphorus (P)	15	2	8	5	
	Sulphur (S)	16	2	8	6	
	Chlorine (Cl)	17	2	8	7	

EXERCISE 1.1

1. What do you mean by a chemical bond?
2. When a bond is formed, each atom acquires a stable configuration similar to _______
3. Generally, metals lose electrons to achieve inert gas electron arrangement. (True or False)
4. Which of the following is not a noble gas?
   (*a*) Helium (*b*) Neon
   (*c*) Hydrogen (*d*) Argon
5. Among, phosphorus, sulphur, and calcium; which element achieves stability by losing electron.

1.2 FORMATION OF IONS FROM ATOMS

ACTIVITY 1.2: Illustrating Formation of Ion

Divide the class into two groups. Half of the students hold positive plank cards and another half hold negative plank cards. Positive plank cards are protons and negative plank cards are electrons. Now perform the following and analyse:

Students with 5 positive and 5 negative plank cards are grouped together. Their total charge being neutral in the group.

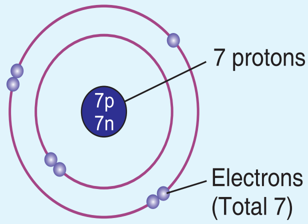
- Now, one electron is removed from the group. 4 students are left holding negative plank cards.

  Can you tell the net charge now of this group?
- Add one electron to the neutral group. 6 students are now holding negative plank cards.

  Can you now tell what is the charge of this group?
- Similarly, perform the above activity with 7 students and analyse the charge.

An atom contains electrons, protons and neutrons.

Protons carry positive charges, electrons carry negative charges and neutrons carry no charges. Every atom contains an equal number of "positively charged protons" and "negatively charged electrons". Thus, an atom is electrically neutral.



Atom

An ion is formed when an atom loses or gains one or more electrons. The atom may be of a metal or a non-metal.

A metal readily loses its outermost electron or electrons to form a positive ion or **cation**. The number of positive charges carried on a cation is equal to the number of electron(s) lost by the metal atom. Examples are given in Table 1.3.

Table 1.3: Some Metals and their Ions

Parent atom (electronic configuration)	Electrons lost	Name and symbol of ion (electronic configuration)
Sodium (2, 8, 1)	1	Sodium ion, Na^+ (2, 8)
Calcium (2, 8, 8, 2)	2	Calcium ion, Ca^{2+} (2, 8, 8)
Aluminium (2, 8, 3)	3	Aluminium ion, Al^{3+} (2, 8)

Metal ions carry positive charges because the number of positively charged protons in the nucleus becomes greater than the number of negatively charged electrons surrounding it. For example, in a sodium atom there are 11 protons in the nucleus and 11 electrons surrounding it. Loss of one electron to form a sodium ion means that there are 11 protons but only 10 electrons. There is a net charge of 1+. This charge is written as a superscript at the right of the symbol of the element (Figure 1.1).

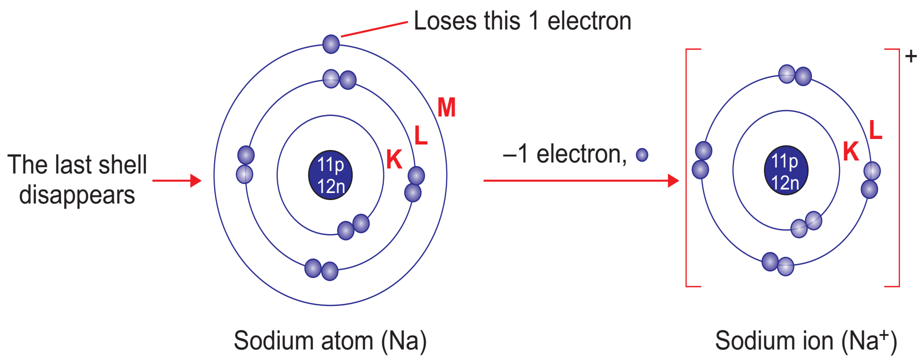


Figure 1.1: Formation of a sodium ion.

Hydrogen atoms can also lose an electron to form an ion with one positive charge.

Some non-metals readily gain one or more electrons into their outermost shell to form a negative ion or **anion**. The number of negative charges an anion carries is equal to the number of electron(s) gained by the non-metal atom. Examples are given in Table 1.4.

Table 1.4: Some Non-metals and their Ions

Parent atom (electronic configuration)	Electrons gained	Name and symbol of ion (electronic configuration)
Chlorine (2, 8, 7)	1	Chloride ion, Cl^- (2, 8, 8)
Oxygen (2, 6)	2	Oxide ion, O^{2-} (2, 8)
Nitrogen (2, 5)	3	Nitride ion, N^{3-} (2, 8)

Non-metal ions carry negative charges because the number of negatively charged electrons surrounding the nucleus becomes greater than the number of positively charged protons in it. For example, in a chlorine atom there are 17 protons in the nucleus and 17 electrons surrounding it. Gain of one electron to form a chloride ion means that there are 18 electrons and only 17 protons. There is a net charge of –1. This charge is written as a superscript at the right of the symbol of the element (Figure 1.2).

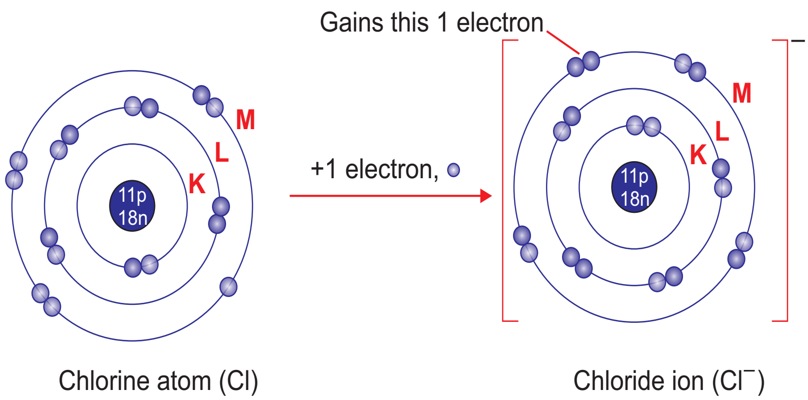


Figure 1.2: Formation of a chloride ion.

Notice that when a non-metal forms an anion, the name changes slightly; chlorine forms a chloride ion, oxygen forms an oxide ion. Several common radicals exist as negative ions including nitrate (NO_3^-), carbonate (CO_3^{2-}) and phosphate (PO_4^{3-}).

EXERCISE 1.2

1. Define ion.
2. Which of the following is an anion?
   (a) Cl^- (b) Na^+
   (c) Mg^{2+} (d) Al^{3+}
3. Why do metal ions carry positive charges?
4. The number of negative charges an anion carries is equal to the number of electrons gained by the non-metal atom. (True or False)
5. Give two examples of each:
   (i) anion (ii) cation

1.3 IONIC BONDING

The compounds which are made up of ions are known as ionic compounds. In an ionic compound, the positively charged ions (cations) and negatively charged ions (anions) are held together by the strong electrostatic forces of attraction. The forces which hold the ions together in an ionic compound are known as **ionic bonds** or **electrovalent bonds**. Since an ionic bond consists of an equal positive and negative charges, the overall charge on an ionic compound is zero. For example, sodium chloride (NaCl) is an ionic compound which is made up of equal number of positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-). Some of the common ionic compounds, their formulae and the ions present in them are given in Table 1.5.

Table 1.5: Formulae and Nomenclature of Some Ionic Compounds

Nomenclature	Formula	Ions present
Aluminium oxide	Al_2O_3	$2Al^{3+}$ and $3O^{2-}$
Ammonium chloride	NH_4Cl	NH_4^+ and Cl^-
Calcium hydroxide	$Ca(OH)_2$	Ca^{2+} and $2OH^-$

Calcium nitrate	$Ca(NO_3)_2$	Ca^{2+} and $2NO_3^-$
Calcium oxide	CaO	Ca^{2+} and O^{2-}
Copper sulphate	$CuSO_4$	Cu^{2+} and SO_4^{2-}
Magnesium chloride	$MgCl_2$	Mg^{2+} and $2Cl^-$
Potassium chloride	KCl	K^+ and Cl^-
Potassium hydroxide	KOH	K^+ and OH^-
Sodium carbonate	Na_2CO_3	$2Na^+$ and CO_3^{2-}
Sodium hydroxide	$NaOH$	Na^+ and OH^-

Ionic compounds are made up of a metal and a non-metal (except ammonium chloride which is an ionic compound made up of only non-metals). So, whenever a bond involves a metal and a non-metal, we call it ionic bond.

EXERCISE 1.3

1. Give two examples of ionic compounds. Write their chemical formulae.
2. The overall charge on an ionic compound is zero. (True or False)
3. Name the ions present in calcium nitrate.
4. Ionic compounds are made up of a ______ and a ______ .
5. Give an example of an ionic compound made up of only non-metals.

1.4 FORMATION OF IONIC BOND

An ionic bond changes the electronic configurations of the atoms. Metal atoms lose their outermost electron(s), forming cations. Non-metal atoms gain electron(s) to fill their outermost shell, forming anions. The electrostatic force of attraction between the oppositely charged ions holds the ions together. For example,

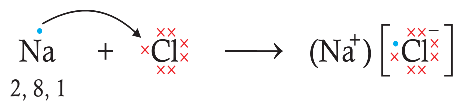
(*a*) When a hot sodium atom is placed in chlorine gas, a reaction takes place resulting in formation of sodium chloride.

$$\underset{2,\,8,\,1}{Na} \longrightarrow \underset{2,\,8}{Na^+} + e^-$$

Sodium ion
(Cation)

$$\underset{2,\,8,\,7}{Cl} + e^- \longrightarrow \underset{2,\,8,\,8}{Cl^-}$$

Chloride ion
(Anion)



Formation of Sodium Chloride

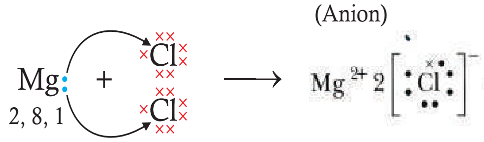
(*b*) When a magnesium atom comes in contact with chlorine gas, it forms magnesium chloride.

$$\underset{2,\,8,\,2}{Mg} \longrightarrow \underset{2,\,8}{Mg^{2+}} + 2e^-$$

Magnesium ion
(Cation)

$$\underset{2,\,8,\,7}{Cl} + e^- \longrightarrow \underset{2,\,8,\,8}{Cl^-}$$

Chloride ion
(Anion)



Formation of Magnesium Chloride

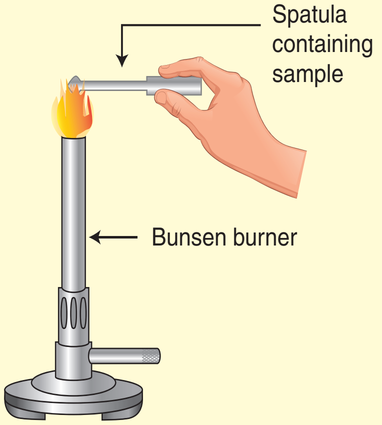
EXERCISE 1.4

1. With the help of dot and cross, show the formation of $CaCl_2$.
2. Which of the following is correct?
   (*a*) $[Na^+]\left[{}^{\bullet}_{\times}\overset{\times\times}{\underset{\times\times}{Cl}}{}^{\times}_{\times}\right]^-$ (*b*) $[Na^-]\left[{}^{\bullet}_{\bullet}\underset{\times\times}{Cl}{}^{\times}_{\times}\right]^-$
   (*c*) $[Na^+]\left[{}^{\times}_{\times}\overset{\times\times}{\underset{\times\times}{Cl}}{}^{\times}_{\times}\right]^+$ (*d*) $[Na]^+\left[{}^{\times}_{\times}\underset{\times\times}{Cl}{}^{\times}_{\times}\right]^-$
3. The electrostatic force of attraction holds the ions together. (True or False)

1.5 PROPERTIES OF IONIC COMPOUNDS

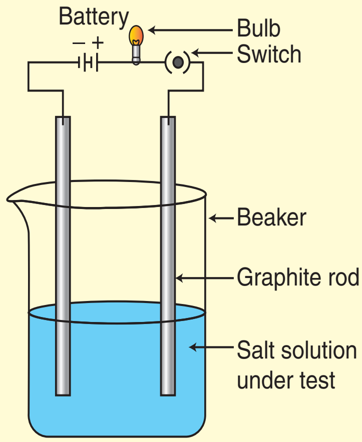
ACTIVITY 1.3: Illustrating Physical Properties of Ionic Compounds

- Take a sample of sodium chloride or any other salt from the science laboratory.
- What is the physical state of this salt?
- Take a small amount of a sample on a metal spatula and heat directly on the flame as shown in figure (a).



(a) Testing melting point of sodium chloride

- What did you observe? Did the sample impart any colour to the flame? Does this compound melt?
- Try to dissolve the sample in water, petrol and kerosene. Is it soluble?
- Make a circuit as shown in figure (b) and insert the electrodes into a solution of salt. What did you observe?



(b) Testing electrical conductivity of salt solution

- What is your inference about the nature of this compound?

You may have observed the following general properties of ionic compounds:

- Ionic compounds are usually crystalline solids.
- Ionic compounds have high melting and high boiling points.

The temperature at which a solid melts into a liquid is called the **melting point** of the solid. The temperature does not change during melting.

Boiling point is the temperature at which a liquid changes into a gas. The temperature of a liquid remains the same once boiling has started.

Table 1.6: Melting and Boiling Points of Some Ionic Compounds

Ionic compounds	Melting point (K)	Boiling point (K)
NaCl	1074	1686
LiCl	887	1600
$CaCl_2$	1045	1900
CaO	2850	3120
$MgCl_2$	981	1685

- Ionic compounds are usually soluble in water but insoluble in organic solvents like petrol and kerosene.
- Ionic compounds conduct electricity when dissolved in water or when melted. When we dissolve the ionic solid in water or melt it, the crystalline structure is broken down to form mobile ions. These ions help in conducting electricity.

EXERCISE 1.5

1. Why do ionic compounds conduct electricity when dissolved in water?
2. Ionic compounds are insoluble in
   (*a*) kerosene
   (*b*) petrol
   (*c*) both (*a*) and (*b*)
   (*d*) neither (*a*) nor (*b*)
3. Ionic compounds have low melting and boiling points. (True or False)
4. Ionic compounds are usually ______ solids.

1.6 COVALENT BONDING

The chemical bond formed by sharing of electrons between two atoms is known as a **covalent bond**. The compounds containing covalent bonds are known as covalent compounds. A covalent bond is formed when both the combining atoms need electrons to achieve the inert gas electron arrangement. Now, the non-metals have usually 5, 6 or 7 electrons in the outermost shells of their atoms. So, all the non-metal atoms need electrons to achieve the inert gas electronic structure. They get these electrons by mutual sharing. Thus, whenever a non-metal combines with another non-metal, covalent bond is formed.

Table 1.7: Formulae and Nomenclature of Some Covalent Compounds

Nomenclature	Formula	Elements present
Methane	CH_4	C and H
Ethane	C_2H_6	C and H
Ethene	C_2H_4	C and H
Ethyne	C_2H_2	C and H
Ammonia	NH_3	N and H
Alcohol (Ethyl alcohol)	C_2H_5OH	C, H and O
Hydrogen sulphide gas	H_2S	H and S
Carbon dioxide	CO_2	C and O
Carbon disulphide	CS_2	C and S

Carbon tetrachloride	CCl_4	C and Cl
Glucose	$C_6H_{12}O_6$	C, H and O
Cane sugar	$C_{12}H_{22}O_{11}$	C, H and O
Urea	$CO(NH_2)_2$	C, O, N and H

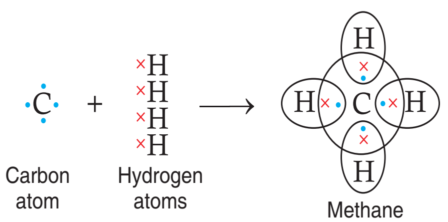
EXERCISE 1.6

1. What do you mean by a covalent bond?
2. Give two examples of covalent compounds. Also write their chemical formulae.
3. When a ______ combines with another ______, covalent bond is formed.
4. Choose the covalent compound(s).
   (*a*) CH_4 (*b*) H_2S (*c*) CS_2 (*d*) All of these

1.7 FORMATION OF COVALENT BOND

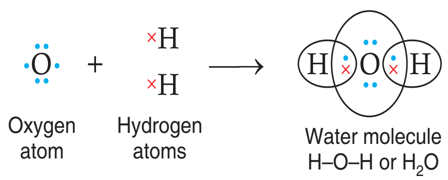
Covalent bonding between atoms of different elements.

(*i*) Carbon atom shares four electrons to form methane.



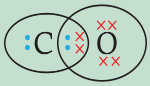
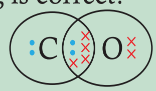
Covalent bonding or sharing of electrons only takes place in outermost shells of atoms to attain inert gas electron arrangement.

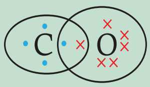
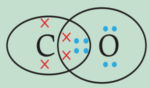
(*ii*) As in water molecule, 2 hydrogen atoms share electrons with oxygen atom.



EXERCISE 1.7

1. Using cross and dot diagram, show the formation of carbon dioxide.
2. Which of the following is correct?

(*a*)  (*b*) 

(*c*)  (*d*) 

3. Carbon tetrachloride and Urea are not covalent compounds. (True or False)
4. In the formation of covalent bonding, ______ of electrons takes place in the ______ shells of atoms.

1.8 PROPERTIES OF COVALENT COMPOUNDS

ACTIVITY 1.4: Illustrating Physical Properties of Covalent Compounds

Let us test some covalent compounds in different ways:

- Take sample of cooking oil. Try to dissolve it in water and ethanol. Does it dissolve?

- Have you ever observed a burning candle wax? If not, take a candle wax and observe it burning. How much time does it take to melt down?
- Take a pan and add water to it. Let it boil. Do you know the boiling point of water?
- Now add two electrodes to the water pan making a circuit. What did you observe? What would have happened if you would have added NaCl salt in the pan?
- What can you now say about these covalent compounds?

You have observed the following properties of covalent compounds:

Covalent compounds are usually liquids,

- Covalent compounds are usually liquids, gases or solids of low melting and boiling points. For example, alcohol, benzene, water and cooking oil are liquids. Methane, ethane and chlorine are gases. Glucose, urea, and wax are solid covalent compounds.
- Covalent compounds have usually low melting points and low boiling points.
- Covalent compounds are usually insoluble in water, but they are soluble in organic solvents. Some of the covalent compounds like glucose, sugar and urea are soluble in water.
- Covalent compounds do not conduct electricity because they do not contain ions.

ACTIVITY 1.5: Detecting an Ionic Bond or Covalent Bond

- Take the sample such as common salt (NaCl) provided.
- Try to dissolve it in water.
- If it dissolves, chances are it is likely to be an ionic compound. But, you already know some covalent compounds like glucose, urea and sugar are soluble in water.
- Now, perform electrical conductivity test.
- If the NaCl sample dissolves in water, arrange a circuit with two electrodes and a bulb.
- Figure out whether the bulb glows or not. According to your observation conclude the bond present in the sample.
- Make a report on the properties of ionic and covalent compounds.

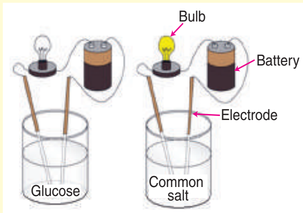


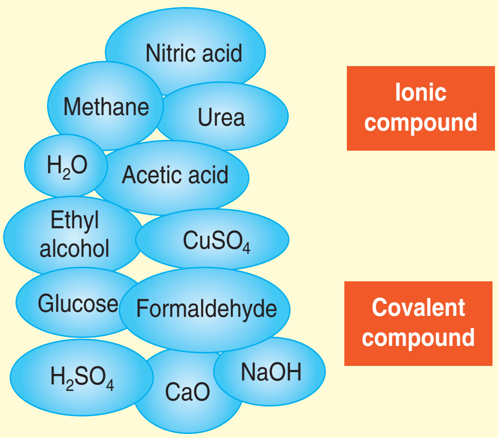
Table 1.8: Differences between Ionic Compounds and Covalent Compounds

Ionic compounds	Covalent compounds
Ionic compounds are usually crystalline solids.	Covalent compounds are usually liquids, gases or solids of low melting and boiling points.
Ionic compounds have high melting points and boiling points. That is, ionic compounds are non-volatile.	Covalent compounds have usually low melting and boiling points.
Ionic compounds conduct electricity when dissolved in water or melted.	Most covalent compounds do not conduct electricity.

Ionic compounds are usually soluble in water.	Covalent compounds are usually insoluble in water (except, glucose, sugar, urea, etc.).
Ionic compounds are insoluble in organic solvents (like alcohol, ether, acetone, etc.)	Covalent compounds are soluble in organic solvents.

ACTIVITY 1.6: Identifying Ionic and Covalent Compounds

Choose the ionic as well as covalent compounds from the bubbles and make a table in your exercise notebook.



EXERCISE 1.8

1. Some covalent compounds are solid. (True or False)
2. Most covalent compounds are ______ in water but ______ in organic solvents.
3. Name two covalent compounds which are soluble in water.
4. Why most covalent compounds do not conduct electricity?
5. Melting and boiling points of covalent compounds are
   (*a*) high
   (*b*) low
   (*c*) between 500°C and 1000°C
   (*d*) cannot be determined.

1.9 GIANT COVALENT STRUCTURES

Diamond, graphite and silicon dioxide have giant covalent structures.

1.9.1 Diamond and its Properties

Diamond is a colourless transparent substance having extraordinary brilliance. Diamond is quite heavy. Diamond is extremely hard. It is the *hardest* natural substance known. Diamond does not conduct electricity. Diamond burns on strong heating to form carbon dioxide. It has a very high melting point. *If we burn diamond in oxygen, then only carbon dioxide gas is formed and nothing is left behind. This shows that diamond is made up of carbon only*. Since diamond is made up of carbon atoms only, its symbol is taken to be C.

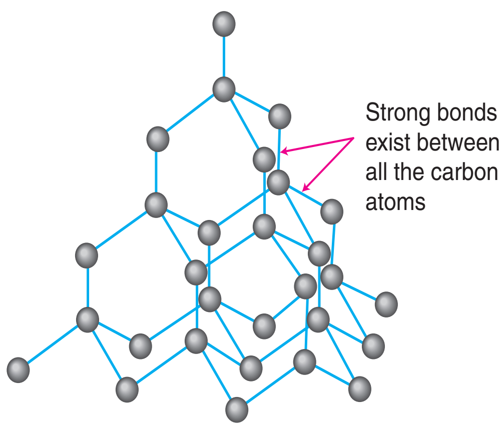


Figure 1.3: Structure of diamond (The black balls represent carbon atoms).

1.9.2 Graphite and its Properties

Graphite is a greyish-black opaque substance. Graphite is lighter than diamond. Graphite is soft and slippery to touch. Graphite conducts electricity. Graphite burns on strong heating to form carbon dioxide. Like diamond, graphite also has very high melting point. *If we burn graphite in oxygen, then only carbon dioxide gas is formed and nothing is left behind. This shows that graphite is made up of carbon only*. Since graphite is made up of carbon atoms only, its symbol is taken to be C.

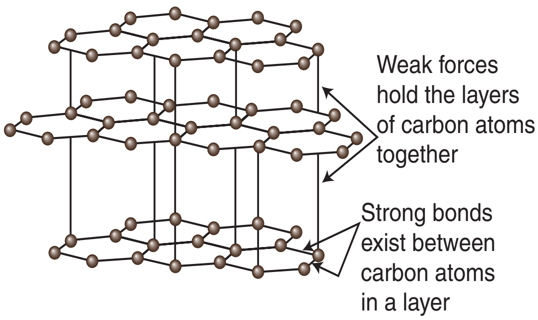


Figure 1.4: Structure of graphite (The black balls represent carbon atoms).

1.9.3 Silicon Dioxide and its Properties

Silicon dioxide (also known as Silica) has a giant covalent structure. Each silicon atom is covalently bonded to four oxygen atoms. Each oxygen atom is covalently bonded to two silicon atoms. This means that, overall, the ratio is two oxygen atoms to each silicon atom, giving the formula SiO_2. Silicon dioxide is very hard. It has a very high melting point (1,610°C) and boiling point (2,230°C). It is insoluble in water, and does not conduct electricity. These properties result from the very strong covalent bonds that hold the silicon and oxygen atoms in the giant covalent structure. Silicon dioxide is found as quartz in granite, and is the major compound in sandstone. The sand on a beach is made mostly of silicon dioxide.

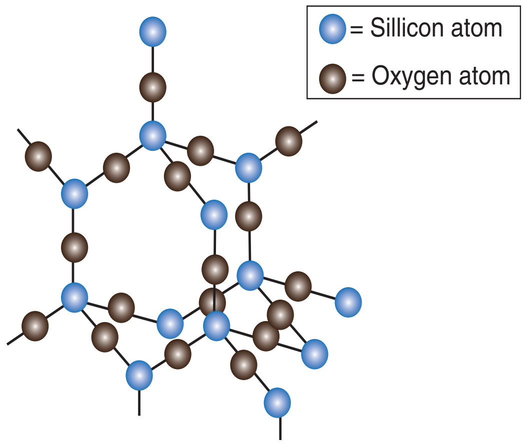


Figure 1.5: Structure of silicon dioxide

1.9.4 Uses of Diamond, Graphite and Silicon Dioxide

Uses of Diamond

- Since diamond is extremely hard, it is used for cutting and grinding other hard materials. It is also used for drilling holes in the earth's rocky layers. Diamond 'dies' are used for drawing thin wires like the tungsten filament of an electric bulb.
- Diamonds are used for making jewellery. The use of diamonds in making jewellery is because of their extraordinary brilliance. Diamond is also used in the tip of glass cutter. A sharp diamond-edged knife called *keratome* is used by eye surgeons to remove cataract from the eyes.

Glass Cutter Keratome Jewellery

Figure 1.6: Some of the uses of diamond.

Uses of Graphite

Due to its softness, powdered graphite is used as a lubricant for fast moving parts of machinery. Graphite can be used as a dry lubricant in the form of graphite powder or mixed with petroleum jelly to form graphite grease. Graphite powder can also be mixed with lubricant oils.

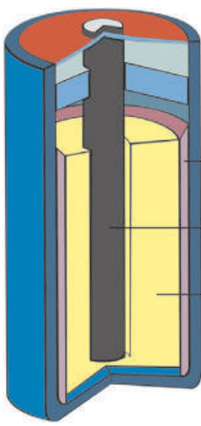


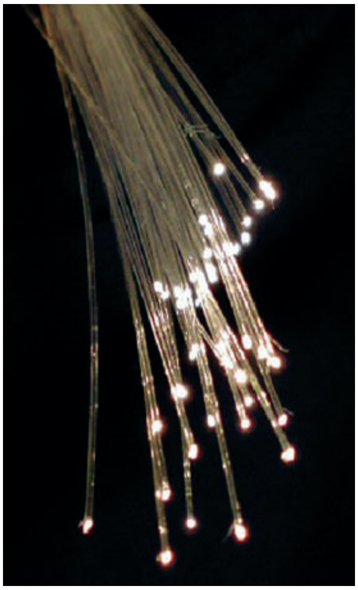
Figure 1.7: Some of the uses of graphite.

- Graphite is a good conductor of electricity due to which graphite is used for making carbon electrodes or graphite electrodes in dry cells and electric arcs. The black coloured 'anode' of a dry cell is made of graphite. The carbon brushes of electric motors are also made of graphite.
- Graphite is used for making the cores of our pencils called 'pencil leads' and black points. Graphite is black in colour and quite soft. So, it marks black lines on paper. Due to this property, graphite is used for making pencil leads. For making pencil leads, graphite is usually mixed with clay.

Uses of Silicon Dioxide

- An estimated 95% of silicon dioxide produced is consumed in the construction industry, e.g. for the production of cement
- Silica is used primarily in the production of glass for windows, drinking glasses, beverage bottles, and many other uses.
- The majority of optical fibres for telecommunication are also made from silica.

Glass made from silicon dioxide (Silica)



Bundle of optical fibres composed of high purity silica

Figure 1.8: Some of the uses of silicon dioxide

EXERCISE 1.9

1. Diamond and Graphite are two common allotropes of Carbon. (True or False)
2. Which of the following is correct?
   (*a*) Diamond is the hardest substance known.
   (*b*) Graphite has very low melting point.
   (*c*) Graphite does not conduct electricity.
   (*d*) Diamond burns on strong heating to form helium gas.
3. Why are diamonds used for making jewellery?
4. Graphite is used for making
   (*a*) pencil lead (*b*) electrodes
   (*c*) both (*a*) and (*b*)
   (*d*) none of these.
5. Diamond and Graphite have very ______ melting point.

1.10 METALLIC BONDING

The force which binds various metal atoms together is called **metallic bond**. The metallic bond is neither a covalent bond nor an ionic bond because these bonds are not able to explain properties of metals.

For example, metals are very good conductors of electricity but in solid state. Both ionic and covalent compounds cannot do so with the exception of graphite.

1.11 FORMATION OF METALLIC BOND

Loreutz proposed the theory of electron gas model or **electron sea model** for metallic bonding.

In this model, the metal is pictured as an array of metal cations in a "sea" of electrons. The atoms in a metallic solid contribute their valence electrons to form a "sea" of electrons that surrounds metallic cations. Delocalised electrons are not held by any specific atom and can move easily throughout the solid. A metallic bond is the attraction between these electrons and the metallic cation.

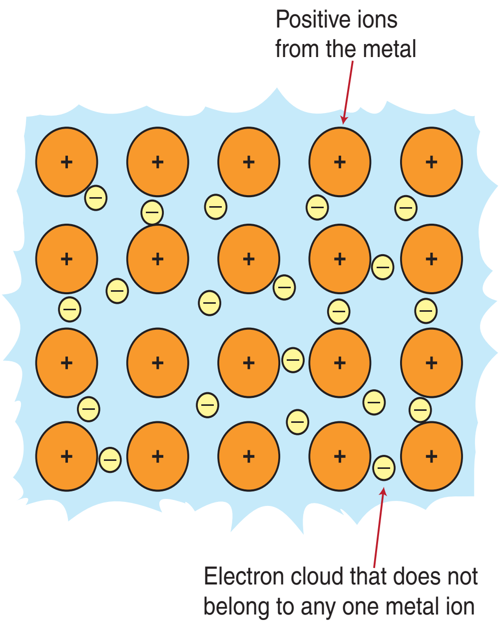


Figure 1.9: Formation of metallic bond.

EXERCISE 1.10

1. Name the scientist who proposed the theory of electron sea model.
2. Metallic bond is neither a covalent bond nor an electrovalent bond. (True or False)
3. The force which binds various metal atoms together is called ______ .
4. Make a 3D structure of electron sea model.
5. Write a short note on formation of metallic bonding.

1.12 PROPERTIES OF METALLIC BOND

ACTIVITY 1.7: Illustrating the Properties of Metals

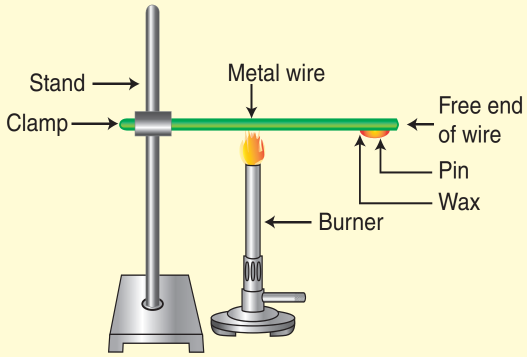
- Take samples of iron, copper, aluminium, sodium, carbon and iodine. Note the appearance of each sample.
- Clean the surface of each sample by rubbing them with sand paper and note their appearance again.
- Try to cut these elements with a sharp knife and note your observations.
- Hold a piece of sodium with a pair of tongs.
  Caution: Always handle sodium with care. Dry it by pressing between the folds of a filter paper.
- Put it on a watch-glass and try to cut it with a knife.
- What do you observe?
- Place any one element on a block of iron and strike it four or five times with a hammer. What do you observe?
- Repeat above steps with other elements.
- Record the change in the shape of these elements.
- Which of the above elements are available in the form of wires?

1.12.1 Properties of Metals

ACTIVITY 1.8: Illustrating Conductivity of Heat and Electricity of Metals

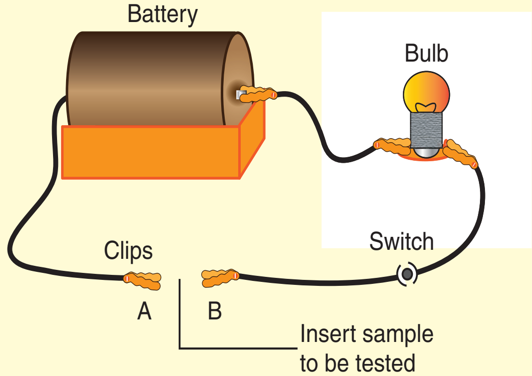
- Take an aluminium or copper wire. Clamp this wire on a stand, as shown in Figure (a).

- Fix a pin to the free end of the wire using wax.
- Heat the wire with a spirit lamp, candle or a burner near the place where it is clamped.
- What do you observe after some time?
- Repeat the same with carbon or sulphur.
- Note your observations.
  Does the element melt?



(a)

- Consider elements aluminium, copper, sulphur and carbon.
- Set up an electric circuit as shown in Figure (b).



(b)

- Place the element to be tested in the circuit between terminals A and B as shown.
  Does the bulb glow? What does this indicate?
- Compile your observations regarding properties of elements in your exercise book.

- Metals are good conductors of heat and electricity. This means that metals allow heat and electricity to pass through them easily. Silver metal is the best conductor of heat. Copper metal is a better conductor of heat than aluminium metal.
- Metals are lustrous (or shiny). This means that metals have a shiny appearance.

- Metals are usually **strong**. For example, iron metal (in the form of steel) is very strong when freshly cut and is used in the construction of bridges, buildings and vehicles. Some metals are not strong. For example, sodium and potassium.
- Metals are **ductile**. This means that metals can be drawn (or stretched) into thin wires.
- Gold and silver are among the best ductile metals.
- Metals are **malleable**. This means that metals can be hammered into thin sheets.

The cooking utensils are made of metals because metals are good conductors of heat.

EXERCISE 1.11

1. Name the metal which is the best conductor of electricity.
2. Aluminium is a better conductor of heat than copper. (True or False)
3. Metals are _______. This means that they can be hammered into thin sheets.
4. Why are cooking utensils made of metals?
5. Which of the following statement(s) is/are correct for metals?
   (*a*) Metals such as sodium and potassium are not strong.
   (*b*) Iron is used in the construction of buildings.
   (*c*) Gold and Silver are among the best ductile metals.
   (*d*) All of these.

1.13 SUMMARY

- An atom achieves a stable electronic configuration by losing, gaining or sharing electrons.
  - ♦ Metal atoms with one, two or three electrons in the outermost shell lose electron(s) to form positively charged ions (cations).
  - ♦ Non-metal atoms with five, six or seven electrons in the outermost shell gain three, two and one electron(s) to form negatively charged ions (anions).
  - ♦ Non-metal atoms with four to seven outermost electrons may attain a stable electronic configuration by sharing electrons with each other..
- A chemical bond is a force that holds ions or atoms together in a compound a molecule or a metal. A bond is formed when each atom acquires a stable electronic configuration like noble gas.
- The electrostatic binding force is called an ionic bond or electrovalent bond.
- Ionic compounds are formed by attraction of positive and negative ions. These compounds are crystalline solid. They conduct electricity. Ionic compounds have high melting and boiling points.
- A covalent bond forms between two or more atoms of non-metals that cannot combine by forming ions.

- Covalent compound is formed when atoms achieve a stable electronic configuration by sharing of electrons. Covalent compounds are solids, liquid or gases. Covalent compounds have low melting and boiling points.
- The two forms of carbon that join covalently to form giant structure are diamond and graphite.
- The force which binds various metal atoms together is called metallic bond.
- Metals are generally hard, lustrous, strong, malleable and ductile. They conduct heat and electricity in both molten and solid state.

1.14 GLOSSARY

- **Anion:** a negatively charged ion.
- **Cation:** a positively charged ion.
- **Crystal:** a solid where the atoms form a periodic arrangement.
- **Diamond:** one of the known allotropes of carbon.
- **Ductile:** able to be drawn out into a thin wire.
- **Electronic configuration:** the distribution of electrons of an atom.
- **Graphite:** a grey crystalline allotropic form of carbon which occurs as a mineral in some rocks.
- **Malleable:** able to be hammered or pressed into thin sheet without breaking or cracking.
- **Noble gas:** the gaseous elements helium, neon, argon, krypton, xenon, and radon.

1.15 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Sulphur, phosphorous and chlorine are all non-metals. which one of them gain one electron to achieve a noble gas configuration
   a) phosphorous b) chlorine c) sulphur d) chlorine and sulphur
2. The electronic configuration of sodium ion is
   (*a*) 2,8,1 (*b*) 2,8,8 (*c*) 2,8 (*d*) 2,8,2
3. The electronic configuration of chloride ion is
   (*a*) 2,8 (*b*) 2,8,8 (*c*) 2,8,7 (*d*) 2,8,3
4. Choose the ionic compound.
   (*a*) Calcium chloride (*b*) Copper sulphate
   (*c*) Sodium hydroxide (*d*) All of these
5. Most ionic compounds are soluble in
   (*a*) water (*b*) petrol (*c*) kerosene (*d*) all of these

6. Which of these is not a covalent compound?
   (*a*) Carbon dioxide (*b*) Methane (*c*) Ammonia (*d*) None of these
7. Choose the correct statement.
   (*a*) Covalent compounds have low melting points
   (*b*) Ionic compounds have high melting points
   (*c*) Urea and glucose are solid covalent compounds
   (*d*) All of these
8. Graphite is used for making ____________.
   (*a*) lubricant oils (*b*) pencil leads (*c*) both (a) and (b) (*d*) jewellery
9. If we burn diamond, the product formed is ____________.
   (*a*) carbon dioxide (*b*) hydrogen gas
   (*c*) hydrogen chloride gas (*d*) oxygen gas
10. The force which binds various metal atoms together is called ____________.
   (*a*) metallic bond (*b*) covalent bond (*c*) ionic bond (*d*) none of these

II. Open Ended Questions

1. How can an atom achieve stability?
2. Distinguish between covalent and ionic bond.
3. Compare between the properties of ionic and covalent compounds.
4. Explain the formation of sodium ion.
5. Give five examples of each
   (*a*) Ionic compounds (*b*) Covalent compounds
6. Compare the conductivity of distilled water with sodium chloride solution.
7. Write two uses of diamond.
8. Draw the structure of graphite.
9. Describe the physical properties of metals.

III. Practical-based Questions

1. Look at the figures and choose the correct statement.

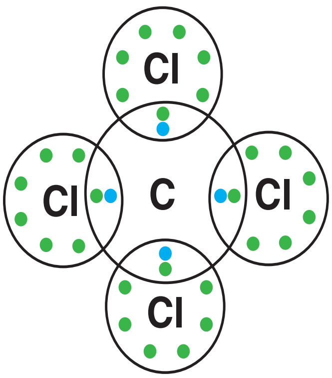


Figure A

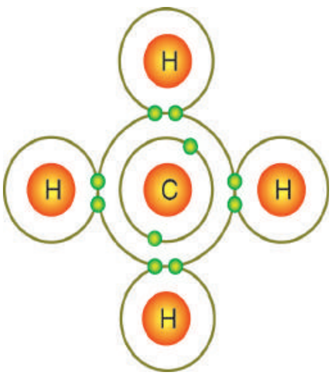
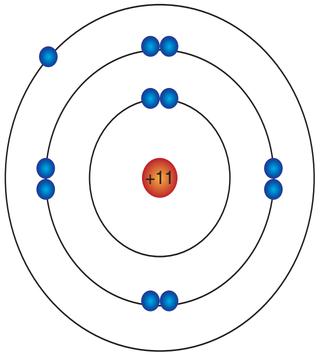


Figure B

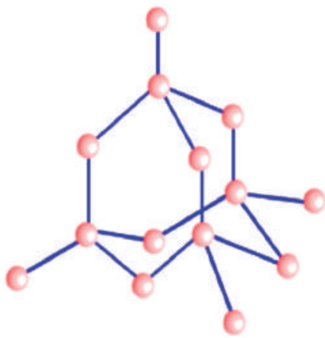
(*a*) Figure A is an example of ionic compound
(*b*) Figure B is not an example of covalent compound
(*c*) Both Figure A and Figure B are covalent compounds
(*d*) None of these

2. The following figure illustrates the electronic configuration of



(*a*) Lithium (*b*) Sodium (*c*) Chlorine (*d*) Helium

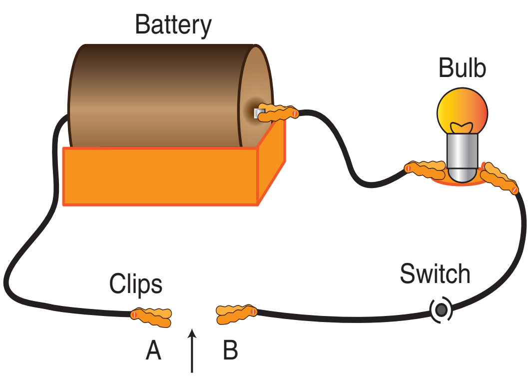
3.



shows the structure of ....

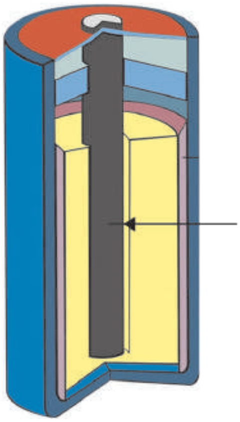
(*a*) the hardest substance known
(*b*) an allotrope of carbon
(*c*) both (*a*) and (*b*)
(*d*) none of these

4. Which of the following materials makes the circuit complete when inserted in between the crocodile clips?



(*a*) Aluminium foil
(*b*) Copper wire
(*c*) Both (*a*) and (*b*)
(*d*) Sulphur

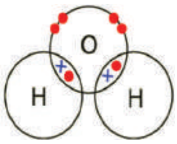
5. In the given figure, arrow shows the



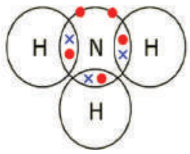
(*a*) carbon rod (*b*) iron rod (*c*) brass rod (*d*) copper rod

6. Which of the following depicts the molecule of water?

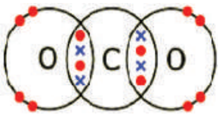
(*a*)



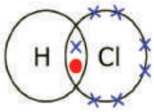
(*b*)



(*c*)



(*d*)



PROJECT

Make a 3D model of diamond and graphite and discuss their physical properties.

Unit 2

Trends in Properties of Elements in the Periodic Table

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- describe trends in reactive elements with acids, water, and halogens.
- explain the trends in the physical properties across a period and down a group.

KNOWLEDGE GAIN

Jons Jakob Berzelius was a Swedish chemist and one of the founders of modern chemistry. He proposed the first letter (or first letter and another letter) of the name of an element as its symbol.

2.1 CLASSIFICATION OF ELEMENTS

ACTIVITY 2.1: Distinguishing Metallic and Non-metallic Objects

Collect five objects made of metals. Also collect five objects made of non-metals. Compare the physical properties of metallic objects and non-metallic objects.

Observe Figure 2.1 of the periodic table. There are 118 chemical elements known at present. These elements are classified into metals, metalloids and non-metals. The metals appear at the left-hand side and middle part of the periodic table. The non-metals appear at the right-hand side of the periodic table (Figure 2.1). Metalloids lie in between metals and non-metals.

ACTIVITY 2.2: Categorising Elements into Metals, Non-metals and Metalloids

In groups, classify all elements of periodic table into metals, metalloids and non-metals.

Metals | Metalloids | Non-metals

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Period 1	1 H																	2 He
Period 2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
Period 3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
Period 4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
Period 5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
Period 6	55 Cs	56 Ba	*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
Period 7	87 Fr	88 Ra	**	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

Lanthanide series (*)	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinide series (**)	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Figure 2.1: Periodic table.

In the periodic table, there is a regular variation in the properties of elements in groups and periods.

2.1.1 Variation in Metallic and Non-metallic Character Across a Period

Metallic character is a measure of the tendency of atoms to lose their electrons.
From this non-metallic character can be defined as a measure of the tendency of atoms to gain electrons.

Observe Figure 2.2 (*a*).

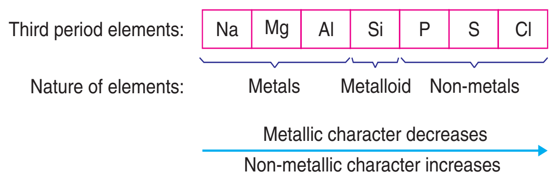


Figure 2.2: (a) Third period of periodic table.

In the third period of the periodic table shown in Figure 2.2(*a*), sodium, magnesium and aluminium are metals. The properties of silicon are in-between those of a metal and a non-metal; therefore, silicon is a metalloid. The next elements – phosphorus, sulphur and chlorine are non-metals. The metallic character decreases from sodium to magnesium to aluminium; silicon is a metalloid; and the non-metallic character increases from phosphorus to sulphur to chlorine. Thus, in the third period of the periodic table, sodium is the most metallic element whereas chlorine is the most non-metallic element. In general, we can say that *the greatest metallic character is found in the elements on the extreme left side of a period and the greatest non-metallic character is found in the elements on the right side of a period.*

On moving from left to right in a period, the metallic character of elements decreases. The non-metallic character increases on moving from left to right in a period. On the left side in a period, we have metals and on the right side we have non-metals. Some elements in-between the metals and non-metals are known as **metalloids**.

EXERCISE 2.1

1. On moving from left to right in a period, the metallic character of elements ____ .
2. How many metals are there in third period of periodic table?
3. The non-metallic character increases on moving from left to right in a period. (True or False)
4. Choose the symbol of metalloid from the following.
   (*a*) Na (*b*) Si
   (*c*) S (*d*) Ne

2.1.2 Variation in Metallic and Non-metallic Character down a Group

Observe Figure 2.2(*b*). *On going down in a group of the periodic table, the metallic character of elements increases.* For example, when we move down in group 1 of the periodic table, the metallic character increases from lithium to francium.

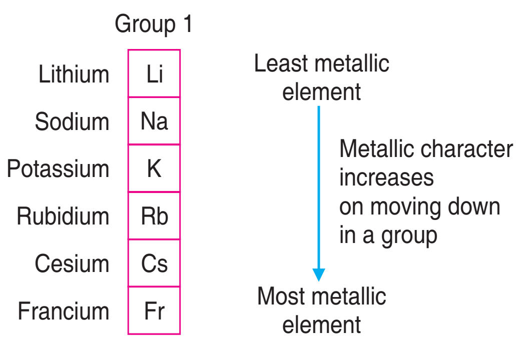


Figure 2.2: (b) Group 1 of periodic table.

Thus, in group 1 of alkali metals, lithium is the least metallic element whereas francium is the most metallic element. It is obvious that *the greatest metallic character is found in the elements in the lowest part of a group.*

We can also say that *on going down in a group of the periodic table, the non-metallic character of elements decreases.* Observe Figure 2.2(*c*). When we go down in group 17 of the halogen elements, the non-metallic character decreases from fluorine to iodine. Thus, out of fluorine, chlorine, bromine and iodine, fluorine has the most non-metallic character whereas iodine has the least non-metallic character.

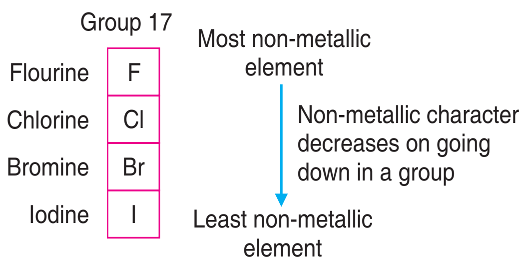


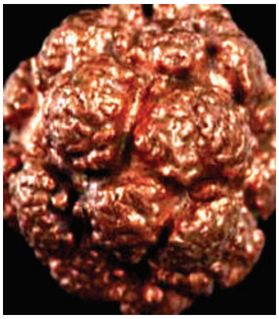
Figure 2.2: (c) Group 17 of periodic table.

EXERCISE 2.2

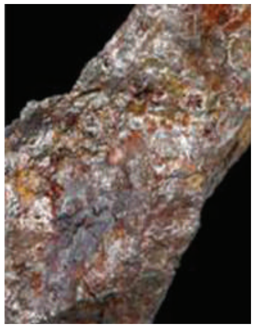
1. On moving down in a group of the periodic table, the non-metallic character of elements ______ .
2. In group 17, Iodine has the least non-metallic character. (True or False)

2.1.3 Metals

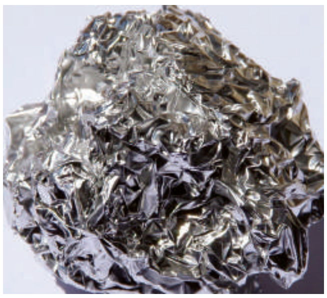
The majority of known elements (about 80%) are metals. All the metals are solid, except mercury which is a liquid metal at room temperature and pressure. Metals are elements which conduct electricity and heat. Metals are also shiny, hard and produce ringing sound when struck. Metals are widely used in our daily life for a large number of purposes. The common metals we use are iron, copper, aluminium, tin, zinc, gold, etc. (Figure 2.3). The electric fan, machines, bicycle, cars, aeroplane, cooking utensils are all made of metals or mixture of metals.



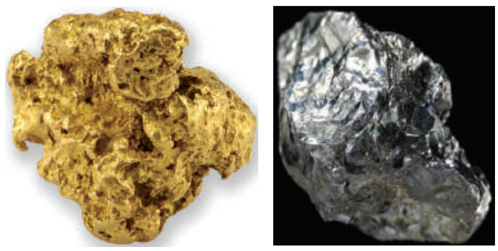
Copper



Iron



Aluminium



Gold and silver

Figure 2.3: (a) Examples of some metals.

During a chemical reaction, metals can form positive ions by losing electrons. **Metals are the elements** which form positive ions by losing electrons.

For example, sodium (Na) is a metal which forms positively charged sodium ion (Na^+) by losing one electron.

Metals are very important for the national economy of every country. The economy and prosperity of a country are dependent on the natural resources produced by the country. The main metal deposits in our country are tin, coltan and tungsten.

The major metals in the earth's crust are aluminium, iron, calcium, sodium, potassium and magnesium. Aluminium is the most abundant metal in the earth's crust.

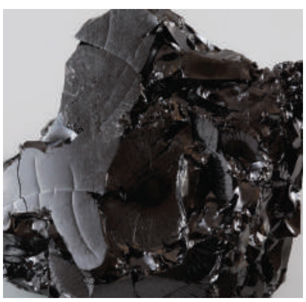
EXERCISE 2.3

1. Name the liquid metal.
2. A majority of known ______ (about 80%) are metals.
3. Metals are very important for the economic growth of every country. (True or False)
4. Which is the most abundant metal in the earth's crust?
   (*a*) Aluminium (*b*) Iron
   (*c*) Sodium (*d*) Calcium
5. The main metal deposits in our country are ______ and _____.

2.1.4 Non-metals

There are only 22 non-metals. Out of these, 10 non-metals are solid, 1 non-metal is liquid and the remaining 11 are gases. Thus, all non-metals are solids and gases, except bromine which is a liquid non-metal.

Non-metals are elements which do not conduct heat and electricity. The only exception is graphite. Non-metals are brittle, and have dull appearance. They are soft, not strong and not shiny.



Carbon

Sulphur

Iodine

Figure 2.3: (b) Examples of some non-metals.

During a chemical reaction, non-metals form negative ions by gaining electrons. So we define non-metals as elements which form negative ions by accepting electrons. For example, Chlorine (Cl) is a non-metal which forms negatively charged chloride ion (Cl^-) by gaining one electron.

Non-metals are small in number as compared to metals but they play an important role in our everyday life. In fact, life would not have been possible without the presence of non-metals on earth. For example, oxygen (non-metal) is essential for breathing to maintain life. Another non-metal carbon is one of the most important elements for existence of life on earth. This is because carbon compounds like proteins, carbohydrates, vitamins and fats are essential for growth and development of living organism. Non-metals are also used to make vegetable oil, acids, fertilisers, and fungicides.

The major non-metals in the earth's crust are oxygen, silicon, phosphorus and sulphur. Oxygen is the most abundant non-metal in the earth's crust.

Non-metals are the major constituents of air, earth's crust and oceans.

For example, oxygen and nitrogen are the major constituents of air.

Oxygen and silicon are the major constituents of earth's crust.

Oxygen and hydrogen are the major constituents of oceans.

EXERCISE 2.4

1. Name the only non-metal which is liquid.
2. How many non-metals are found in gaseous state?
3. The only non-metal which conducts electricity is graphite. (True or False)
4. Which is the most abundant non-metal in the earth's crust?
   (*a*) Carbon (*b*) Oxygen
   (*c*) Sulphur (*d*) Nitrogen
5. Non-metals are used to make ______ and ______ .

2.1.5 Metalloids

Metalloids are the elements found along the stair-step line that distinguishes metals from non-metals. Metalloids have properties of both metals and non-metals. Some of the metalloids, such as silicon and germanium, are semiconductors. This means that they can carry an electrical charge under special conditions. This property makes metalloids useful in computers and calculators.

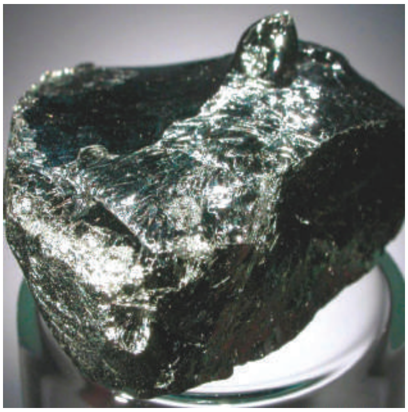
The metalloids (Figure 2.4) are:

- Boron
- Silicon
- Germanium
- Arsenic
- Antimony
- Tellurium
- Polonium

Metalloids tend to be economically important because of their unique conductivity properties (they only partially conduct electricity), which makes them valuable in the semiconductor and computer chip industry.

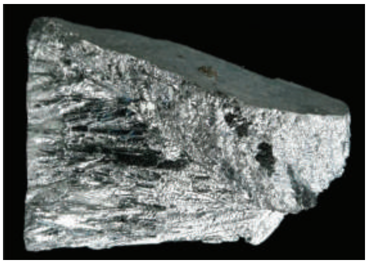
Boron

Silicon

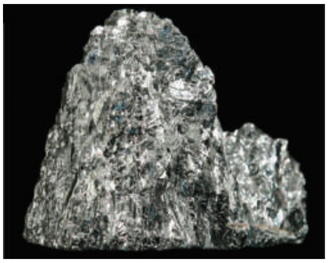


Germanium

Arsenic



Tellurium



Antimony

Polonium

Figure 2.4: Metalloids.

EXERCISE 2.5

1. Name two metalloids.
2. Metalloids have properties of both metals and non-metals. (True or False)
3. Silicon is used in making ______.
4. Which of the following is/are metalloid(s)?
   (*a*) Boron (*b*) Germanium (*c*) Antimony (*d*) All of these

2.2 PHYSICAL PROPERTIES OF METALS

ACTIVITY 2.3: Illustrating Physical Properties of Metals

- Collect some metallic objects.
- Note physical properties of collected objects (Get more information from internet, if available).
- Make a report on the findings.

All metals, such as gold, silver, iron,copper, aluminium, etc are solid at room temperature except mercury which is a liquid.

Copper is reddish-brown in colour whereas gold is yellow.

The other important physical properties of metals are:

1. **Metals are lustrous:** Metals are lustrous, that is, they have a shining surface. The shining surface of metals makes them useful in making jewellery and decorative items. For example, gold and silver are used for making jewellery.

Gold jewellery

Figure 2.5: Lustrous objects.

A metal has a shining surface only when it is freshly cut. On exposure to air, metals lose their brightness due to the formation of oxide and carbonate on their surface. This is known as corrosion. If we rub the corroded metal (dull surface of metal) with sand paper, the outer corroded layer is removed and the metal object becomes shiny once again.

ACTIVITY 2.4: Illustrating the Appearance of Metals

Take samples of iron, copper, aluminium and magnesium. Note the appearance of each sample.

Clean the surface of each sample by rubbing them with sand paper and note their appearance again.

2. **Metals are hard and strong:** Most of the metals are hard but all metals are not equally hard. The hardness of metals varies from metal to metal. Metals such as iron, aluminium, and copper are very hard. They cannot be cut with a knife. Sodium and potassium are soft metals. These metals can be cut easily with a knife.

   Generally metals are strong. But some metals like sodium and potassium are soft. Metals can hold large weight without breaking. For example, iron in the form of steel is very strong. Due to this, iron metal is used in machines, chains and vehicles.

ACTIVITY 2.5: Illustrating the Variation of Hardness in Metals

- Take small pieces of iron, copper, aluminium, and magnesium. Try to cut these metals with a sharp knife and note your observations.
- Hold a piece of sodium metal with a pair of tongs.

  Caution: Always handle sodium metal with care.

  Dry it by pressing between the folds of a filter paper.

  Put it on a watch glass and try to cut it with a knife.

  What do you observe?

3. **Metals are malleable:** Metals can be beaten (hammered) into very thin sheets without breaking. This property of metals is called malleability. Gold and silver are the most malleable metals. These can be hammered into very thin sheets called foils. Aluminium foils are used for packing medicines and cigarettes.

ACTIVITY 2.6: Illustrating the Malleability of Metals

- Take pieces of iron, zinc, lead and copper.
- Place any one metal on a block of iron and strike it four or five times with a hammer. What do you observe?
- Repeat with other metals.
- Record the change in the shape of these metals.

4. **Metals are ductile:** Metals are ductile. They can be drawn into thin wires. All the metals are not equally ductile. Gold and silver are the best ductile metals. Just one gram of gold can be stretched into a wire of about 2 km length. Copper and aluminium are also very ductile. Their wires are used in electrical wiring.

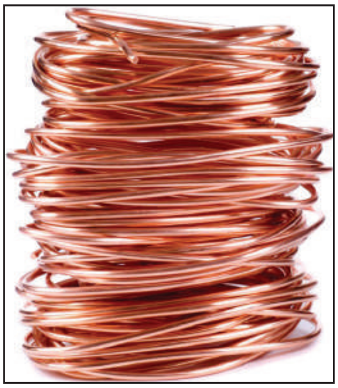


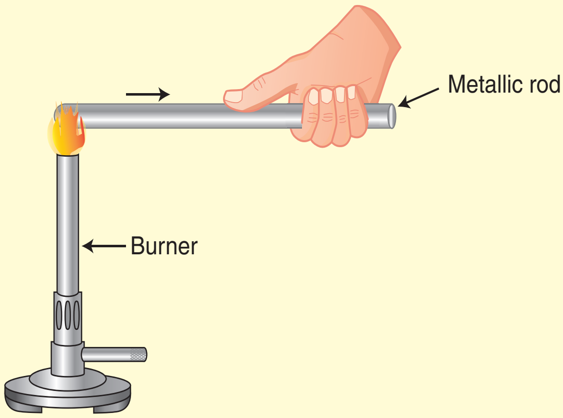
Figure 2.6: Copper wire.

5. **Metals are good conductors of heat:** Metals allow heat to pass through them easily. *The conduction of heat is also known as thermal conductivity*.

ACTIVITY 2.7: Illustrating Thermal Conductivity in Metals

Caution: Be careful while heating the objects.

- Take a steel spoon, a brass key, aluminium or copper wire (10 cm), and iron rod.
- Light the burner.
- Hold one end of iron rod in your hand.
- Keep the other end of iron rod to the flame of burner for 3 to 4 minutes as shown in figure.



Conduction of heat in metals

What do you feel?

- Repeat the activity with other metallic objects.
- State your observation in each case. Does the metal wire melt?

Deduction

The activity tells that metals are good conductors of heat and have high melting points. The best conductors of heat are silver and copper. Lead and mercury are comparatively poor conductors of heat.

6. **Metals are good conductors of electricity:** Metals allow electricity to pass through them.

Silver metal is the best conductor of electricity, copper metal is the next best conductor of electricity followed by gold, aluminium and tungsten.

Metals are good conductors of electricity because they contain free electrons. These free electrons can move easily through the metal and conduct electric current. Thus, electrical conductivity is another characteristic property of metals. From the above discussion we conclude that metals are good conductors of heat and electricity.

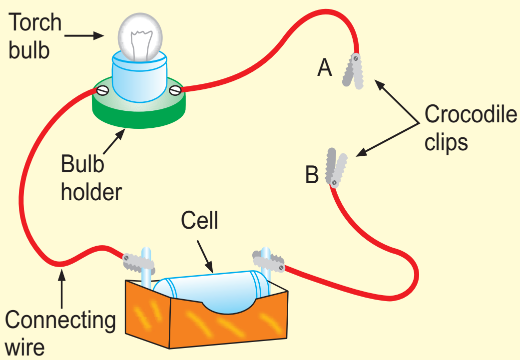
The electric wires that carry current in our homes have a covering of plastic such as poly vinyl chloride (PVC). Polyvinyl chloride is an insulator. It does not allow electric current to pass through it. The electric wires have a covering of an insulating material (like PVC) around them so that even if we happen to touch them, the current will not pass through our body and hence we will not get an electric shock.

ACTIVITY 2.8: Showing that a Metal Conducts Electricity

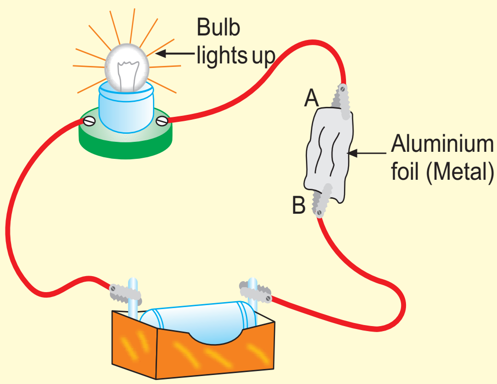
We take a dry cell, a torch bulb fitted in a holder and some connecting wires (copper wires) with crocodile clips, and connect them [as shown in Figure *a*] to make an electric circuit. There is a gap between the ends of the crocodile clips *A* and *B* so no current flows in the incomplete circuit shown in Figure (*a*) and hence the bulb does not light up. Let us now insert

(i) a piece of aluminium foil between the ends of crocodile clips *A* and *B* as shown in Figure (*b*)

(ii) an iron rod between the ends of crocodile clips A and B to replace aluminium foil in Figure (b).



(*a*)



(*b*)

Deduction

In both cases, we see that the bulb lights up at once. This means that both aluminium foil and iron rod allow electric current to pass through them. In other words, aluminium metal is a good conductor of electricity. Note that the connecting wires used in this experiment are made of copper metal. Since these copper connecting wires allow electric current to pass through them, copper metal is also a good conductor of electricity.

7. **Metals have high melting points:** Metals melt and turn into liquid at very high temperatures. However, there are some exceptions. Sodium and potassium have low melting points. Melting points of some metals are given in Table 2.1.

8. **Metals have high densities:** The density of a substance is defined as mass of the substance per unit volume. Metals have high densities. Thus, metals are heavy substances. However, aluminium, sodium and potassium have low densities. Densities of some metals are given in Table 2.1.

DO YOU KNOW?

The melting point of gallium and caesium metals are so low that they start melting at temperatures greater than 27.76°C.

Table 2.1: Melting Points and Densities of Metals

Metals	Melting points	Densities (at room temperature)	
Iron	1538°C	Iron	7.8 g/cm^3
Copper	1084°C	Copper	8.9 g/cm^3
Aluminium	660°C	Aluminium	2.7 g/cm^3
Sodium	98°C	Sodium	0.9 g/cm^3
Potassium	64°C	Potassium	0.8 g/cm^3
Gallium	30°C	Gallium	5.9 g/cm^3
Caesium	28°C	Caesium	1.9 g/cm^3

9. **Metals are sonorous:** Metals are sonorous means that they are capable of producing a ringing sound. The property of metals of being sonorous is called sonority. It is due to their ability to produce sound when hit. Metals are used for making bells and wires of violin.

African bells

Violin

Figure 2.7: Sonorous metals.

EXERCISE 2.6

1. Which metal is yellow in colour?
2. Write two uses of metals.
3. Define malleability.
4. Why are metals good conductors of electricity?
5. Metals have high

    (*a*) melting points (*b*) densities

    (*c*) both (*a*) and (*b*) (*d*) None of these.

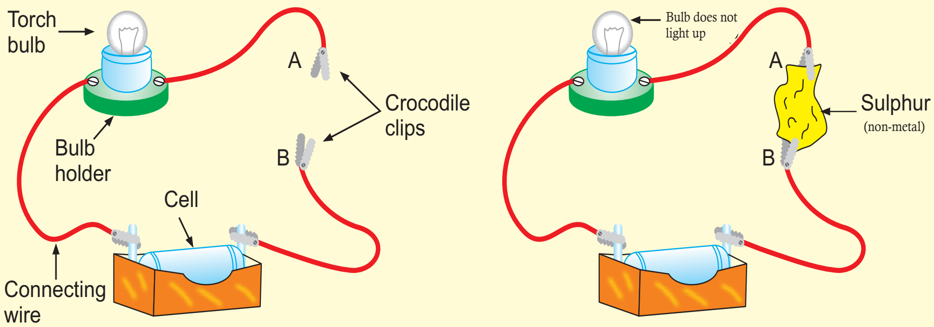
2.3 PHYSICAL PROPERTIES OF NON-METALS

ACTIVITY 2.9: Showing that a Non-metal does not Conduct Electricity

We take a dry cell, a torch bulb fitted in a holder and some connecting wires (copper wires) with crocodile clips, and connect them as shown in Figure (a) to make an electric circuit. There is a gap between the ends of crocodile clips *A* and *B* so no current flows in the open circuit shown in figure (a). Let us now insert

(i) a piece of sulphur (which is a non-metal) between the crocodile clips *A* and *B* as shown in figure (b)

(ii) Piece of dry wood between the ends of crocodile clips *A* and *B* replace sulphur in figure (b).



(*a*) (*b*)

Explanation

In both cases, we see that the bulbs do not light up at all. This means that both sulphur and piece of dry wood not allow electric current to pass through them and no current flows in the circuit. This activity shows that non-metals do not conduct electricity.

Non-metals exist in all three physical states: solid, liquid and gaseous. For example, carbon, sulphur and phosphorus are solid. Bromine is a liquid. Oxygen, hydrogen and nitrogen are gases. Non-metals have many different colours.

For example, phosphorus is red, black and white. Sulphur is yellow, chlorine is green.

Some non-metals such as oxygen and hydrogen are colourless.

The other important physical properties of non-metals are:

1. **Non-metals are not lustrous:** Non-metals do not have a shining surface. The only non-metals having a shining surface is iodine.
2. **Non-metals are neither hard nor strong:** Most of the solid non-metals are soft, they can be broken easily. For example, sulphur and phosphorus. Only one non-metal carbon in the form of diamond is very hard. Diamond is the hardest substance known on earth.
3. **Non-metals are neither malleable nor ductile:** Non-metals are brittle which means that they break into pieces when hammered or stretched.

   Therefore, non-metals cannot be hammered with a hammer to form thin sheets. They cannot be stretched to form wires.

   The property of breaking easily is called **brittleness**. Brittleness is the characteristic property of non-metals.

   Note: Brittleness is not applicable to liquid and gaseous non-metals.

4. **Non-metals do not conduct heat and electricity:** Non-metals do not conduct heat and electricity because they have no free electrons which are necessary to conduct heat and electricity. However, there is one exception. Carbon in the form of graphite is a good conductor of electricity. Therefore, graphite is used for making electrodes.

ACTIVITY 2.10: Illustrating Thermal Conductivity in Non-metals

Repeat **Activity 2.7** with carbon rod (taken out from used cell) and a lump of sulphur.

Activity 2.9 shows that non-metals do not conduct electricity.

In Activity 2.10, you will observe that both carbon and sulphur are poor conductors of heat.

5. **Non-metals have low melting and boiling points:** Non-metals have comparatively low melting and boiling points. Only graphite and diamond (the allotropes of carbon) have high melting points.The melting point of graphite is 3600C.
6. **Non-metals have low density:** The density of non-metals is low, that is, they are light substances.

   Densities of some non-metals are given in Table 2.2.

Table 2.2. Densities of Non-metals

Carbon	2.6 g/cm^3
Sulphur	2.0 g/cm^3
Bromine	3.1 g/cm^3
Phosphorus	1.8–2.6 g/cm^3

7. **Non-metals are non-sonorous:** Non-metals do not produce ringing sound when hit with an object.

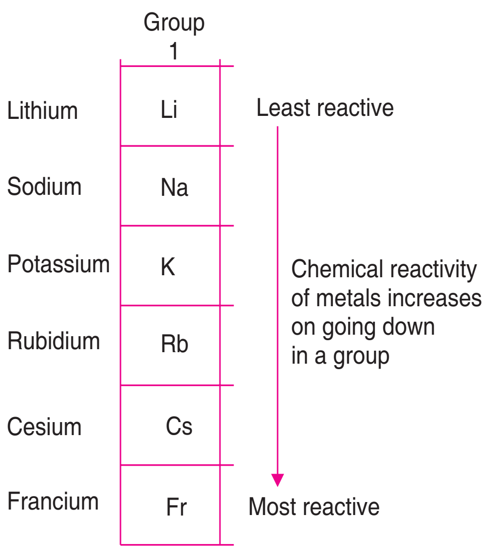
EXERCISE 2.7

1. Non-metals exist in all three physical states. (True or False)
2. ______ and ______ are colourless non-metals.
3. Which non-metal is used for making electrodes?
4. Why are non-metals bad conductors of electricity and heat?
5. Non-metals have low
   (*a*) densities
   (*b*) melting points
   (*c*) boiling points
   (*d*) All of these.

2.4 TRENDS IN REACTIVITY FOR METALS AND NON-METALS

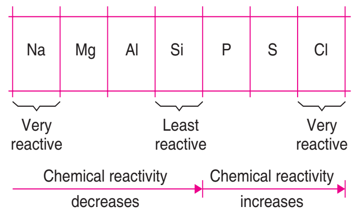
2.4.1 Reactivity of Metals

The chemical reactivity of metals increases on going down in a group of the periodic table. For example, in group 1 of alkali metals, the chemical reactivity increases from lithium to francium (radioactive).

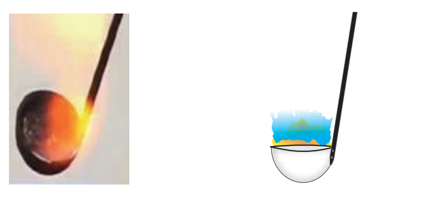


Thus, *as we go down in a group of metals, the tendency of their atoms to lose electrons increases, as the distance between the nucleus and the outermost electron increases, and hence their chemical reactivity also increases.*

On moving from left to right in a period, the chemical reactivity of elements first decreases and then increases.



In the third period of elements shown above, the reactivity of metals decreases due to the decrease of atomic size and the tendency of losing electrons, thus sodium is a very reactive element, magnesium is less reactive, whereas aluminium is still less reactive. Silicon is the chemically least reactive element in the third period. Now, phosphorus is quite reactive, sulphur is still more reactive, whereas chlorine is very reactive because the atomic size decreases and the tendency of gaining electrons increases. From this discussion we conclude that in the third period of the periodic table, chemical reactivity first decreases from sodium to silicon and then increases from phosphorus to chlorine.



(a) sodium (b) sulphur powder

Figure 2.8: Sodium reacts vigorously whereas sulphur reacts less vigorously with oxygen.

2.4.2 Reactivity of Non-metals

The chemical reactivity of non-metals decreases on going down in a group of the periodic table. For example, in group 17 of halogen elements (which are non-metals), the chemical reactivity decreases from fluorine to iodine.

	Group 17	
Fluorine	F	Most reactive
Chlorine	Cl	Chemical reactivity of non-metals decreases on going down in a group
Bromine	Br	
Iodine	I	Least reactive

Thus, *as we go down in a group of non-metals, the tendency of their atoms to gain electrons decreases, due to which their reactivity also decreases.*

EXERCISE 2.8

1. Explain, why chemical reactivity of metals increases on going down in a group of periodic table?

2. Which of the following element is least reactive?

(*a*) Sodium (*b*) Silicon (*c*) Sulphur (*d*) Chlorine

3. The chemical reactivity of non-metals decreases on going down in a group of periodic table. Why?

4. In the third period of elements shown below

Na	Mg	Al	Si	P	S	Cl

Sodium reacts vigorously but sulphur reacts less vigorously with oxygen.

(True or False)

2.5 CHEMICAL PROPERTIES OF METALS

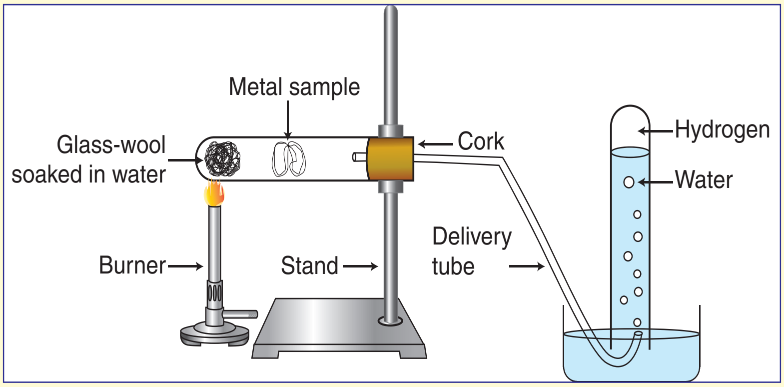
2.5.1 Reaction of Metals with Water

ACTIVITY 2.11: Illustrating Reaction of Metals with Water

Caution: Do not touch sodium and potassium with bare hands. They cause severe burns.

- Collect samples of sodium, potassium, calcium, aluminium, iron, magnesium, zinc and copper.
- Put small pieces of samples separately in beakers half filled with cold water.

- Observe which metals reacted with cold water.
  Did any metal produce fire in water?
  Did any metal start floating after sometime?
- Put the metals that do not react with cold water in beakers filled with hot water.
- Observe which metals reacted with hot water.
- For those metals which did not react with hot water arrange the apparatus as shown in figure.



Action of steam on a metal

- Observe which metals react with steam.
  Did any metal not react even with steam?
- Make an appropriate report on reaction of metals with water.

In Activity 2.11, you have observed that all metals are not equally reactive.

Metals react with water to form a metal hydroxide and hydrogen gas. Some metals react with cold water. For example, sodium, potassium and calcium. Some other metals such as magnesium react with hot water. It does not react with cold water. Metals such as aluminium, iron and zinc do not react with either cold or hot water. They react with steam* to form metal oxide and hydrogen. There are some metals that do not react even with steam. For example, copper, gold, silver and mercury.

When a metal reacts with **cold water** or **hot water**, the products formed are metal hydroxide and hydrogen gas.

Metal + Water → Metal hydroxide + Hydrogen

* Steam is a gaseous form of water. It is very hot.

When metals such as magnesium, aluminium, zinc and iron react with **steam**, the products formed are metal oxide and hydrogen gas.

Metal + Steam → Metal oxide + Hydrogen

Note: Metal oxides are basic in nature. Their solutions in water turn red litmus into blue. Some metal oxides react with water to form alkali. Some of them such as ZnO and PbO are amphoteric

Example 1

Sodium reacts vigorously with cold water to form sodium hydroxide and hydrogen gas.

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\substack{\text{Water}\\\text{(cold)}}}{2H_2O(l)} \longrightarrow \underset{\substack{\text{Sodium}\\\text{hydroxide}}}{2NaOH(aq)} + \underset{\text{Hydrogen}}{H_2(g)} + \text{Heat}$$

The reaction of sodium metal with water is also highly exothermic (heat producing). This is the reason why hydrogen gas formed during the reaction catches fire and burns causing little explosions. Thus, sodium is a very reactive metal.

Example 2

Potassium reacts violently with cold water to form potassium hydroxide and hydrogen gas.

$$\underset{\text{Potassium}}{2K(s)} + \underset{\substack{\text{Water}\\\text{(cold)}}}{2H_2O(l)} \longrightarrow \underset{\substack{\text{Potassium}\\\text{hydroxide}}}{2KOH(aq)} + \underset{\text{Hydrogen}}{H_2(g)} + \text{Heat}$$

The reaction of potassium metal with water is highly exothermic (heat producing). This is the reason why hydrogen gas formed during the reaction catches fire immediately. Thus, potassium is also a very reactive metal.

Note: Potassium is more reactive than sodium.

Example 3

Calcium reacts with cold water to form calcium hydroxide and hydrogen gas.

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\substack{\text{Water}\\\text{(Cold)}}}{2H_2O(l)} \longrightarrow \underset{\text{Calcium hydroxide}}{Ca(OH)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

The reaction of calcium with water is less violent. The heat produced is not enough for the hydrogen to catch fire. Calcium starts floating in water because bubbles of hydrogen formed during the reaction stick to surface of metal.

Note: Calcium is less reactive than sodium.

Example 4

Magnesium metal does not react with cold water. It reacts with both hot water and steam. Magnesium reacts with hot water to form magnesium hydroxide and water.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\substack{\text{Water}\\\text{(Hot)}}}{2H_2O(l)} \longrightarrow \underset{\substack{\text{Magnesium}\\\text{hydroxide}}}{Mg(OH)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

In this reaction, the piece of magnesium metal starts floating on water due to the bubbles of hydrogen gas sticking to its surface.

> **Note:** Calcium reacts with cold water but magnesium reacts only with hot water. This shows that magnesium is less reactive than calcium.

Magnesium reacts very rapidly with steam to form magnesium oxide and hydrogen.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Steam}}{H_2O(g)} \longrightarrow \underset{\text{Magnesium oxide}}{MgO(s)} + \underset{\text{Hydrogen}}{H_2(g)}$$

When magnesium reacts with hot water, it forms ***magnesium hydroxide*** and hydrogen.

In this reaction, magnesium reacts with steam to form *magnesium oxide* and hydrogen.

Example 5

(*i*) Aluminium reacts with steam to form aluminium oxide and hydrogen gas.

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\text{Steam}}{3H_2O(g)} \longrightarrow \underset{\substack{\text{Aluminium}\\\text{oxide}}}{Al_2O_3(s)} + \underset{\text{Hydrogen}}{3H_2(g)}$$

(*ii*) Zinc reacts with steam to form zinc oxide and hydrogen gas.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Steam}}{H_2O(g)} \longrightarrow \underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Hydrogen}}{H_2(g)}$$

(*iii*) Iron reacts with steam to form iron oxide and hydrogen gas.

$$\underset{\text{Iron}}{3Fe(s)} + \underset{\text{Steam}}{4H_2O(g)} \longrightarrow \underset{\text{Iron oxide}}{Fe_3O_4} + \underset{\text{Hydrogen}}{4H_2(g)}$$

Example 6

Copper does not react with water (or steam)

$$\underset{\text{Copper}}{Cu(s)} + \underset{\substack{\text{Water}\\\text{(or steam)}}}{H_2O(g)} \longrightarrow \text{No reaction}$$

2.5.2 Reaction of Metals with Acids

ACTIVITY 2.12: Illustrating Reaction of Metals with Dilute Acids

Caution: Do not touch dilute hydrochloric and sulphuric acids with your bare hands.

- Collect small pieces of magnesium, aluminium, zinc, copper and iron.
- Clean the sample metals with sand paper.
- Put these metal pieces in separate test tubes.
- Add 10 ml dilute hydrochloric acid to each test tube.
- Observe carefully the rate of formation of hydrogen gas bubbles.
  Did any metal react with dilute hydrochloric acid?
- Repeat the activity again with dilute sulphuric acid.
  Did any metal not react with dilute sulphuric acid?
- Make an appropriate report on reaction of metals with dilute acids.

In Activity 2.12, you must have observed that the rate of formation of hydrogen bubbles was the fastest in magnesium. The reactivity decreases in the order Mg > Al > Zn > Fe.

In case of copper, no bubbles were seen. This shows that copper does not react with dilute hydrochloric acid. However, it reacts with Sulfuric acid.

Metals usually react with dilute acids to give a metal salt and hydrogen gas. Some metals react violently (explosively) with dilute acids whereas some react rapidly. Sodium reacts violently with dilute acids and magnesium reacts rapidly. Some metals react slowly with dilute acids whereas a few metals do not react with acids at all. Aluminium, iron and zinc react slowly with dilute acids; whereas gold and silver do not react at all.

DO YOU KNOW?

Aqua regia, (Latin for 'royal water') is a freshly prepared mixture of concentrated hydrochloric acid and concentrated nitric acid in the ratio 3 : 1. It can dissolve gold, even though neither of these acids can do so alone. *Aqua regia* is a highly corrosive, fuming liquid. It is one of the few reagents that is able to dissolve gold and platinum.

When a metal reacts with dilute hydrochloric acid, the products formed are metal chlorides and hydrogen gas.

$$\text{Metal} + \underset{\text{(dilute)}}{\text{Hydrochloric acid}} \longrightarrow \underset{\text{(Salt)}}{\text{Metal chloride}} + \text{Hydrogen gas}$$

When a metal reacts with dilute sulphuric acid, the products formed are metal sulphate and hydrogen.

$$\textbf{Metal} + \underset{\textbf{(dilute)}}{\textbf{Sulphuric acid}} \longrightarrow \underset{\textbf{(Salt)}}{\textbf{Metal sulphate}} + \textbf{Hydrogen gas}$$

Example 1

(*i*) Sodium metal reacts violently with dilute hydrochloric acid to give sodium chloride and hydrogen gas

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\substack{\text{Dilute}\\\text{Hydrochloric acid}}}{2HCl(aq)} \longrightarrow \underset{\substack{\text{Sodium}\\\text{chloride}}}{2NaCl\,(aq)} + \underset{\text{Hydrogen}}{H_2\,(g)}$$

(*ii*) Sodium metal reacts with dilute sulphuric acid to give sodium sulphate and hydrogen gas.

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\substack{\text{Dilute sulphuric}\\\text{acid}}}{H_2SO_4(aq)} \longrightarrow \underset{\substack{\text{Sodium}\\\text{sulphate}}}{Na_2SO_4(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Example 2

(*i*) Magnesium reacts with dilute hydrochloric acid to give magnesium chloride and hydrogen gas.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\substack{\text{Dilute hydrochloric}\\\text{acid}}}{2HCl\,(aq)} \longrightarrow \underset{\substack{\text{Magnesium}\\\text{chloride}}}{MgCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

(*ii*) Magnesium reacts with dilute sulphuric acid to form magnesium sulphate and hydrogen gas.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\substack{\text{Dilute sulphuric}\\\text{acid}}}{H_2SO_4(aq)} \longrightarrow \underset{\substack{\text{Magnesium}\\\text{sulphate}}}{MgSO_4(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Example 3

(*i*) Calcium reacts with dilute hydrochloric acid to form calcium chloride and hydrogen gas.

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\substack{\text{Hydrochloric}\\\text{acid (dilute)}}}{2HCl(aq)} \longrightarrow \underset{\substack{\text{Calcium}\\\text{chloride}}}{CaCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

(*ii*) Calcium reacts with dilute sulphuric acid to give calcium sulphate and hydrogen gas.

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\substack{\text{Sulphuric}\\\text{acid (dilute)}}}{H_2SO_4(aq)} \longrightarrow \underset{\substack{\text{Calcium}\\\text{sulphate}}}{CaSO_4(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Example 4

(*i*) Aluminium reacts with dilute hydrochloric acid to give aluminium chloride and hydrogen gas

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\substack{\text{Hydrochloric}\\ \text{acid (dilute)}}}{6HCl(aq)} \longrightarrow \underset{\substack{\text{Aluminium}\\ \text{chloride}}}{2AlCl_3(aq)} + \underset{\text{Hydrogen}}{3H_2(g)}$$

(*ii*) Aluminium reacts with dilute sulphuric acid to form aluminium sulphate and hydrogen gas.

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\substack{\text{Sulphuric}\\ \text{acid (dilute)}}}{3H_2SO_4(aq)} \longrightarrow \underset{\substack{\text{Aluminium}\\ \text{sulphate}}}{Al_2(SO_4)_3(aq)} + \underset{\text{Hydrogen}}{3H_2(g)}$$

Example 5

(*i*) Zinc reacts with dilute hydrochloric acid to form zinc chloride and hydrogen gas.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\substack{\text{Hydrochloric}\\ \text{acid (dilute)}}}{2HCl(aq)} \longrightarrow \underset{\substack{\text{Zinc}\\ \text{chloride}}}{ZnCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

(*ii*) Zinc metal reacts with dilute sulphuric acid to give zinc sulphate and hydrogen gas.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\substack{\text{Sulphuric}\\ \text{acid (dilute)}}}{H_2SO_4(aq)} \longrightarrow \underset{\text{Zinc sulphate}}{ZnSO_4(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Example 6

(*i*) Iron reacts slowly with cold dilute hydrochloric acid to form iron chloride and hydrogen gas.

$$\underset{\text{Iron}}{Fe(s)} + \underset{\substack{\text{Hydrogen}\\ \text{chloride (dilute)}}}{2HCl(aq)} \longrightarrow \underset{\substack{\text{Iron}\\ \text{chloride}}}{FeCl_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

(*ii*) Iron reacts with sulphuric acid (dilute) to give iron sulphate and hydrogen gas.

$$\underset{\text{Iron}}{Fe(s)} + \underset{\substack{\text{Sulphuric}\\ \text{acid (dilute)}}}{H_2SO_4(aq)} \longrightarrow \underset{\text{Iron (II) sulphate}}{FeSO_4(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Example 7

(*i*) Copper does not react with dilute hydrochloric acid.

$$Cu(s) + HCl(aq) \longrightarrow \text{No reaction}$$

(*ii*) Copper does not react with dilute sulphuric acid.

$$Cu(s) + H_2SO_4(aq) \longrightarrow \text{No reaction}$$

Hydrogen gas is not evolved when a metal (e.g., Cu) reacts with nitric acid (HNO_3).

It is because nitric acid is a strong oxidising agent. It oxidises the hydrogen produced to water and nitric acid itself is reduced to any of the nitrogen oxides. The examples of nitrogen oxides are nitrogen monoxide (NO), nitrogen dioxide (NO_2) and dinitrogen monoxide (N_2O). Only magnesium and manganese react with very dilute nitric acid to evolve hydrogen gas.

The reaction of magnesium and manganese metals with very dilute nitric acid are:

- Magnesium reacts with very dilute nitric acid to form magnesium nitrate and hydrogen gas

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\substack{\text{Nitric acid}\\ \text{(very dilute)}}}{2HNO_3(aq)} \longrightarrow \underset{\substack{\text{Magnesium}\\ \text{nitrate}}}{Mg(NO_3)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

2.5.3 Reaction of Metals with Halogens

The elements of group 17 in the periodic table are called **halogens**. Fluorine (F), Chlorine (Cl), bromine (Br), Iodine (I) and Astatine (At) are halogens.

Metals react with halogens to form ionic halide.

Metal halides are usually solid and conduct electricity in solution. Let us see some equations for reaction of metals with chlorine.

(*i*) Sodium reacts with chlorine to form sodium chloride

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{Sodium chloride}}{2NaCl(s)}$$

(*ii*) Calcium reacts with chlorine to form calcium chloride

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{Calcium chloride}}{CaCl_2(s)}$$

(*iii*) Aluminium reacts with chlorine to form aluminium chloride.

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\text{Chlorine}}{3Cl_2(g)} \longrightarrow \underset{\substack{\text{Aluminium}\\ \text{chloride}}}{2AlCl_3(s)}$$

(*iv*) Iron reacts with chlorine to form iron chloride.

$$\underset{\text{Iron}}{2Fe(s)} + \underset{\text{Chlorine}}{3Cl_2(g)} \longrightarrow \underset{\substack{\text{Iron III}\\ \text{chloride}}}{2FeCl_3(s)}$$

(*v*) Copper reacts with chlorine to form copper chloride.

$$\underset{\text{Copper}}{Cu(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{Copper (II) chloride}}{CuCl_2(s)}$$

(*vi*) Zinc reacts with chlorine to form zinc chloride.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2(s)}$$

Note: All the metal chlorides are ionic compounds.

2.5.4 Reaction of Metals with Oxygen

ACTIVITY 2.13: Illustrating Reaction of Metals with Oxygen

Safety: The activity needs the teacher's assistance. Students should wear eye protection.

Collect small pieces of potassium, sodium, magnesium, aluminium, zinc, copper and iron. Also collect some iron filings.

- Hold any of the samples taken above with a pair of tongs and try burning over a flame. Repeat with other metal samples.
- Collect the product if formed.
- Let the products and the metal surface cool down.

  Which metals burn easily?

  How does the metal surface appear after burning?

  Are the products soluble in water?
- Make an appropriate report on reaction of metals with oxygen.

You will observe in Activity 2.13 that almost all metals combine with oxygen to form metal oxides, but all metals do not react with oxygen at the same rate. Different metals show different reactivity towards oxygen.

Metals such as potassium and sodium react so vigorously that they catch fire if kept in the open. Hence, to protect them and to prevent accidental fires, they are kept immersed in kerosene oil. At ordinary temperature, the surfaces of metals such as magnesium, aluminium, zinc, lead, etc., are covered with a thin layer of oxide. The protective oxide layer prevents the metal from further oxidation. Iron does not burn on heating but iron filings burn vigorously when sprinkled in the flame of the burner. Copper does not burn, but the hot metal is coated with a black coloured layer of copper(II) oxide. Silver and gold do not react with oxygen even at high temperatures.

Let us see some equations for reaction of metals with oxygen.

(*i*) Sodium reacts with oxygen to form sodium oxide.

$$\underset{\text{Sodium}}{4Na(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Sodium oxide}}{2Na_2O(s)}$$

(*ii*) Potassium reacts with oxygen to form potassium oxide.

$$4K(s) + O_2(g) \longrightarrow 2K_2O(s)$$

(*iii*) Magnesium reacts with oxygen to form magnesium oxide

$$\underset{\text{Magnesium}}{2Mg(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Magnesium oxide}}{2MgO(s)}$$

EXERCISE 2.9

1. ______ is a gaseous form of water.
2. Metals react with water to form metal oxide and hydrogen gas. (True or False)
3. Metals react with steam to form metal hydroxide and hydrogen gas. (True or False)
4. Complete and balance the following equations:
   (*i*) $Na(s) + H_2O(l) \longrightarrow ?$
   (*ii*) $Mg(s) + HCl(aq) \longrightarrow ?$
   (*iii*) $Cu(s) + H_2SO_4(aq) \longrightarrow ?$
   (*iv*) $Al(s) + Cl_2(g) \longrightarrow ?$
   (*v*) $K(s) + O_2(g) \longrightarrow ?$
5. Gold and Silver ______ react with dilute acids.
6. Hydrogen gas is not evolved when a metal reacts with ______ .
7. All metal chlorides are ionic in nature. (True or False)

2.6 CHEMICAL PROPERTIES OF NON-METALS

Non-metals neither react with water nor with dilute acids. In other words, non-metals do not displace hydrogen gas from acids and water. Some exceptions are fluorine and chlorine

2.6.1 Reaction of Non-metals with Halogen (Chlorine)

Non-metals react with chlorine to form covalent chlorides. Non-metal chlorides are usually liquids or gases. They do not conduct electricity.

Example 1

(*i*) $$\underset{\text{Carbon}}{C(s)} + \underset{\text{Chlorine}}{2Cl_2(g)} \longrightarrow \underset{\text{Carbon tetrachloride}}{CCl_4(g)}$$

(*ii*) $$\underset{\text{Phosphorus}}{P_4(s)} + \underset{\text{Chlorine}}{6Cl_2(g)} \longrightarrow \underset{\text{Phosphorus chloride (Covalent chloride)}}{4PCl_3(l)}$$

2.6.2 Reaction of Non-metals with Oxygen

Non-metals react with oxygen to form acidic oxides or neutral oxides.

(*i*) Carbon reacts with oxygen to form carbon dioxide

$$\underset{\text{Carbon}}{C(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Carbon dioxide (Acidic oxide)}}{CO_2(g)}$$

(*ii*) Hydrogen reacts with oxygen to form water.

$$\underset{\text{Hydrogen}}{2H_2(g)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Water (Neutral oxide)}}{2H_2O(l)}$$

EXERCISE 2.10

1. Non-metals do not displace hydrogen gas from acids. (True or False)
2. Non-metals react with ______ to form covalent chlorides.
3. Non-metals react with oxygen to form
   (*a*) acidic oxides
   (*b*) neutral oxides
   (*c*) both (*a*) and (*b*)
   (*d*) none of these
4. Complete the following equations:
   (*i*) $H_2(g) + O_2(g) \longrightarrow ?$
   (*ii*) $C(s) + Cl_2(g) \longrightarrow ?$
5. Non-metal chlorides do not conduct electricity. (True or False)

2.7 COMPARISON AMONG THE PHYSICAL AND CHEMICAL PROPERTIES OF METALS AND NON-METALS

We have studied the physical and chemical properties of metals and non-metals. Let us see the main points of difference between the metals and non-metals.

Table 2.3: Differences between Physical Properties of Metals and Non-metals

Metals	Non-metals
1. Metals are malleable and ductile. That is, metals can be hammered into thin sheets and drawn into thin wires.	Non-metals are brittle (break easily). They are neither malleable nor ductile.
2. Metals are good conductors of heat and electricity.	Non-metals are bad conductors of heat and electricity (except *graphite* which is a good conductor of electricity).
3. Metals are lustrous (shiny) and can be polished.	Non-metals are non-lustrous (dull) and cannot be polished (except *iodine* which is a lustrous non-metal).
4. Metals are solid at room temperature (except *mercury* which is a liquid metal).	Non-metals may be solid, liquid or gaseous at room temperature.
5. Metals are strong and tough.	Non-metals are neither strong nor tough.

Table 2.4: Differences between Chemical Properties of Metals and Non-metals

Metals	Non-metals
1. Metals form basic or amphoteric oxides.	Non-metals form acidic oxides or neutral oxides.
2. Metals displace hydrogen from water (or steam).	Non-metals do not react with water (or steam) and hence do not displace hydrogen from water (or steam).
3. Metals displace hydrogen from dilute acids.	Non-metals do not react with dilute acids and hence do not displace hydrogen from dilute acids.
4. Metals form ionic chlorides with chlorine.	Non-metals form covalent chlorides with chlorine.
5. Metals usually do not combine with hydrogen. Only a few reactive metals combine with hydrogen to form ionic metal hydrides.	Non-metals react with hydrogen to form stable, covalent hydrides.

2.8 USES OF METALS AND NON-METALS

Table 2.5: Uses of Metals and Non-metals

Uses of Metals	Uses of Non-metals
Metals like ***copper*** and ***aluminium*** are used for making electric wires and cooking vessels.	***Hydrogen*** and ***carbon*** are the essential constituents of the cells of all living beings. Without these elements, life is not possible. Furthermore, carbon in the form of coal is used for generating heat energy for houses as well as factories.
Iron and ***steel*** are extensively used for making machinery, automobiles, trains, boilers, bridges and buildings. ***Stainless steel*** is used for making cooking vessels and surgical instruments.	***Nitrogen*** in the form of fertilisers is essential for the growth and development of plants. ***Urea*** and ***calcium ammonium nitrate*** are common fertilisers used by the farmers.
Aluminium is extensively used for making airplanes and automobile engines.	***Oxygen*** is absolutely essential for the respiration of all kinds of plants and animals. It is also used in the manufacture of steel and for artificial respiration on high altitudes, in space and deep sea diving.
Sodium is used in nuclear reactors.	***Silicon*** is used for making microchips used in making all kinds of electronic goods.
Tin is used for coating iron containers for packaging food.	***Phosphorus*** is used in match industry and fertilisers.
Mercury is used in thermometers.	***Sulphur*** is used for making fire crackers, gun powder and sulphuric acid.
Steel is used in making all kinds of weapons, ships, tanks, guns, etc.	***Chlorine*** is used in the disinfection of drinking water.
Gold and ***silver*** are used for making ornaments.	***Iodine*** dissolved in alcohol [tincture iodine] is antiseptic in nature and is used in dressing wounds. ***Helium*** is used in balloons. Argon is used electrical bulbs and fluorescent tubes.

EXERCISE 2.11

1. State True or False
   (*a*) Mercury is used in thermometers.
   (*b*) Gold and Silver are used for making electric wires.
   (*c*) Helium is used in balloons.
   (*d*) Phosphorus is used in fertilisers.
   (*e*) Chlorine is not used in the disinfection of drinking water.
2. Gap filling:
   (*a*) Sulphur is used for making ______.
   (*b*) Silicon is used for making ______.
   (*c*) Sodium is used in ______ reactors.
3. Name the Noble gas used in electric bulbs.
4. Which element is used for coating iron containers for packaging food?
5. Name two metals used for making machines.

2.9 SUMMARY

- Elements can be classified as metals, metalloids and non-metals.
- On moving from left to right in a period, the metallic character of elements decreases whereas non-metallic character increases.
- On going down in a group of the periodic table, the metallic character of elements increases whereas non-metallic character decreases.
- Metals are lustrous, malleable, ductile and good conductors of heat and electricity. They are solids at room temperature, except mercury which is a liquid.
- Non-metals have properties opposite to that of metals. They are neither malleable nor ductile. They are bad conductors of heat and electricity, except for graphite, which conducts electricity.
- Metals form positive ions by losing electrons.
- Non-metals form negative ions by accepting electrons.
- Elements in-between the metals and non-metals are known as metalloids.
- The chemical reactivity of metals increases on going down in a group of the periodic table.
- The chemical reactivity of non-metals decreases on going down in a group of the periodic table.
- Metals react with water to form metal hydroxide or metal oxide and hydrogen gas.
- Metals usually react with dilute acids to give a metal salt and hydrogen gas.

Metal + Hydrochloric acid → Metal chloride + Hydrogen

Metal + Sulphuric acid → Metal sulphate + Hydrogen.

- Copper does not react with dilute acids.
- Metals react with halogen to give ionic halides.
- Metals react with oxygen to form metal oxides (basic or amphoteric oxides).
- Different metals have different reactivities with water and dilute acids.
- Non-metals react with chlorine to form covalent chlorides.
- Non-metals react with oxygen to give acidic oxides or neutral oxides.

2.10 GLOSSARY

- **Antiseptic:** a substance that prevents the growth of disease-causing micro-organisms.
- **Carbohydrates:** a large group of organic compounds occurring in foods including sugars, starch, and cellulose.
- **Fertilisers:** a chemical or natural substance added to soil or land to increase its fertility.
- **Fluorescent:** emitting light.
- **Fungicides:** a chemical that destroys fungus.
- **Metal:** a material that is hard, opaque, shiny and has good electrical and heating conductivity.
- **Metalloid:** a chemical element with properties in between those of metals and non-metals.
- **Non-metal:** a material that does not have properties of metals.
- **Nuclear reactor:** a device used at nuclear power plants for electricity generation.
- **Ornaments:** a thing used or serving to make something look more attractive.
- **Thermometer:** an instrument for measuring temperature.
- **Weapon:** something designed or used for inflicting physical injury or damage.

2.11 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Which of the following metal exists in the liquid state?
(*a*) Sodium (*b*) Silver (*c*) Mercury (*d*) Neon

2. On going down in a group of the periodic table, the metallic character of elements ....
(*a*) increases (*b*) decreases (*c*) both (*a*) and (*b*) (*d*) neither (*a*) nor (*b*)

3. Silicon and germanium are examples of ....
(*a*) metals (*b*) metalloids (*c*) non-metals (*d*) none of these

4. All non-metals are solids and gases except:
(*a*) fluorine (*b*) chlorine (*c*) bromine (*d*) iodine

5. Which metal is the best conductor of electricity?
(*a*) Silver (*b*) Copper (*c*) Aluminium (*d*) Iron

6. Which of the following statement(s) is/are not correct.
   (*a*) The hardest substance known on earth is diamond.
   (*b*) Graphite – a non-metal – conducts electricity.
   (*c*) The melting points of gallium and caesium are very high.
   (*d*) Brittleness is the characteristic property of non-metal.
7. On moving from left to right in a period, the chemical reactivity of elements ....
   (*a*) increases (*b*) decreases
   (*c*) first decreases and then increases (*d*) first increases and then decreases
8. Name the gas produced when metals react with dilute acids.
   (*a*) Oxygen (*b*) Hydrogen (*c*) Nitrogen (*d*) Chlorine
9. The molecular formula of manganese nitrate is
   (*a*) $Mn\,(NO_3)_2$ (*b*) $Mn_2\,(NO_3)$ (*c*) $Mn_3\,(NO)_2$ (*d*) MnNO
10. Carbon reacts with oxygen to give
   (*a*) carbon dioxide (*b*) carbon monoxide
   (*c*) both (*a*) and (*b*) (*d*) none of these

II. Open Ended Questions

1. What is meant by saying that the metals are malleable and ductile? Explain with examples.
2. With the help of example, describe how metals differ from non-metals.
3. Name one metal and one non-metal which exist in liquid state at room temperature.
4. (*a*) Name the most abundant metal in the earth's crust.
   (*b*) Name the most abundant non-metal in the earth's crust.
   (*c*) Name one metal which has low melting point.
   (*d*) Name one non-metal which is kept under water.
   (*e*) Name one metal which is stored in kerosene oil.
5. Complete and balance the following equations:
   (*i*) $K(s) + H_2O\,(l) \rightarrow ? + ?$ (*ii*) $Mg\,(s) + H_2O\,(g) \rightarrow ? + ?$
   (*iii*) $Mg\,(s) + H_2O\,(l) \rightarrow ? + ?$ (*iv*) $Ca\,(s) + H_2SO_4\,(aq) \rightarrow ? + ?$
   (*v*) $Fe\,(s) + Cl_2\,(g) \rightarrow ?$ (*vi*) $P_4\,(s) + ? \rightarrow PCl_3\,(l)$
6. Describe the trends and patterns in the properties of elements in groups and periods.
7. Describe electrical conductivity of metals and non-metals.
8. Distinguish between the chemical properties of metals and non-metals.
9. Can you explain why copper does not react with water?
10. Give three uses of non-metals in daily life.

III. Practical-based Questions

1. Which of the following is an example of non-metals?

(*a*) 

(*b*) 

(*c*) 

(*d*) 

2. Which is the most reactive metal in the given table?

Li
Na
K
Rb

(*a*) Li (*b*) Na (*c*) K (*d*) Rb

3. Which is the least reactive metal in the given table?

Na	Mg	Al	Si	P	S	Cl

(*a*) Na (*b*) Cl (*c*) Si (*d*) P

4. How many non-metals are there in the following figure?

1 **H**							2 **He**
3 **Li**	4 **Be**	5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**	13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**

(*a*) Five (*b*) Eleven (*c*) Six (*d*) Seven

5. In the following periodic table, the yellow colour represents ....

Periodic Table of the Elements

I	II											III	IV	V	VI	VII	VIII
1	2											13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.811	6 **C** 12.011	7 **N** 14.007	8 **O** 15.999	9 **F** 18.998	10 **Na** 20.180
11 **Na** 22.990	12 **Mg** 24.305	3	4	5	6	7	8	9	10	11	12	13 **Al** 26.982	14 **Si** 28.086	15 **P** 30.974	16 **S** 32.066	17 **Cl** 35.453	18 **Ar** 39.940
19 **K** 39.098	20 **Ca** 40.078	21 **Sc** 44.956	22 **Ti** 47.88	23 **V** 50.942	24 **Cr** 51.996	25 **Mn** 54.938	26 **Fe** 55.847	27 **Co** 58.933	28 **Ni** 58.69	29 **Cu** 63546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.61	33 **As** 72.61	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
37 **Rb** 85.468	38 **Sr** 87.62	39 **Y** 88.906	40 **Zr** 91.224	41 **Nb** 92.906	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.905	46 **Pd** 106.42	47 **Ag** 107.868	106 **Cd** 114.411	49 **In** 114.82	50 **Sn** 118.710	51 **Sb** 121.757	52 **Te** 121.757	53 **I** 125.905	54 **Xe** 131.29
55 **Cs** 132.905	56 **Ba** 137.327	71 **Lu** 174.967	72 **Hf** 178.49	73 **Ta** 180.948	74 **W** 183.85	75 **Rs** 186.207	76 **Os** 190.2	106 **Ir** 192.22	77 **Pt** 195.08	79 **Au** 196.967	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.2	83 **Bi** 208.980	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
1 **Fr** (223)	38 **Ra** 226.025	88 **Lr** 226.025	104 **Rf** (261)	10.5 **Dh** (262)	106 **Sg** (263)	108 **Bh** (265)	106 **Hs** (263)	109 **Mt** (268)	110 (269)	11 273							

57 **La** 138.906	58 **Ce** 140.908	59 **Pr** 140.908	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.965	64 **Gd** 157.25	65 **Tb** 162.50	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 167.934	70 **Yb** 173.04
89 **La** 227.028	90 **Th** 232.038	91 **Pa** 232.086	92 **U** 238.029	93 **Np** 237.048	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **BK** (247)	98 **Cf** (251)	99 **Fm** (2.57)	99 **Fm** (2.57)	101 **Md** (2.57)	102 **Na** (2.99)

(*a*) Metals (*b*) Non-metals (*c*) Metalloids (*d*) Liquid metals

Unit 3

Water Pollution

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- define water pollution.
- identify the main water pollutants.
- describe the dangers of polluted water.
- suggest the ways of preventing water pollution.

KNOWLEDGE GAIN

Tap water which is considered safe for drinking sometimes includes harmful microbes. These microbes do not alter colour and odour of water but are very harmful.

3.1 WATER POLLUTION

ACTIVITY 3.1: Showing Awareness about Water Pollution

Investigate the level of awareness about water pollution in your area. Collect data on the sources of drinking water and polluted water from newspaper and magazines.

What are the common water-borne diseases in the community? You can consult your local doctor/health worker for this.

Which are the governmental and non-governmental organisations working in this field? What are the measures being taken by them for generating awareness?

Prepare an illustrative presentation on "water pollution" from the data collected. Present it in class.

Water is essential for life. Without water there would be no life. The purity of water is always taken for granted. Most of the water which we use comes from rivers and lakes.

Everyday, many unwanted and harmful substances are thrown (or discharged) into the rivers and lakes. They make the water of rivers and lakes impure (or contaminated). So, we say that the water has been polluted. *The contamination of water of rivers, lakes and ponds, etc., with unwanted and harmful substances is called water pollution*. Water is said to be polluted when it becomes unfit for drinking or bathing. Pollution of water originates from human activities and sometimes from nature (animals and plants). Through different paths, pollution reaches surface or ground water. Easily identified source or place of pollution is called *point source*. For example, municipal and industrial discharge pipes where pollutants enter the water-source. *Non-point sources* of pollution are those where a source of pollution cannot be easily identified. For example, agricultural run off (from farm, animals and crop-lands), acid rain, storm-water drainage (from streets, parking lots and lawns), etc.

3.1.1 Causes of Water Pollution

(i) Pathogens: The most serious water pollutants are the disease causing agents called pathogens. Pathogens include bacteria and other organisms that enter water from domestic sewage and animal excreta. Human excreta contains bacteria such as *Escherichia coli* and *Streptococcus faecalis* which cause gastrointestinal diseases.

(ii) Organic wastes: The other major water pollutant is organic matter such as leaves, grass, trash, etc. They pollute water as a consequence of run off. Excessive phytoplankton growth within water is also a cause of water pollution. These wastes are biodegradable.

Figure 3.1: How water is contaminated.

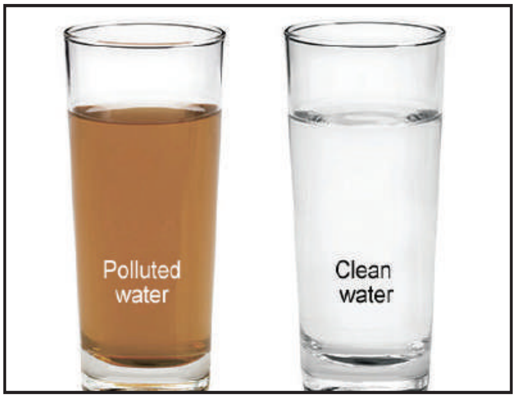


Figure 3.2: Polluted water and clean water.

ACTIVITY 3.2: Illustrating Effects of Quantities

Pour a cup of black ink into a river. What do you observe? Are you able to see it? Now do the same in a bucket. Do you find any change?

When you poured a cup of black ink into a river, the ink quickly disappeared into the river's much larger volume of clean water. The ink would still be there in the river, but in such a low concentration that you would not be able to see it. At such low levels, the chemicals in the ink probably would not present any real problem. However, when you poured a cup of ink into a bucket, the bucket quickly turned black. The chemicals in the ink could very quickly have an effect on the quality of the water.

Thus, water pollution is all about quantities of pollutants. It depends on how much of a polluting substance is released and how big a volume of water it is released into.

EXERCISE 3.1

1. Most of the water which we use comes from ______ and ______ .
2. What do you mean by water pollution?
3. Name two pathogens which cause gastrointestinal disease.
4. Pollution of water originates from human activities. (True or False)
5. What are the causes of water pollution?

3.2 MAIN WATER POLLUTANTS

ACTIVITY 3.3: Illustrating Major Pollutants of Water

Visit your nearby water body. Collect information on the major pollutants added to the water body. Illustrate with pictures the main idea of water pollution in the water body.

The substances which cause water pollution are known as water pollutants. Most water pollution does not begin in the water itself. For example, in oceans around 80% of pollution enters from the land. The main water pollutants are:

3.2.1 Sewage

ACTIVITY 3.4: Illustrating Disposal of Sewage

Visit in groups, the sewage disposal system of your locality. Try to find out the answer to the following questions:

- How is sewage collected from your home?
- Where does it go thereafter?

Sewage is a water carrying waste. It generally consists of faeces, urine and laundry waste. It also contains harmful micro-organisms such as bacteria, protozoa, fungi, viruses and parasites.

Figure 3.3: Sewage.

Disposal of sewage is a big problem in developing countries. Most of the people do not have access to proper sanitation facilities. It affects people's immediate environment causing various water-borne diseases such as diarrhoea. Even if there are flush toilets, the problem still continues. When you flush the toilet, the waste has to go somewhere. Even after it leaves sewage treatment works, there is still waste to dispose off. Sometimes sewage waste is pumped untreated into the sea.

EXERCISE 3.2

1. Sewage generally consists of ______, ______ and ______.
2. Disposal of sewage is not a problem in developing countries. (True or False)
3. Name three harmful micro-organisms sewage contains.

3.2.2 Nutrient-rich Waste Water

The farmers use large amounts of fertilisers in the fields to increase the crop yields. These are rich in nitrates and phosphates. The excess fertilisers dissolve in water and run into rivers, lakes and ponds. Fertilisers are plant nutrients. They cause rapid growth of tiny, green, water plants called algae in the water body. Algae cover the entire water body like a green sheet. Algae compete with other organisms in the water for dissolved oxygen. As a result, there is a threat to the aquatic life.

Figure 3.4: Nutrient-rich waste water.

EXERCISE 3.3

1. Nutrient-rich waste water causes rapid growth of green plants called ______ in the water body.
2. Is nutrient-rich waste water a threat to aquatic life?
3. Name the organism which covers the entire water body like a green sheet.

3.2.3 Chemical Waste

Almost all the industries produce poisonous chemicals as their waste products. These are called **chemical waste** or **industrial wastes**. These wastes are discharged untreated in nearby water bodies. In this way, the water bodies get polluted with chemicals. The chemicals present are the compounds of harmful metals such as mercury, cadmium, lead, arsenic and nickel. These may also include detergents and polychlorinated biphenyls (PCBs). These chemicals can kill aquatic animals and plants. They also cause severe disorders in humans such as cancer and nervous disorders.

Figure 3.5: Industrial waste.

EXERCISE 3.4

1. What do you mean by chemical waste?
2. Chemical waste can cause ______ and ______ in humans.
3. Compounds of which elements are present in chemical waste?

3.2.4 Radioactive Waste

ACTIVITY 3.5: Showing Radioactive Waste

Look at the figure. Observe the sign on the dustbin. Answer the questions raised.

Describe the picture. Make a report and submit it to your teacher.

Radioactive waste is a waste that contains radioactive substance. A radioactive substance is unstable and produces dangerous kinds of radiation. People view radioactive waste with great alarm – and for good reason. At high concentration it can kill, at lower concentrations it causes diseases like cancer. They are carried into water from nuclear power plants, wastes of uranium and thorium during their mining and refining processes and also from medical and scientific institutions.

EXERCISE 3.5

1. What do you mean by radioactive substance?
2. Radioactive waste are generated from ______ and _______ .
3. How are radioactive wastes harmful?

3.2.5 Oil Pollution

Oil and oil wastes enter water bodies from different sources such as oil refineries, storage tanks, automobile waste oil, and industries. Spillage of oil from ships also results in pollution. The pollution caused by oil and oil wastes is termed as oil pollution. Oil is insoluble in water; it floats and spreads rapidly into a thin layer. This layer prevents oxygen transfer from atmosphere. As a result of this, less oxygen is available for aquatic life.

Figure 3.6: (a) Spillage of oil.

Figure 3.6: (b) Oil pollution affecting aquatic life.

At sea, oil layer is responsible for the death of birds. The oil penetrates the bird feathers thereby affecting their floating and flying abilities.

EXERCISE 3.6

1. From where do oil wastes enter into water bodies?
2. The pollution caused by oil and oil wastes is termed ______ .
3. How is oil pollution responsible for death of aquatic animals and plants?

3.2.6 Plastic

ACTIVITY 3.6: Say No to Plastics

Look at the banner below. Have you seen this before? What is meant by polythene? Why and when was it banned in Rwanda? How did Rwanda accomplish it? Make a report to be presented in the class.

Note: Polythene bags have been banned in Rwanda since 2008.

Polythenes or polyethylenes are the most common plastics. Plastic is far and away the most common substance that washes up with the waves. There are three reasons for this: plastic is one of the most common materials, used for packaging, and making any kind of manufactured object from clothing to automobile parts; plastic is light and floats easily. So, it can travel enormous distances across the oceans; most plastics are not biodegradable (they do not break down naturally in the environment). Once in a water body they amass in landfills, litter streets, obstruct sewers and hurt aquatic life.

Figure 3.7: (a) Polythene causes water pollution.

Figure 3.7: (b) Plastics cause water pollution.

A plastic bottle can survive an estimated 450 years in the ocean and plastic fishing line can last up to 600 years.

EXERCISE 3.7

1. ______ is the most common plastics.
2. Plastics are biodegradable. (True or False)
3. Why is use of plastics prohibited?

3.2.7 Alien Species

Alien species (sometimes known as invasive species) are animals or plants from one region that have been introduced into a different ecosystem where they do not belong. The water hyacinth which was introduced as an ornamental plant has since invaded lakes in Rwanda.

It has invaded from Muhazi to Rweru from the river Nyabarongo, and even reached Lake Victoria through Akagera river. The water hyacinth is a major biodiversity problem in the ecosystem of the Lake Victoria Basin.

Rampant growth of water hyacinth can destroy native wetlands and waterways, killing native fish and other wildlife. Water hyacinth can form dense mats that spread out across water surfaces eventually choking the entire water body. Heavy weed cover also prevents the exchange of air, which normally occurs on an open water surface. This stagnation affects water quality and may result in the death of aquatic animals.

Figure 3.8: Water hyacinth.

EXERCISE 3.8

1. What do you mean by alien species?
2. ______ is a major biodiversity problem of the Lake Victoria Basin.
3. Growth of ______ can destroy aquatic life.

3.2.8 Other Forms of Pollution

This category includes the most common forms of pollution – but by no means the only ones. Heat or Thermal pollution from factories and power plants also causes problems in the river. By adding hot water into the water body it raises the temperature. The rise in the temperature has an adverse effect on the animals and plants living in it.

Figure 3.9: Thermal pollution affecting river.

3.3 DANGERS OF POLLUTED WATER

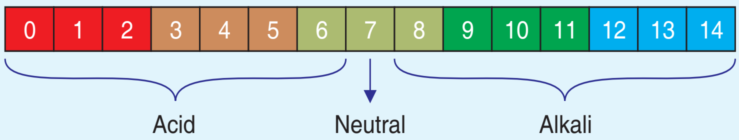
ACTIVITY 3.7: Illustrating Effects of Polluted Water

Learners plan a field visit to nearby areas and collect various samples of water. Pour each into separate glass containers. Compare the samples for smell, acidity and colour. Complete the following table in your exercise book.

	Smell	Acidity/pH	Colour
Tap water			
Polluted water			
Pond water			
River water			
Well water			
Lake water			

What is the possible cause of smell, acidity and colour in the water? What is the possible cause of this difference? Make a comparative report of your observations.

Acidity of water is measured as pH (potential of hydrogen) of the water body. It is a figure between 0 and 14 defining how acidic or basic a body of water is along the pH scale. The lower the number, the more acidic the water is. The higher the number, the more basic it is. A pH of 7 is considered neutral. The pH of pure water is 7. You can use a pH strip to measure pH. The colour of the strip after dipping in water will give its pH.



Addition of pollutants to water changes its physical, chemical and biological properties. Water from different sources is likely to have different pollutants. For example, a river situated near an industry is more likely to be affected by its discharge. Water pollution is very harmful to humans, animals and water life. The effects can be catastrophic, depending on the kind of chemicals, concentrations of the pollutants and where they are polluting. Dangers of polluted water include:

3.3.1 Eutrophication

The entry of nutrient-rich water results in a thick growth of algae (tiny plant) called algal bloom, and many other weeds. Rapid growth of these plants covers the entire surface of water. This is called **eutrophication**. Eutrophication may be defined as the process of nutrient enrichment of water bodies and the subsequent overgrowth of plants on the surface of water. The algae use up a lot of oxygen that other aquatic animals die due to lack of it. It also blocks light to reach under water affecting aquatic plants. Eutrophication hence results in loss of aquatic life. Slowly, it results in the death of "lake or river".

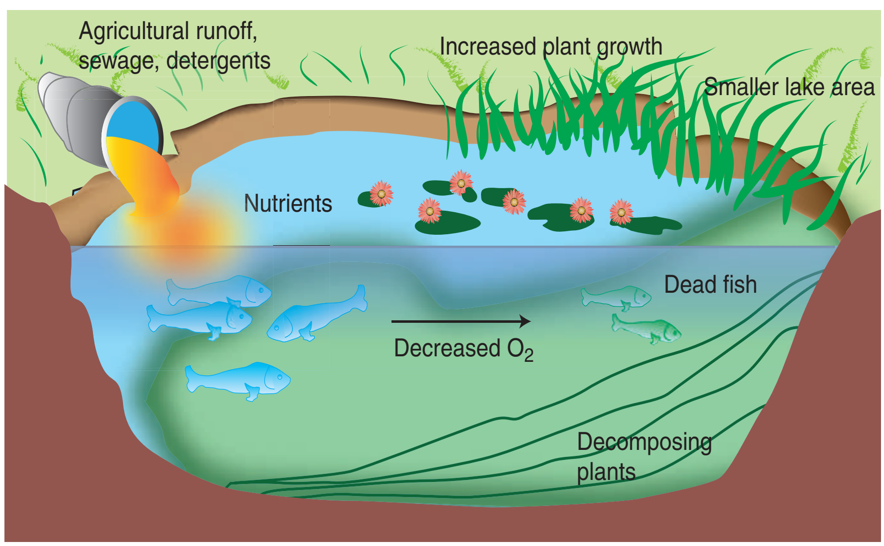


Figure 3.10: Eutrophication.

3.3.2 Acidification

The oceans are normally a natural carbon sink, absorbing carbon dioxide from the atmosphere. Carbon dioxide (CO_2) is released into the oceans as a result of water pollution by nutrients. It enhances the unwanted changes in ocean acidity due to atmospheric increases in CO_2. It impacts primarily the ecosystems and fish communities that live in the ocean. In particular, the rising levels of CO_2 acidify the ocean. Even though the ocean can absorb carbon dioxide that originates from the atmosphere, the carbon dioxide levels are steadily increasing. The ocean's absorbing mechanisms, due to the rising of the ocean's temperatures, are unable to keep up with the pace. This results in acidification of oceans. Due to this, there are concerns that structures made of calcium carbonate may become vulnerable to dissolution, affecting corals and the ability of shellfish to form shells.

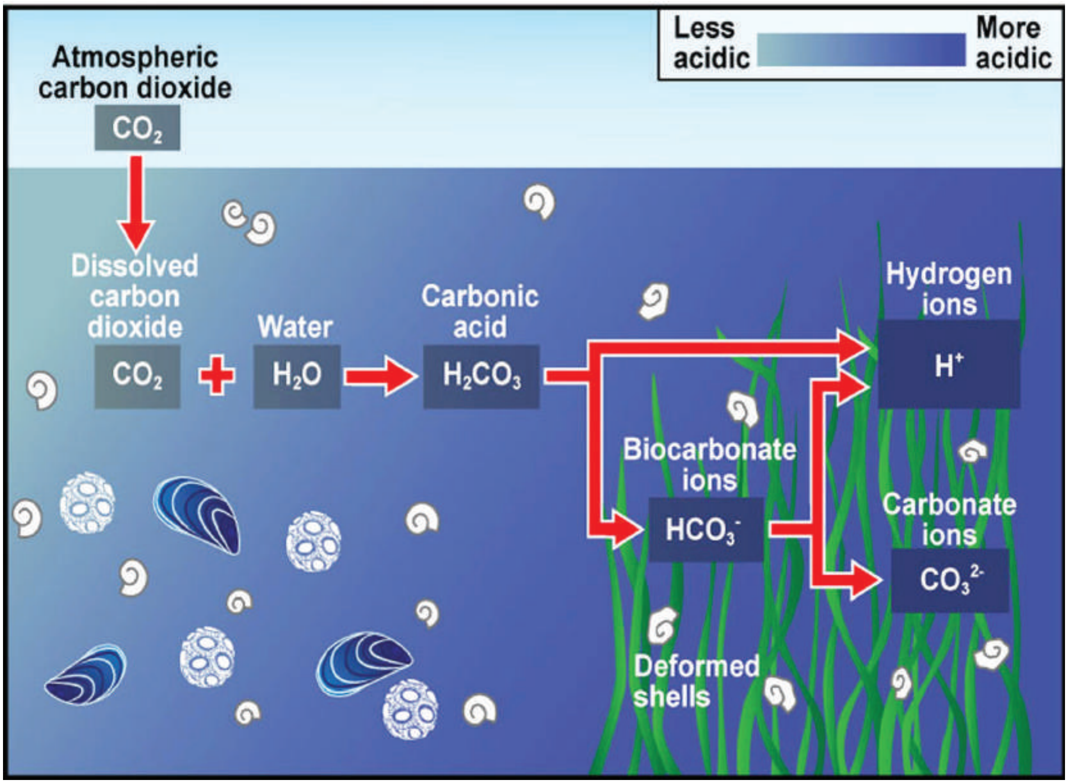


Figure 3.11: Ocean acidification.

3.3.3 Health Hazards

ACTIVITY 3.8: Diseases Caused by Polluted Water

Design a questionnaire to find out how many students in your class have been affected by one of the following diseases:


Diarrhoea


Vomiting


Typhoid


Diphtheria


Hepatitis


Kidney damage


Nerve disorders


Skin lesions

Include in your questionnaire the cause of their disease as diagnosed by their doctors.

Virtually all types of water pollution are harmful to the health of humans and animals. Water pollution may not damage our health immediately but can be harmful after long term exposure. People cannot survive without drinking water, and if their freshwater resources are polluted, they can fall ill by drinking it. Different types of pollutants affect human health in different ways:

- Heavy metals from industrial processes can accumulate in nearby lakes and rivers. These are toxic to aquatic life such as fish and shellfish, and subsequently to the humans who eat them. Heavy metals can slow down development; resulting in birth defects and some are carcinogenic, *i.e.*, can cause cancer.
- Industrial waste often contains many toxic compounds that damage the health of aquatic animals and those who eat them. Some of the toxins in industrial waste may only have a mild effect whereas others can be fatal. They can affect immune system, reproductive system or cause poisoning.
- Microbial pollutants from sewage often result in water-borne diseases that infect aquatic life and terrestrial life through drinking water. Microbial pollutants include bacteria, virus and protozoa.

Cause	Water-borne diseases
Bacterial infections	Typhoid Cholera Paratyphoid fever Bacillary dysentery
Viral infections	Infectious Hepatitis (jaundice) Poliomyelitis
Protozoal infections	Amoebic dysentery

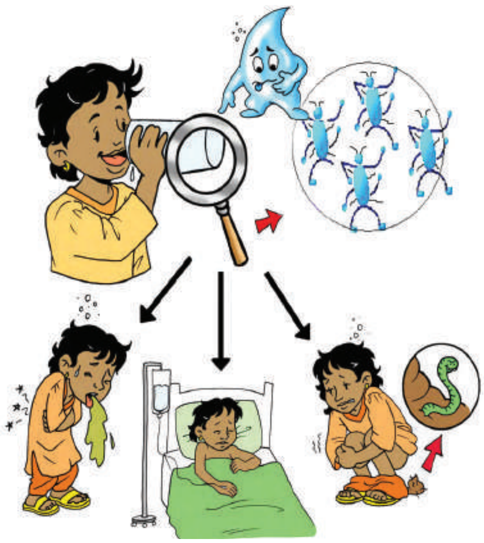
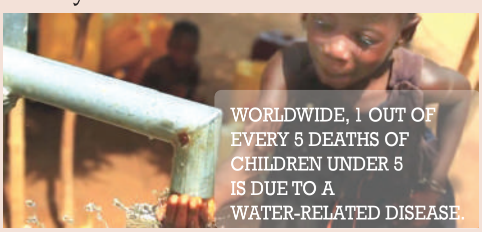


Figure 3.12: Contaminated water contains microbes that make us sick.

DO YOU KNOW?

Microbial water pollution is a major problem in the developing world. These illnesses are particularly dangerous for young children; in fact, they account for almost 60 per cent of early childhood deaths worldwide.



EXERCISE 3.9

1. Addition of pollutants to water changes its physical and chemical properties. (True or False)
2. The pH of pure water is ______ .
3. Water pollution is harmful to _______ and ______ .
4. What are the dangers of water pollution?
5. Name two water-borne diseases.

3.4 PREVENTION OF WATER POLLUTION

ACTIVITY 3.9: Illustrating Prevention of Water Pollution

'Water water everywhere but not a drop to drink'.

Comment on the statement given above. Make a poster/PowerPoint presentation on how you can save water. Display it in class.

There is no easy way to solve water pollution; if there were, it would not be so much of a problem. Broadly speaking, there are three different things that can help to tackle the problem—education, laws, and economics—and they work together as a team.

3.4.1 Education

Making people aware of the problem is the first step towards solving it. Education can help people determine their best strategies to avoid contaminating local water sources: avoiding urinating or defecating in or near the water; building toilets/sites for waste downhill from wells to reduce risks of contaminating groundwater; employing household water treatment and safe storage techniques are examples. Greater public awareness can make a positive difference. Awareness helps to prevent disposal of solid and human waste and chemical and industrial waste into waterways as much as possible. It also includes treating wastes before they go into waterways. This can be achieved by setting up of educational camps.

Figure 3.13: A volunteer educating learners about water pollution.

3.4.2 Laws

One of the biggest problems with water pollution is its transboundary nature. Many rivers cross countries, while seas span whole continents. Pollution discharged by factories in one country can cause problems in neighbouring nations, even when they have tougher laws and higher standards. Environmental laws can make it tougher for people to pollute, but to be really effective they have to operate across national and international borders. Proper implementations of national and international laws is another issue faced by the government. Without tougher implementation it is difficult to solve the problem of water pollution. As in Rwanda, the ban of polythenes is successfully implemented with inspection officers.

Figure 3.14: Say no to polythene bag.

3.4.3 Economics

Most environmental experts agree that the best way to tackle pollution is through something called the **polluter pays principle**. This means that whoever causes pollution should have to pay to clean it up, in one way or another. Polluter pays can operate in all kinds of ways. It could mean that shoppers should have to pay for their plastic grocery bags, as it is now common in Ireland, to encourage recycling and minimise waste. Or it could mean that factories that use rivers must have their water inlet pipes downstream of their effluent outflow pipes, so if they cause pollution, they themselves are the first people to suffer. Ultimately, the polluter pays principle is designed to hinder people from polluting. It makes it less expensive for them to behave in an environmentally responsible way.

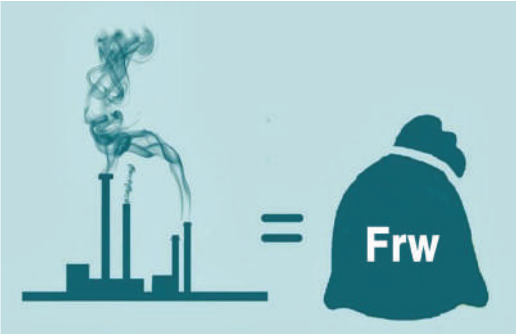


Figure 3.15: Polluter pays principle.

3.5 OUR CLEAN FUTURE

Life is ultimately about choices—and so is pollution. We can live with dirty surroundings, dead rivers, and fish that are too poisonous to eat. Or we can work together to keep the environment clean so the plants, animals, and people who depend on it remain healthy. We can take individual action to help reduce water pollution. These actions are:

1. **Use less water:** This might sound simplistic, but decreasing your water consumption is one of the keys to minimise water pollution. By reducing the amount of water you use, you will reduce the amount of water that flows into sewage treatment systems.

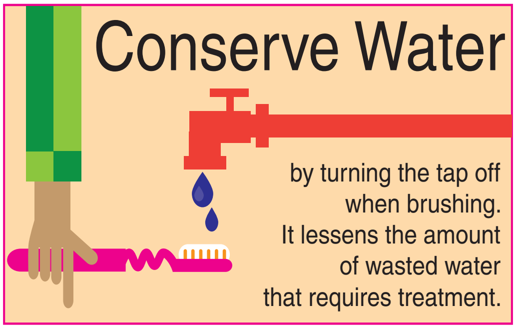


Figure 3.16: (a) A poster on water conservation.

2. **Use environment-friendly household products:** Don't use household products that contain chemicals. Instead, use green products, like biodegradable soap and all-natural toiletries.

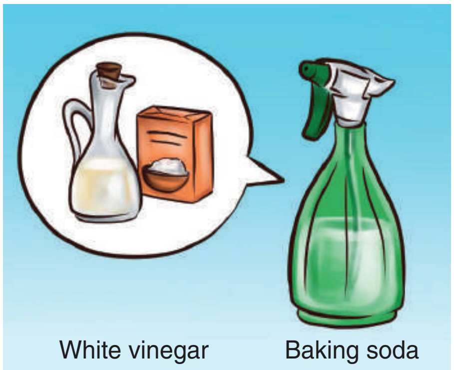


Figure 3.16: (b) Switch to white vinegar and baking soda to clean your home.

3. **Apply natural pesticides and fertilisers:** The use of chemical pesticides and fertilisers leads to water pollution because contaminated water seeps into ground water and runs off into nearby water sources.

Figure 3.16: (c) Avoid using chemical pesticides.

4. **Don't litter:** Avoid littering in rivers, lakes, and oceans.

Figure 3.16: (d) Do not throw litter in water bodies.

5. **Dispose off toxic products with care:** Make sure to dispose off toxic products, such as paints, solvents, and polishes, in the proper area. Do not pour them down your drain.

Figure 3.16: (e) Do not dispose toxic liquids in the sink.

6. **Follow the three Rs to prevent water pollution.**

3 Rs to Prevent Water Pollution
Refuse: Say No to water pollution
Recycle: Recycle water
Reduce: Minimise use of water

EXERCISE 3.10

1. What are different ways to solve water pollution?
2. Can awareness help to prevent disposal of human waste?
3. How should we conserve water?
4. What are the 3Rs to prevent water pollution?
5. We should avoid littering in
   (*a*) rivers (*b*) lakes
   (*c*) oceans (*d*) all of these.

3.6 SUMMARY

- Water pollution is the addition of any foreign substance (organic, inorganic, radioactive or biological) to water which produces harmful effect and decreases the usefulness of water.
- The substances which cause water pollution are known as water pollutants.
- The major pollutants of water include – Sewage, Nutrient-rich waste water, Chemical waste, Radioactive waste, Oil pollution, Plastic, Alien species and other forms.
- Sewage is the term used for wastewater that often contains faeces, urine and laundry waste. Sewage disposal is a major problem in developing countries as many people in these areas don't have access to sanitary conditions and clean water.
- The fertilisers used by farmers run off from the fields adding nutrients to water. This addition of nutrient results in eutrophication of water bodies.

- Waste from industries is discharged without treatment in water. This waste is termed as chemical waste as it is rich in chemicals like heavy metals and polychlorinated biphenyls (PCBs). These chemicals are responsible for various disorders in humans. Addition of CO_2 causes acidification in oceans.
- Radioactive waste is a waste that contains radioactive substance. They are released from nuclear plants, mines, refining processes and scientific institutions. These wastes are highly toxic and can cause genetic defects on exposure.
- Oil and oil wastes enter water bodies from refineries, storage tanks, automobile waste oil, industries and accidental spillage from ships. It affects aquatic life as well as birds.
- Plastic is the most common waste found in water. It is non-biodegradable, because of this it is banned in many countries.
- Alien species such as water hyacinth form dense mats on water bodies. They choke the water bodies resulting in death of aquatic animals.
- Other forms of pollution include heat pollution which is adding hot water to the water bodies.

3.7 GLOSSARY

- **Acidification:** reduction in the pH of the ocean over an extended period of time, caused primarily by uptake of carbon dioxide (CO_2) from the atmosphere.
- **Eutrophication:** the process of nutrient enrichment of water bodies and the subsequent overgrowth of plants on the surface of water.
- **Phytoplankton:** these are single-celled organisms of lakes, streams and oceans that make their own food from sunlight through photosynthesis.
- **Pollutant:** substances that cause pollution.
- **Polluter pays principle:** whosoever causes pollution should have to pay to clean it up, in one way or another.
- **Pollution:** contamination of the environment with substances which are harmful to living beings.
- **Radioactive substance:** a radioactive substance is unstable and produces dangerous kinds of radiation.
- **Sewage:** wastewater that often contains faeces, urine and laundry waste.
- **Toxic products:** capable of causing injury or death, especially by chemical means; poisonous.
- **Water-borne diseases:** diseases caused by micro-organisms in water.
- **Water pollution:** addition of foreign substances which produce harmful effects and decrease usefulness of water.

3.8 UNIT ASSESSMENT

I. Mutiple Choice Questions

1. The drinking water becomes polluted by the addition of
   (*a*) fertilisers (*b*) oil (*c*) sewage (*d*) all of these
2. Nutrient-rich waste water causes
   (*a*) acidification (*b*) distillation (*c*) sedimentation (*d*) eutrophication
3. Polychlorinated biphenyl is a type of
   (*a*) sewage waste (*b*) chemical waste (*c*) plastic waste (*d*) oil waste
4. Oil pollution is caused by
   (*a*) mining (*b*) invasion (*c*) institutes (*d*) none of these
5. Which of the following is non-biodegradable?
   (*a*) polythene (*b*) paper (*c*) oil (*d*) all of these
6. Heat pollution ________ the temperature of the water body.
   (*a*) raises (*b*) lowers (*c*) has no effect (*d*) both (*a*) and (*b*)
7. The three Rs to prevent water pollution include
   (*a*) reduce, recycle, reuse (*b*) refuse, recycle, reduce
   (*c*) restart, recycle, reduce (*d*) restart, refuse, reuse
8. ____________ water pollution is responsible for wide number of deaths of infants.
   (*a*) Microbial (*b*) Radioactive (*c*) Oil (*d*) Heat
9. Ways of prevention of water pollution include
   (*a*) education (*b*) law (*c*) economic (*d*) all of these
10. Choose water-borne disease.
    (*a*) Typhoid (*b*) Cholera (*c*) Both (*a*) and (*b*) (*d*) None of these

II. Open Ended Questions

1. What do you understand by water pollution?
2. What is a pollutant? Name the major water pollutants.
3. Why was polythene banned in Rwanda?
4. How can you say that alien species cause water pollution?
5. Comment on the dangers associated with water pollution.
6. What is Eutrophication? How is it caused? What are its effects?
7. What are the health hazards associated with drinking of polluted water?
8. What role do education and awareness play in preventing water pollution?
9. What do you understand by polluter pays principle?
10. What role can you play in preventing water pollution?

III. Practical-based Questions

1. Which of the following glasses of water is harmful for health?

Glass A

Glass B

(*a*) Water in Glass A is harmful for health
(*b*) Water in Glass B is harmful for health
(*c*) Both are harmful for health
(*d*) None of these

2. The symbol shown in figure is used for ....

(*a*) Radioactive waste (*b*) Chemical waste
(*c*) Waste water (*d*) Dirty clothes

3. Causes of water pollution are

(*a*)

(*b*)

(*c*) both (*a*) and (*b*) (*d*) none of these

PROJECT

Construct a simple water filter by using commonly available materials.

Unit 4

Effective Ways of Waste Management

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- describe the steps involved in effective waste management
- explain the importance and benefits of waste recycling
- discuss the various effects of waste materials and poor waste disposal.

KNOWLEDGE GAIN

Segregate the waste as biodegradable and non-biodegradable.

4.1 STEPS TO EFFECTIVE WASTE MANAGEMENT

ACTIVITY 4.1: Steps to Waste Management

You must have seen an overflowing dustbin and held your breath when you crossed it on the street. What made you hold your breath? Have you ever thought that you are also involved in creation of this foul smell? Discuss this with your friends. Also think how you may contribute in reduction of this foul smell.

Overflowing dustbin

Waste materials (or rubbish) especially household wastes, are called **garbage**. Every time you throw something, you produce garbage. The peels of fruit and vegetables, left over cooked food and fallen leaves of plants are thrown away as garbage.

Unwanted, unused objects are thrown as garbage. Can you list out some waste materials you contribute to garbage? The garbage we add to the dustbin rots to produce foul smell. You can help in lessening the foul smell and making the dustbin look more attractive by adopting effective ways of waste management.

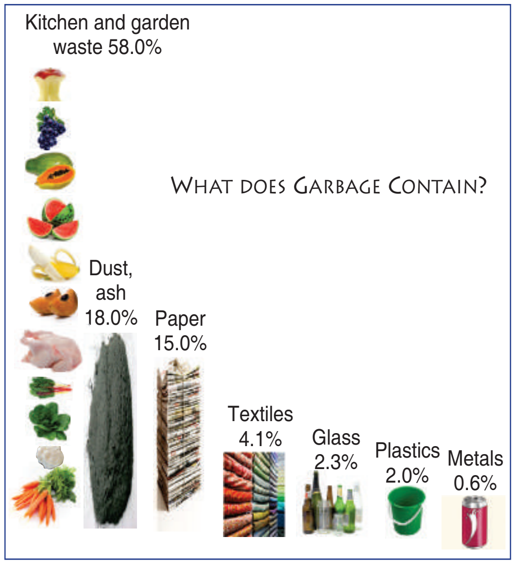


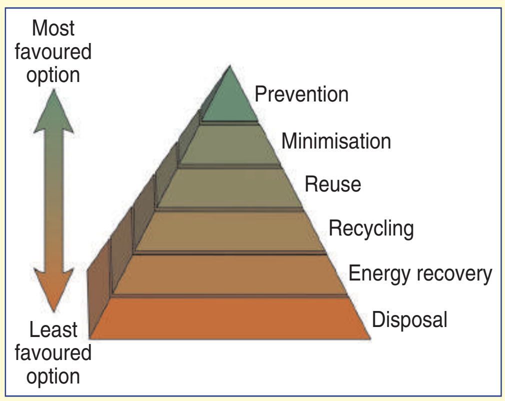
Figure 4.1: Components of garbage.

Many countries in Africa are facing significant challenges in relation to waste management. Waste generation is increasing. Only a sizeable portion of it is disposed on improperly located and operated dumpsites. It is having adverse impacts on environment and health.

ACTIVITY 4.2: Illustrating Waste Hierarchy

Observe the diagram. Sit in groups and discuss the terms given in waste hierarchy to manage waste.

Make presentation with your group on the same. Give your views and suggestions in presentation.



Waste hierarchy

Waste management includes all the processes of handling waste and reducing it. There are six steps to achieve effective waste management. These steps are studied as waste hierarchy. Waste hierarchy focuses more on waste minimisation. All the steps revolve around the same motive. These steps are:

4.1.1 Prevention

ACTIVITY 4.3: Illustrating Prevention of Waste

At the end of the day, after you have taken your dinner and the kitchen has been cleaned, investigate the dustbin of your house. Look at the waste generated in your house. Make a list of the type of items in the waste. Which of the items do you think could have been avoided? Discuss with your family members and try to avoid the waste you could.

Investigating the dustbin of your house

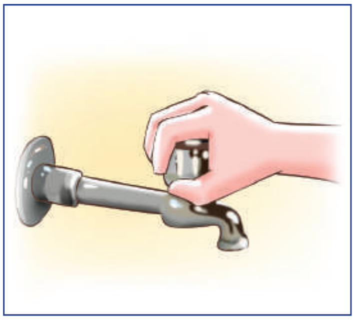
Prevention or avoidance means to try by all possible means to not produce the waste completely. It seeks to prevent waste from being generated. Waste prevention strategies include using less packaging, designing products to last longer. For example, we can stop using plastic carrier bags, and instead long-life bags, such as canvas bags.

This is less practical and very expensive. It requires research to completely avoid or replace a substance or process. It is like, completely stopping vehicles like car in the dream to stop pollution. It is the most environmentally preferred strategy.

4.1.2 Minimisation

The easiest method of waste management is to reduce creation of waste materials. It is done by reducing the amount of waste going to dustbins. It is a very cost-effective method. Reducing waste is self-explanatory. Waste minimisation involves redesigning products and/or changing societal patterns, concerning consumption and production. It can be achieved in homes using more efficient appliances or using their electric supply more effectively. The following are some ways by which you can contribute to minimise waste

- By not leaving heating appliances/gadgets on constantly .
- By closing water running too long.
- By boiling a kettle with the right amount of water in it.
- By switching off the lights whenever not in use.



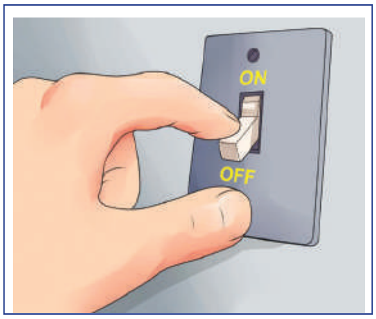


Figure 4.2: Close the running tap and switch off electricity when not in use.

4.1.3 Reuse

ACTIVITY 4.4: Showing Reuse of Waste

Students will take a nature walk, make observations, and collect natural objects for an art activity. Also collect some waste products which you can see. Now make some useful products from the materials you have collected. Showcase them in an art exhibition. Following are some things you may use.

This is simple; reduce waste by not wasting something, but reusing it. We should reuse items we normally throw away.

For example, we should use sacks or clothed bags for multiple times instead of throwing them away after just one use. It is a cost-effective method as it can reduce on our purchases. It includes refilling. Ink cartridges are an example, in which the only bit replaced is the stuff that is used up (*i.e.*, the ink). Glass milk bottles are also often collected, washed, refilled and delivered many times in their life. Can you name some other items that you may reuse? Reuse is often seen along with the value which is retained of the product.

I can reuse

Make a list of items you can reuse

1. ____________________ 4. ____________________

2. ____________________ 5. ____________________

3. ____________________ 6. ____________________

Name ____________________________

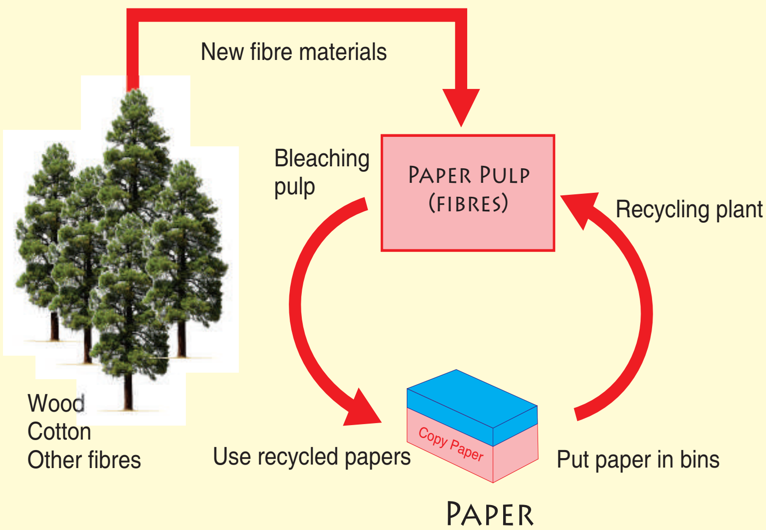
4.1.4 Recycling

ACTIVITY 4.5: Illustrating Recycling of Waste

Making a Model

To make your own recycled paper

- Take waste paper from old newspapers and magazines, and tear them into small pieces.



Recycling paper

- Take warm water in a bucket and add a little starch to it.
- Soak the paper in the water for 5–6 hours.
- Take the paper out of the water and pound it with a mortar and pestle till it becomes soft and fluffy.
- Add more starch to it to thicken it.
- Spread this pulp on a fine wire mesh, and press it to squeeze out the excess water.
- Carefully turn the wire mesh upside down on a smooth surface, and put some weights on it.
- Let the pulp dry for several hours.

Once it dries up, your hand-made recycled paper is ready! You may not be able to write on the recycled paper, but you can draw on it.

We should remember to recycle items that are recyclable. Recycling is a series of activities that includes collecting used, reused, or unused items that would otherwise be considered waste. It includes sorting and processing the recyclable products into raw materials. Further, re-manufacturing the recycled raw materials into new products. Recycling involves putting energy into a waste item to convert it to something else entirely, sometimes with lower grade and value. Consumers provide the last link in recycling by purchasing products made from recycled content. It is the preferred option when reuse is not an option, if an item is broken, or in a poor condition that means it cannot be reused. In the milk bottle example, reusing involves washing and refilling, while recycling would involve grinding the glass and melting it again, turning the raw material into a different grade of glass.

Figure 4.3: Recycling trash.

EXERCISE 4.1

1. What are the effective ways of waste management?
2. Prevention means preventing or avoiding of the ______ completely.
3. The easiest method of waste management is to reduce creation of waste materials. (True or False)
4. Name four items that you can reuse.
5. What are the components of garbage?

4.1.5 Energy Recovery

ACTIVITY 4.6: To Make a Biogas Generator

Materials

- 1-litre clear plastic bottle (sports drink bottles and other wide-mouthed receptacles work particularly well for this activity, but you can use whatever is in your recycling bin)
- A few balloons
- Duct tape
- 1/3 cup of raw vegetable scraps and grass
- 1/3 cup of soil from the outdoors (not bagged potting soil)
- Permanent marker
- Scoop or large spoon
- Funnel
- Ruler
- String

How to make it

- Mix the vegetable scraps, grass, and soil.
- Using the funnel, pour it into the bottle.
- Stretch a balloon carefully over the opening of the bottle, and duct tape around the balloon's base to seal it to the bottle and keep outside air out.
- Over the next few days, the microbes in the soil will digest the mixture and create methane gas, which will fill the balloon.
- Every other day, measure the amount of mixture in the bottle, and measure the circumference of the balloon by wrapping the string around it, marking it, and measuring the string.

Energy recovery from waste is the conversion of non-recyclable waste materials into usable heat, electricity, or fuel through a variety of processes, including heating. This process is often called waste-to-energy (WTE). It converts non-recyclable waste materials into electricity and heat. After energy is recovered, approximately ten per cent of the volume remains as ash, which is generally sent to disposal.

Energy recovery can also be achieved with bio-mass waste. We can either burn it directly, or compost it to capture methane, which can be burned to produce energy (biogas).

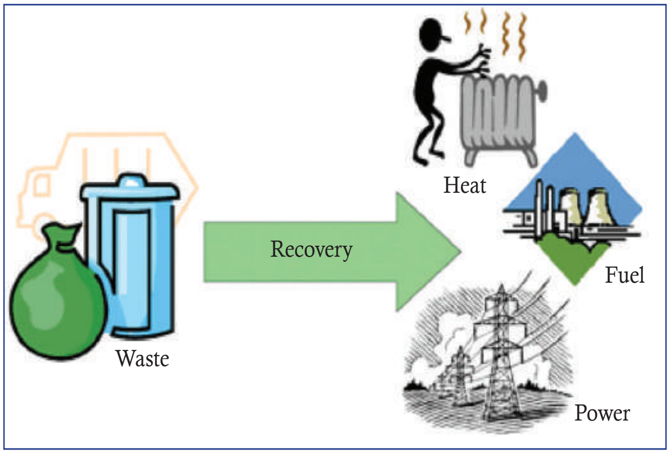


Figure 4.4: Waste to energy.

How to find out the valuable "wastes" from daily wastes and convert the "wastes" into treasure?

As for high calorific solid wastes, briquetting is one of the ways to turn the wastes into treasure. Briquetting technology is used to densify the loose combustible materials into solid composites. It can be made of different shapes and sizes with the presence of pressure and binding agents. Generally, there are a wide range of materials that can be used to make briquettes, such as waste paper, cardboard, water hyacinth, agricultural residues, charcoal dust, and wood wastes like sawdust, etc.

(*Source*: http://operation-boost.com/act/green-innovation/fuel-briquettes-from-waste/)

Try to find on internet and in library and research how briquetting is done.

4.1.6 Disposal

ACTIVITY 4.7: Showing Waste Disposal System

Visit a landfill and an incinerator site. Ask the people working there about description and details of the processes. Prepare a report and present in the class.

"Disposal" is also known as "burying it". Mostly, it is least favoured because we don't get anything back from what we put in the ground. Once something is in disposal and buried, there is no real harm the waste can do as it degrades naturally – although this can take thousands of years. Some CO_2 or other greenhouse gases may be released, however. Overall, sending something to disposal should be the last resort when the above five options have been tried.

Landfills are the most common form of waste disposal and are an important component of an integrated waste management system. In this method, the non-useful garbage is burned at a high temperature in a special kind of furnace called "incinerator".

Figure 4.5: A landfill site in Rwanda.

Composting is also a kind of disposal system. It is the process of converting plant and animal waste materials into manure. The biodegradable domestic garbage includes fruit and vegetable peels, left-over food and fallen leaves, etc. Compost is a natural fertiliser.

ACTIVITY 4.8: Preparing Compost

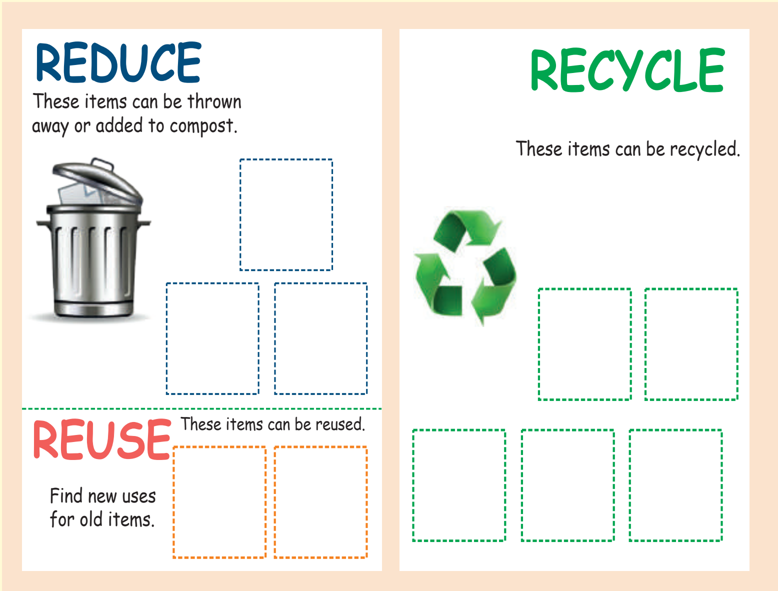
Dig a pit in the ground about 30 centimetres deep in a corner of the garden. Dump the plant wastes such as fruit and vegetable peels, left-over cooked food and fallen leaves into the pit on daily basis. The animal wastes like cow dung may also be added to the pit (see Figure below).

Preparation of compost

When the pit gets filled to the top with plant and animal wastes, a few buckets of water are added to the pit. The pit is then covered with a paste of soil and left undisturbed for about 3 months. The micro-organisms present in soil decompose the plant and animal waste materials buried in the pit to form compost. The compost formed can be dug up from the pit and used as a manure in the garden to grow plants. Please note that we should not dump plastics, glass pieces or metal objects in the compost pit because these are non-biodegradable waste materials which cannot be converted into compost by the micro-organisms present in the soil.

ACTIVITY 4.9: Game of Waste Management

Collect different pictures/images of waste items. Make cards by pasting them on cardboard. Mix them all. Sort the items (game pieces) into the category on the game board that you think is best for the earth.



Different waste items

ACTIVITY 4.10: Importance of 3Rs

Make a table showing wastes which can be reduced, recycled and reused. Discuss the importance of 3Rs in the class.

EXERCISE 4.2

1. Energy recovery can also be achieved with bio-mass waste. (True or False)
2. Disposal is also known as _____________.
3. What is the most common form of waste disposal?
4. Give one importance of waste disposal.
5. Compost is a natural __________ .

4.2 IMPORTANCE AND BENEFITS OF WASTE RECYCLING

ACTIVITY 4.11: Practising Recycling Week

Practise recycling week in your routine. Discuss the benefits of recycling in your class during the week. At the end of the week, present your findings as report.

If your approach is not towards disposing materials and other wastes, this recycling method will be a good approach for you. This is a process aimed to make undesirable objects into useful ones. By understanding what you can recycle and how you can recover valuable resources, you can be a very best part of helping the country to attain a positive environmental future.

Figure 4.6: Things to be recycled.

The idea behind recycling is to reduce energy usage, reduce volume of landfills, reduce air and water pollution, reduce greenhouse gas emissions and preserve natural resources for future use. Some of its benefits are:

4.2.1 Environmental Protection

- **Recycling includes reducing deforestation:** Recycling reduces the need for raw materials, so that our forests can be preserved.
- **Recycling helps reduce pollution:** The manufacturing process (including the extraction of raw materials from the earth) for many products releases waste that pollutes the environment. For example, the power plants that provide the energy needed in the manufacturing process produce gases that pollute the air.

  The more we buy (as opposed to reducing), and the more we throw away (instead of reusing and recycling them), the more waste we create. This waste releases poisonous gases and chemicals into the environment during the disposal process.

  For example, when the non-biodegradable products in our waste are burned, they often emit gases that deplete the ozone layer in the atmosphere. This ozone layer depletion allows more ultraviolet radiation to reach our living atmosphere, causing harm to the living organism such as skin cancer.

DO YOU KNOW?

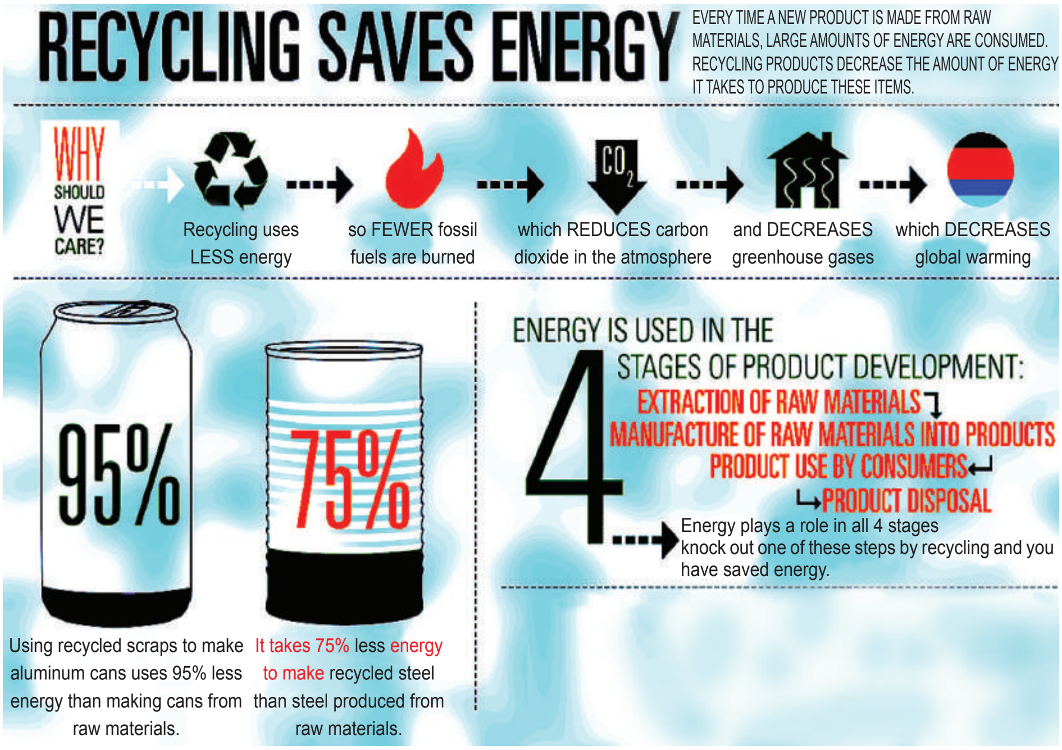


Figure 4.7: Recycling saves energy.

- **Recycling also aids in reducing land pollution:** Imagine if the various materials (especially the non-biodegradable ones) in our waste could be sent for recycling, the amount of waste that needs to be incinerated or buried in landfills would be reduced. With less waste, there would also be less need for landfills and incinerators. These lands could then be freed up for other uses.

4.2.2 Conservation of Natural Resources

By recycling, there is less need for raw materials. Less need for the resources that have to be planted, grown and harvested or extracted from ground (ores and minerals) and less need for space. Using recovered steel materials (through recycling) in place of raw ones, there is less demand for mining practices, which are environmentally devastating.

Figure 4.8: Recycling in steel plant.

Recycling also helps us make full use of the precious materials that we have spent much resources, energy and effort to extract from the earth.

4.2.3 Energy Saving

It takes much less energy to make products using recycled materials as compared to making products from raw materials. This is because when products are made from raw materials, energy is not only required in the production process, but also in the extraction of the raw materials from the earth and the transportation of these materials to the manufacturing plants. Whereas for production, using recycled materials, the raw materials are already available. The recyclables are hardly broken down into their basic components for use in manufacturing.

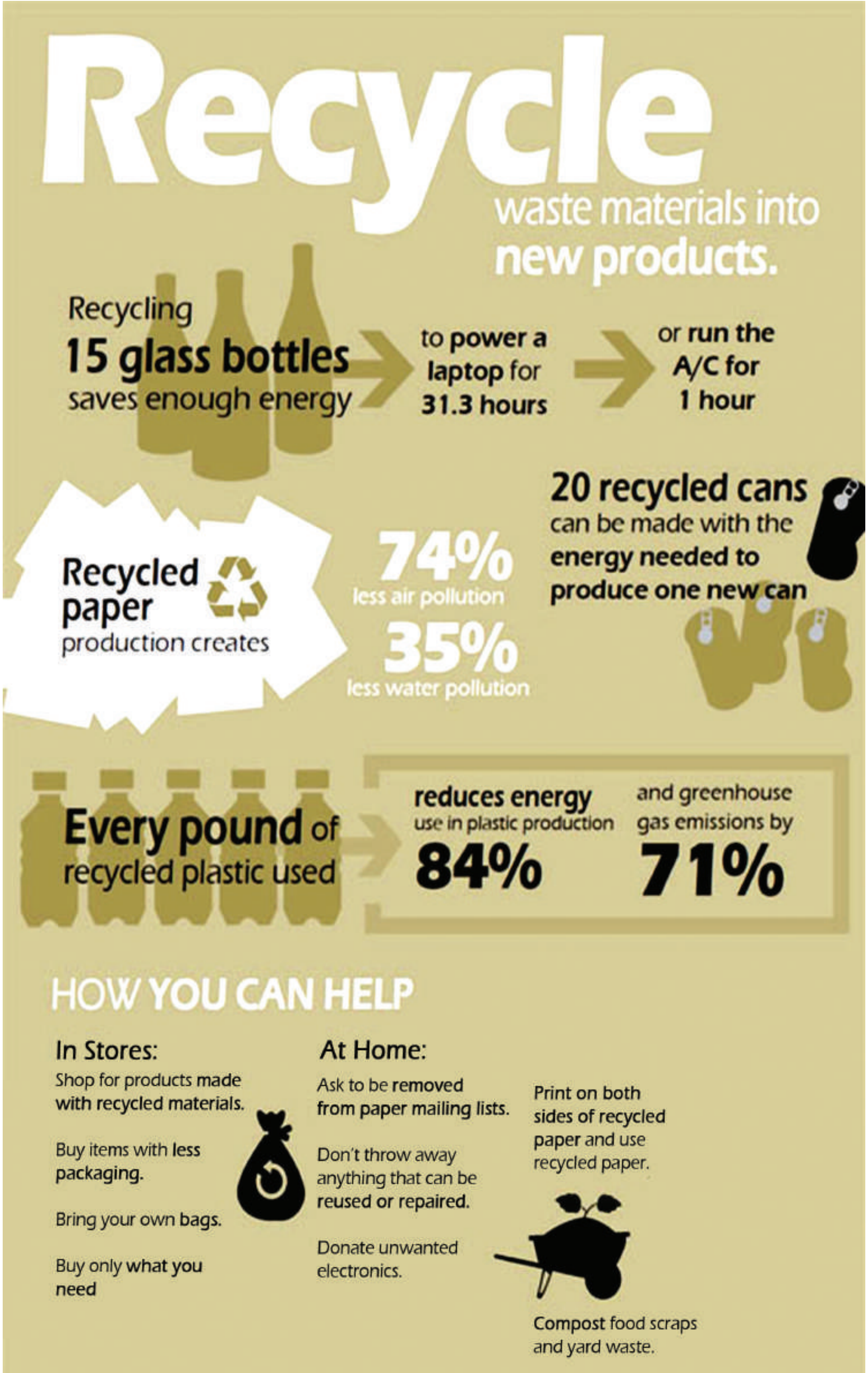


Figure 4.9: Recycling saves energy.

4.2.4 Job Creation

Recycling is more labour-intensive than landfilling or incineration. This means that building the recycling industry is a way to create more jobs. More people will be engaged while you recycle. This will also help in improving the economy of the country. The various recycling industries include paper recycling units, steel recycling plants, plastic recycling plants and glass recycling units.

Figure 4.10: Recycling helps in flourishing jobs.

EXERCISE 4.3

1. What are the benefits of waste recycling?
2. Recycle does not help reduce pollution. (True or False)
3. How do wastes pollute our environment?
4. Recycle waste materials into ______ .
5. Recycling industry is a way to create jobs. How?

4.3 EFFECTS OF WASTE AND POOR DISPOSAL

ACTIVITY 4.12: Group Discussion on Poor Disposal of Garbage

Imagine we all throw garbage, junk and rubbish away anyhow. Imagine there was no authority to supervise waste management activities from all the sources mentioned earlier. Imagine we all just sent our rubbish to the landfill, or just dumped them in a nearby river. What do you think will happen? A disaster! Engage in a group discussion to talk about the same.

Poor disposal of garbage

Modernisation and progress has had its share of disadvantages and one of the main aspects of concern is the pollution it is causing to the earth – be it land, air, and water. With increase in the global population and the rising demand for food and other essentials, there has been a rise in the amount of waste being generated daily by each household. Solid wastes, when improperly disposed can be an environmental hazard because, the surrounding environment and the fish in the nearby ponds get affected. This improper disposal can lead to death of fish as well as diseases to man, *e.g.*, dysentery, cholera and some other fatal diseases.

Some of these wastes can also be very harmful to the atmosphere. These wastes when improperly dumped into the atmosphere can lead to the destruction of the ozone layer and may cause diseases such as cancer. As a result, of late a problem has cropped up, *i.e.*, global warming. Air pollution can also lead to formation of acidic rain which is dangerous to crop life since it leads to the removal of soil fertility from the surface of the ground.

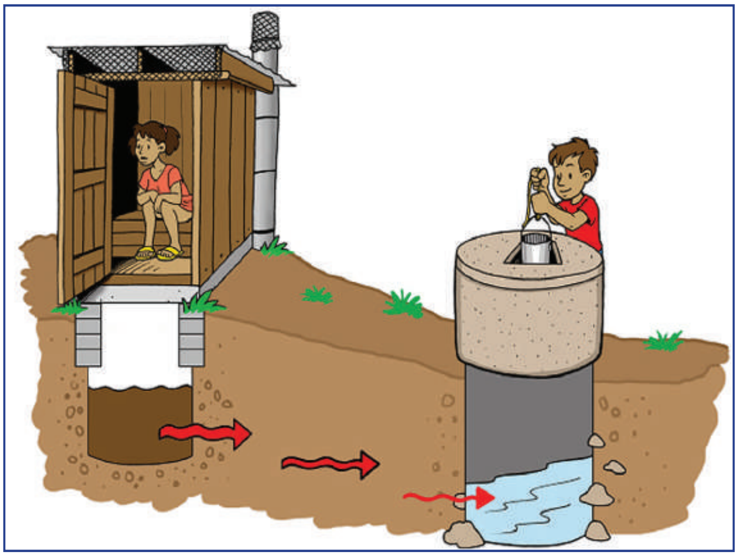


Figure 4.11: Poor sanitation contaminates water.

It also affects drainage. When solid wastes are dumped in drainage channels and gutters, they block the flow of the sewage. This may cause flood. At the same time, solid wastes also affect soil drainage which hinders the growing of crops. Since some of the waste materials are waterproof, they can be dangerous to the aeration system of the soil and hinder agriculture. It also leads to the reduction of fertile cultivatable land in the form of dumping sites. This in turn affects the country since Rwanda depends on agriculture for exports. It also pollutes the underground water which we use for various activities.

Waste materials like toxic chemicals if consumed by animals can be very dangerous to life and worse still if these wastes are dumped in water bodies. They are dangerous to aquatic life. Poor solid waste also leads to the death of animals (especially domestic animals). Death of animals like cattle leads to poverty and the death of animals like dogs, leads to insecurity in homes.

Figure 4.12: Open sewage and sanitation conditions.

Poor Domestic Waste management also displays an ugly scenario of the environment. This can affect the tourism industry, as the tourists may not get attracted to visit the country. It also leads to the spread of diseases in such a way that when wastes like broken bottles are dumped anywhere, they collect water in them (when it rains) and this may become a breeding ground for mosquitoes. Wastes like human stool causes diseases when poorly dumped, as the flies will carry the germ from the stool. It can also lead to human injury. For example, when a person walks and steps on the broken bottles or nails or even pins (sharp objects), he can get injured which may lead to bleeding.

Figure 4.13: Poor domestic waste management hindering scenic beauty.

Uncontrolled dumping of solid waste can lead to wastage of land where we find lots of land being used as dumping sites for wastes. These same pieces of land are later on neglected by the inhabitants of the area.

ACTIVITY 4.13

Go out on field trips and study tours with your teacher to different industrial sites. Visit garbage pits or biogas plants. Understand and know the prevalent problem of waste and its proper utilisation. Now submit a poster to your teacher showing techniques of waste management.

EXERCISE 4.4

1. Improper disposal of waste can lead to death of ______ as well as _______ .
2. Name two diseases caused by poor disposal of wastes.
3. Poor domestic waste management shows an ugly scenario of environment. (True or False)
4. What are the effects of poor waste disposal?
5. How does poor disposal of wastes cause flood?

4.4 SUMMARY

- Waste material, especially household wastes are called garbage. It is very important to manage this waste material in a proper way. Waste management includes all the processes of handling waste and reducing it. There are six steps to follow:
  - **Prevention:** It means avoiding the waste completely. It is less practical and expensive.
  - **Minimisation:** It means reducing the creation of waste material. It is very cost-effective.
  - **Reuse:** It means reducing waste by not wasting it but reusing it. It is also an effective method.
  - **Recycling:** It means collecting used, reused or unused items and processing them to form new products. It is used when reuse is not an option. It helps in environmental protection, conservation of natural resources, energy saving and job creation.
  - **Energy Recovery:** It means converting waste into usable heat, electricity or fuel. It is an effective method to produce energy.
  - **Disposal:** It means burying. It can take two forms—landfill and composting.
- Waste can cause deleterious effects not only to living beings but also to the environment. It causes diseases like cholera, dysentery and also pollutes air, land and water. The pollution created by waste also affects the tourism industry by affecting the scenic beauty of the country.

4.5 GLOSSARY

- **Biodegradable waste:** waste which can be broken down, by micro-organisms.
- **Breeding ground:** an area where birds, fish, or other animals habitually breed.
- **Composting:** the process of converting plant and animal waste materials into manure.
- **Energy recovery:** the conversion of non-recyclable waste materials into usable heat, electricity, or fuel through a variety of processes.
- **Garbage:** waste materials (or rubbish), especially household wastes, are called garbage.
- **Global warming:** the increase of Earth's average surface temperature due to effect of greenhouse gases.
- **Incinerator:** a vessel where incineration occurs. Incineration is a waste treatment process that involves the combustion of organic substances contained in waste materials.
- **Landfill:** a site for the disposal of waste materials by burial.
- **Non-biodegradable waste:** waste that cannot be broken down. It remains in the environment as such.
- **Ores:** the ores are extracted from the earth through mining; they are then refined to extract the valuable element, or elements.
- **Waste management:** activities and action required to manage the waste.
- **Waste recycling:** the processing of used materials (*waste*) into new, useful products.

4.6 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Landfill is an effective way of
   (*a*) Disposal (*b*) Recycle (*c*) Reuse (*d*) Reduce
2. Wastes which can be broken down by microorganisms are said to be
   (*a*) Toxic (*b*) Biodegradable
   (*c*) Non-biodegradable (*d*) Deleterious
3. The most favoured option of waste hierarchy is
   (*a*) Minimisation (*b*) Energy recovery (*c*) Prevention (*d*) Disposal
4. Glass bottles can be
   (*a*) Reused (*b*) Recycled (*c*) Reduced (*d*) All of these
5. Recovering the waste can produce
   (*a*) Fuel (*b*) Water (*c*) Heat (*d*) Both (*a*) and (*c*)
6. Generation of fertilisers from waste is
   (*a*) Briquetting (*b*) Composting (*c*) Landfill (*d*) None of these

7. Recycling benefits in ………………
   (*a*) protection of environment
   (*b*) developing better quality product
   (*c*) both (*a*) and (*b*)
   (*d*) none of these
8. Waste disposal affects
   (*a*) Government (*b*) Electricity (*c*) Rain (*d*) None of these
9. Using canvas bags instead of plastic is an example of
   (*a*) Prevention (*b*) Minimisation (*c*) Disposal (*d*) Energy recovery
10. ……………… is more labour intensive.
   (*a*) Incineration (*b*) Landfill (*c*) Recycling (*d*) Reuse

II. Open Ended Questions

1. Add the following waste to the appropriate bin.

2. What is briquetting?
3. State the benefits of waste recycling.
4. What is waste management?
5. Name the six steps of effective waste management.
6. How can you contribute to minimise the waste?
7. What is waste to energy?
8. How does poor waste handling cause diseases?
9. How does poor disposal technique affect tourism sector?
10. State two examples on how energy is recovered from recycling.

PROJECT

Make a presentation showing benefits of waste recycling.

Unit 5

Categories of Chemical Reactions

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- explain the difference between a decomposition reaction and combination reaction.
- explain single displacement, double displacement (precipitation and neutralisation) and combustion reactions.
- write and balance ionic equations.

KNOWLEDGE GAIN

A firework involves many different chemical reactions occurring at the same time.

ACTIVITY 5.1: Illustrating Chemical Reactions

Consider the following situations of daily life and think about what happens when:

- milk is left at room temperature on a hot day.
- an iron pan/nail is left exposed to humid atmosphere.
- grapes get fermented.
- food is cooked.
- food gets digested in our body.
- we respire.

Explanation

In all these situations, the nature and the identity of the initial substance have somewhat changed. We know that whenever a chemical change occurs, a chemical reaction has taken place.

In a chemical reaction, the reactants are converted into products. *The conversion of reactants into products in a chemical reaction is often accompanied by some features which can be observed easily.* The easily observable features (or changes) which take place as a result of chemical reactions are known as **characteristics of chemical reactions**. The important characteristics of chemical reactions are:

(*i*) Evolution of a gas

(*ii*) Formation of a precipitate

(*iii*) Change in colour

(*iv*) Change in temperature

(*v*) Change in state

Any one of these general characteristics can tell us whether a chemical reaction has taken place or not. For example, if on mixing two substances, any of the above characteristics occurs, then we can say that a chemical reaction has taken place.

Note: A change in temperature can take place even when a physical change occurs. Example: Dissolving NH4NO3 in water or NaOH in water

5.1 TYPES OF REACTIONS

There are millions of known chemical reactions. Many of these chemical reactions have common aspects, so they can be grouped into specific classes. The majority of chemical reactions (not all) fall into the following major categories:

- Combination reactions
- Decomposition reactions
- Single replacement reactions
- Double displacement reactions
- Combustion reactions

5.1.1 Combination Reactions

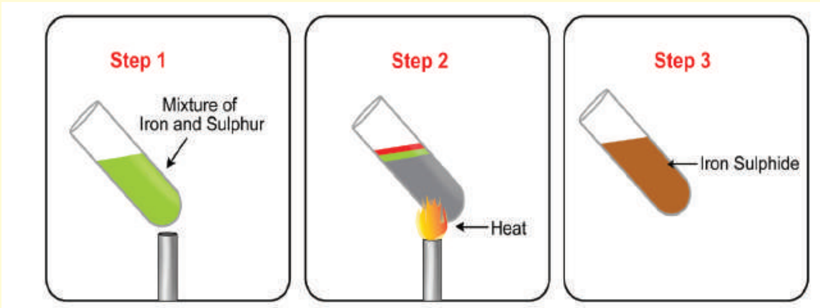
ACTIVITY 5.2: Combination of Iron and Sulphide (Combination Reaction)

Materials Required

Iron powder (filings) (7 grams), sulphur (4 grams), test tube, Bunsen burner, and pair of tongs.

Procedure

- Prepare a mixture containing iron powder and sulphur powder in the ratio 7:4 by mass.
- Note the appearance of the pure elements and the mixture.
- Take about 0.5 g of the mixture into a hard glass test tube.
- Heat the mixture at the base of the test tube—gently at first and then more strongly (use a blue flame throughout)
- Allow the test tube to cool down.
- Observe the product.



Can you write the chemical equation for this reaction?

Safety

- It is advisable to wear protective gloves and heat the mixture using tongs.
- Eye protection required.

In this activity, you will observe that when we start heating a mixture of iron filing and sulphur, the sulphur melts and reacts with the iron filings to form the compound iron (II) sulphide. In this reaction, two different elements combine to form a single product. This is an example of **combination reaction** or **a synthesis reaction.**

Combination reactions are those reactions in which a single product is produced from two (or more) reactants. The general equation for a combination reaction involving two reactants is

$$X + Y \longrightarrow XY$$

The reactants X and Y can be elements or compounds, or a compound and an element. The product XY is always a compound.

Combination reactions may involve

- the combination of two elements to form a compound.
- the combination of a compound and an element to form a new compound.
- the combination of two compounds to form a new compound.

We will now discuss some examples of combination reactions:

Element + Element ⟶ Compound

Example 1

When iron powder is heated with sulphur, iron sulphide is formed.

$$\underset{\text{Iron}}{Fe\,(s)} + \underset{\text{Sulphur}}{S\,(s)} \xrightarrow{\text{Combination}} \underset{\text{Iron sulphide}}{FeS\,(s)}$$

In this reaction, two elements, iron and sulphur are reacting together to form a single compound—iron sulphide. So it is a combination reaction.

Example 2

When sodium metal reacts with chlorine, sodium chloride is formed.

$$2Na\,(s) + Cl_2\,(g) \xrightarrow{\text{Combination}} 2NaCl\,(s)$$

Sodium   Chlorine   Sodium chloride

In this reaction, two elements, sodium and chlorine, combine together to form a single compound—sodium chloride. So, this is a combination reaction.

Example 3

Hydrogen combines with chlorine to form hydrogen chloride.

$$H_2\,(g) + Cl_2\,(g) \xrightarrow{\text{Combination}} 2HCl\,(g)$$

Hydrogen   Chlorine   Hydrogen chloride

In this reaction, hydrogen and chlorine react together to form a single compound, hydrogen chloride gas. So, this is an example of combination reaction. This combination reaction is used in industry for the manufacture of hydrochloric acid (Hydrogen chloride gas on dissolving in water forms hydrochloric acid).

Compound + Element → Compound

Example 4

Carbon monoxide reacts with oxygen to form carbon dioxide.

$$2CO(g) + O_2(g) \xrightarrow{\text{Combination}} 2CO_2(g)$$

Carbon monoxide   Oxygen   Carbon dioxide

In this reaction, carbon monoxide compound reacts with oxygen to form a new compound carbon dioxide. So, this is a combination reaction.

Example 5

Phosphorus trichloride reacts with chlorine to form phosphorus pentachloride.

$$PCl_3\,(l) + Cl_2\,(g) \xrightarrow{\text{Combination}} PCl_5\,(s)$$

Phosphorus trichloride   Chlorine   Phosphorus pentachloride

In this reaction, phosphorus trichloride reacts with chlorine to form a new compound phosphorus pentachloride. So, this is a combination reaction.

Compound + Compound → Compound

Example 6

Calcium oxide reacts with carbon dioxide to form calcium carbonate.

$$CaO\,(s) + CO_2\,(g) \xrightarrow{\text{Combination}} CaCO_3\,(s)$$

Calcium oxide (quicklime)   Carbon dioxide   Calcium carbonate

In this reaction, calcium oxide (quicklime) and carbon dioxide combine together to produce a new compound calcium carbonate. So, this is a combination reaction.

Example 7

Ammonia reacts with hydrogen chloride to form ammonium chloride.

$$NH_3\,(g) + HCl\,(g) \xrightarrow{\text{Combination}} NH_4Cl\,(s)$$

Ammonia   Hydrogen chloride   Ammonium chloride

In this reaction, two compounds, ammonia and hydrogen chloride, combine together to produce a new compound, ammonium chloride. So, this is a combination reaction.

EXPERIMENT 1

Aim

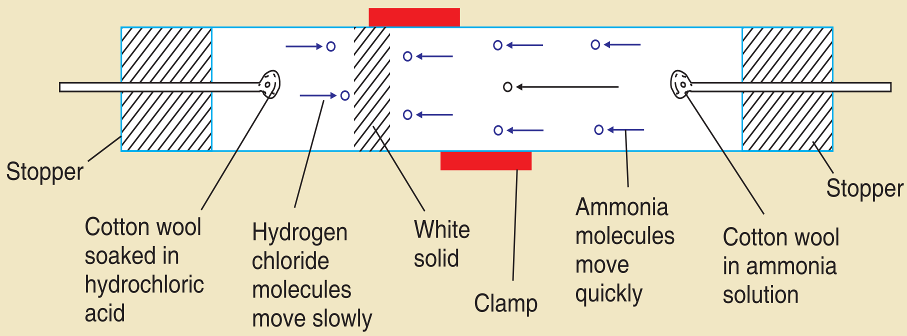
Combination of ammonia and hydrogen chloride.

Materials Required

Ammonia solution, hydrochloric acid, glass rods (two), cotton buds, long glass tube, stopper (two) and clamp.

Procedure

- Fix a cotton bud on two glass rods.
- Mark the glass rods as A and B.
- Dip rod A into concentrated ammonia solution and rod B into concentrated hydrochloric acid.
- Arrange all apparatuses as shown in diagram below.



Note: Insert the two stoppers at the end of glass tube (see diagram)

- Observe the glass tube after a few minutes.
  Write the balanced chemical equation for this reaction.

Safety

- Do not soak the cotton buds into chemicals with hand.
- The experiment must be done carefully in the presence of teacher.
- Make sure that the glass tube is clamped properly.
- Amonia attacks nose, nose protection is needed

Explanation

In the experiment, you will observe that when ammonia reacts with hydrochloric acid, a white solid appears inside the glass tube. This solid is ammonium chloride (NH_4Cl). Here, two reactants react (or combine) to give one product. So this is a **combination reaction.**

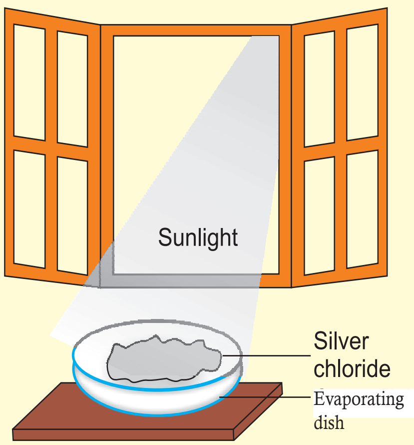
EXERCISE 5.1

1. When iron powder is heated with sulphur, ______ is formed.
   This is an example of ______ .
2. Complete the reaction: $H_2(g) + Cl_2(g) \longrightarrow ?$
3. Calcium oxide reacts with sulphur dioxide to form calcium carbonate. (True or False)
4. Name the compound formed when ammonia reacts with hydrogen chloride.
5. Hydrogen chloride gas on dissolving in water forms ______ acid.

5.1.2 Decomposition Reactions

ACTIVITY 5.3: Decomposition of Silver Chloride

- Take about 2 g of silver chloride in a evaporating dish.
- Note the colour of silver chloride.
- Place the China dish in sunlight for 10–30 minutes.
- Observe the colour of silver chloride after 30 minutes.



In Activity 5.3, you will observe that white silver chloride turns grey in sunlight. This is due to the decomposition of silver chloride into silver and chlorine by light.

Reaction

$$2AgCl(s) \xrightarrow{Sunlight} 2Ag(s) + Cl_2(g)$$

Decomposition reactions are those reactions in which a single reactant breaks down into two (or more) simpler substances (elements or compounds). The general equation for a decomposition reaction in which there are two products, is

$$XY \longrightarrow X + Y$$

The reactant XY is always a compound. The products X and Y may be elements or compounds.

In other words, decomposition reactions are opposite of combination reactions. These reactions often involve an energy source such as heat, light, or electricity which breaks apart the bonds of compounds.

The products of decomposition reactions may be

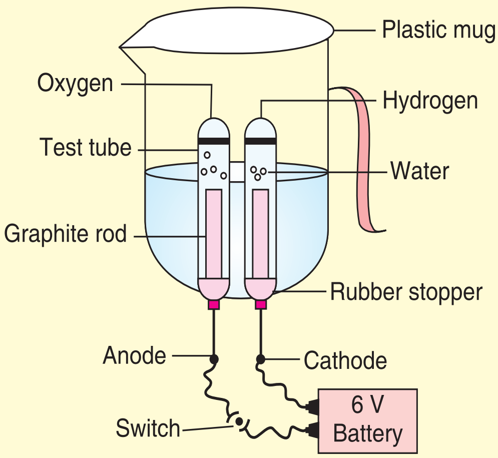
- two elements
- one (or more) elements and one (or more) compounds
- two (or more) compounds.

Some examples of decomposition reactions are:

Compound → Element + Element

ACTIVITY 5.4: Electrolysis of Water

- Take a plastic mug. Drill two holes at its base and fit rubber stoppers in these holes. Insert carbon electrodes in these rubber stoppers as shown in figure.
- Connect these electrodes to a 6-volt battery.
- Fill the mug with water such that the electrodes are immersed. Add a few drops of dilute sulphuric acid to the water.
- Take two test tubes filled with water and invert them over the two carbon electrodes.
- Switch on the current and leave the apparatus undisturbed for some time.
- You will observe the formation of bubbles at the electrodes. These bubbles displace water in the test tubes.



Electrolysis of water

- Is the volume of the gas collected the same in both test tubes?
- Once the test tubes are filled with the respective gases, remove them carefully.
- Test these gases one by one by bringing a burning candle close to the mouth of the test tubes.

  Caution: This step must be performed carefully by the teacher.
- What happens in each case?
- Which gas is present in each test tube?

Example 1

When electricity is passed through acidified water, it decomposes into hydrogen and oxygen.

$$\underset{\text{Water}}{2H_2O(l)} \xrightarrow[\text{Decomposition}]{\text{Electricity}} \underset{\text{Hydrogen}}{2H_2(g)} + \underset{\text{Oxygen}}{O_2(g)}$$

In this reaction, water splits up to form hydrogen and oxygen. This decomposition reaction takes place by the action of electricity. This reaction is called electrolysis of water.

Compound → Compound + Element

Example 2

In the presence of light, hydrogen peroxide decomposes into water and oxygen.

$$2H_2O_2\ (l) \xrightarrow{\text{light}} 2H_2O\ (l) + O_2\ (g)$$

The rate of decomposition of hydrogen peroxide is increased by the catalyst MnO_2. A 30% solution of hydrogen peroxide rapidly decomposes to oxygen and water when treated with manganese dioxide. This decomposition is exothermic.

$$\underset{\text{Hydrogen peroxide}}{2H_2O_2\ (l)} \xrightarrow{MnO_2} \underset{\text{Steam}}{2H_2O(g)} + \underset{\text{Oxygen}}{O_2(g)}$$

Example 3

When Sodium nitrate is heated, it decomposes to produce sodium nitrite and oxygen. This reaction takes place at a temperature of 380–500°C.

$$\underset{\text{Sodium nitrate}}{2NaNO_3(s)} \xrightarrow[\Delta]{380-500°C} \underset{\text{Sodium nitrite}}{2NaNO_2(s)} + \underset{\text{Oxygen}}{O_2(g)}$$

Example 4

When potassium chlorate is heated in the presence of manganese dioxide catalyst, it decomposes to give potassium chloride and oxygen.

$$\underset{\text{Potassium chlorate}}{2KClO_3} \xrightarrow[\text{Decomposition}]{\text{Heat}} \underset{\text{Potassium chloride}}{2KCl(s)} + \underset{\text{Oxygen}}{3O_2(g)}$$

In this reaction, potassium chlorate decomposes into potassium chloride and oxygen. So this is a decomposition reaction. In this reaction, catalyst manganese dioxide (MnO_2) allows the decomposition to occur at a lower temperature. This reaction is used for preparing small amount of oxygen in the laboratory.

Compound → Compound + Compound

Example 5

When calcium carbonate is heated, it decomposes to give calcium oxide and carbon dioxide.

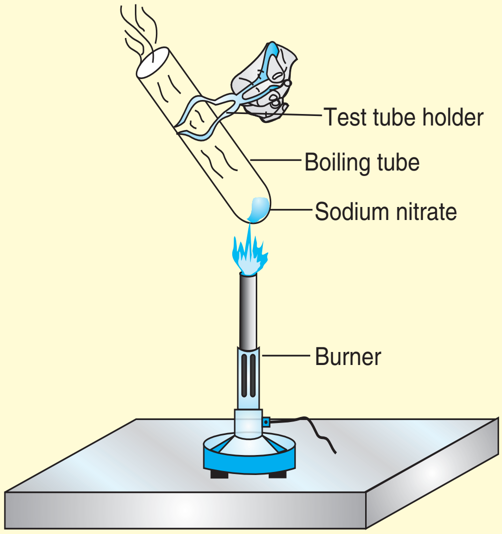
$$\underset{\text{Calcium carbonate}}{CaCO_3} \xrightarrow{\text{Heat}} \underset{\text{Calcium oxide (Quicklime)}}{CaO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

In this reaction, calcium carbonate breaks up into two simpler compounds—calcium oxide and carbon dioxide. So, this is a decomposition reaction.

- Calcium carbonate is also called limestone.
- Calcium oxide is also called lime or quicklime.
- Calcium oxide obtained from the decomposition of calcium carbonate is used in the manufacture of glass and cement.

ACTIVITY 5.5: Decomposition of Sodium Nitrate

- Take about 2 g sodium nitrate powder in a boiling tube.
- Hold the boiling tube with a pair of tongs and heat it over a flame, as shown in figure.
- What do you observe? Note down the change, if any.



Thermal decomposition of sodium nitrate

Explanation

You will observe a white or slightly yellow residue of sodium nitrite in the test tube.

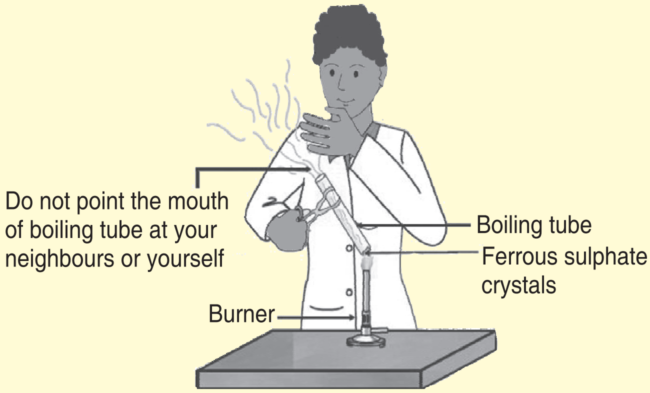
Reaction

$$\underset{\text{Sodium nitrate}}{2NaNO_3(s)} \xrightarrow[\text{Decomposition}]{\text{Heat}} \underset{\text{Sodium nitrite}}{2NaNO_2(s)} + \underset{\text{Oxygen}}{O_2(g)}$$

ACTIVITY 5.6: Heating Ferrous Sulphate Crystals

- Take about 2 g ferrous sulphate crystals in a dry boiling tube.
- Note the colour of the ferrous sulphate crystals.
- Heat the boiling tube over the flame of a burner or spirit lamp as shown in figure.
- Observe the colour of the crystals after heating.

Notice that the green colour of the ferrous sulphate crystals changes.



Correct way of heating the boiling tube containing crystals of ferrous sulphate

In Activity 5.6, when ferrous sulphate is heated strongly, it decomposes into ferric oxide, sulphur dioxide and sulphur trioxide.

$$\underset{\text{Ferrous sulphate}}{2FeSO_4(s)} \xrightarrow[\text{Decomposition}]{\text{Heat}} \underset{\text{Ferric oxide}}{Fe_2O_3(s)} + \underset{\text{Sulphur dioxide}}{SO_2(s)} + \underset{\text{Sulphur trioxide}}{SO_3(g)}$$

In this reaction, you can observe that a single reactant breaks down to give simpler products. So this is a decomposition reaction. Ferrous sulphate crystals ($FeSO_4{\cdot}7H_2O$) lose water when heated and the colour of the crystals changes. It then decomposes to ferric oxide (Fe_2O_3), sulphur dioxide (SO_2) and sulphur trioxide (SO_3). Ferric oxide is a solid, whereas SO_2 and SO_3 are gases.

EXPERIMENT 2

Aim

Decomposition of Hydrated Copper (II) Sulphate.

Materials Required

Copper sulphate crystals (2 g), evaporating dish, burner, tripod, balance, wire gauze.

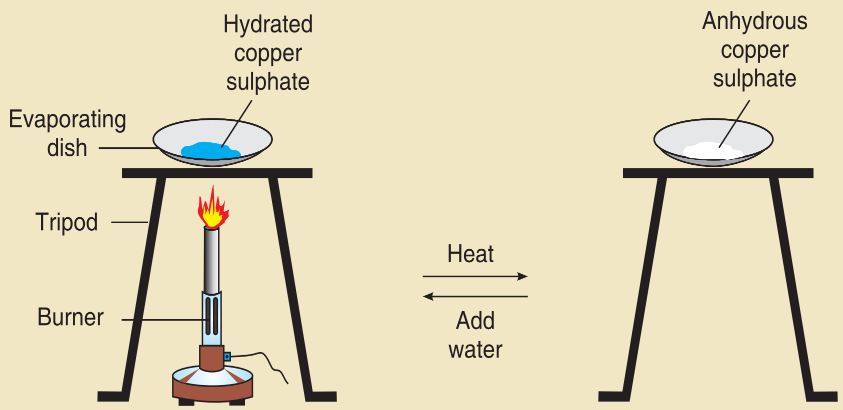
Procedure

- Measure and note the weight of evaporating dish.
- Take 2 g of copper sulphate in an evaporating dish.
- Measure and note the weight again.
- Fix all apparatus as shown in figure on page 103.
- Heat hydrated copper sulphate gently until the colour disappears.

Explanation

The heat causes the hydrated copper sulphate (blue) to split into anhydrous copper sulphate and water. Anhydrous copper sulphate is white in colour.

If we add water to the anhydrous copper sulphate, the white powder becomes blue again.



Safety

- Wear goggles, do not look directly into the evaporating dish.
- Keep away from the burner.
- Do not touch the warm dish.
- This experiment must be performed carefully in the presence of teacher.

When hydrated copper sulphate is heated, it decomposes into anhydrous copper sulphate and water.

$$\underset{\text{Hydrated copper sulphate (blue)}}{CuSO_4.5H_2O} \xrightleftharpoons{\text{Heat}} \underset{\text{Anhydrous Copper sulphate (White)}}{CuSO_4} + \underset{\text{Water}}{5H_2O}$$

In this reaction, when the blue crystals of hydrated copper (II) sulphate are heated, water molecules evaporate leaving anhydrous copper sulphate.

We have seen that the decomposition reactions require energy either in the form of heat, light or electricity for breaking down the reactants.

Note: Reactions in which energy is absorbed are known as *endothermic reactions*.

The digestion of food in the body is an example of decomposition reaction. The major constituents of our food such as carbohydrates, fats, proteins, etc. decompose to give simple glucose and amino acids. The glucose further combines with oxygen to release large amount of energy which keeps our body working.

Note: Reactions in which energy is released are known as *exothermic reactions*.

EXERCISE 5.2

1. What will you observe when 0.5 g lead nitrate is heated?
2. The colour of anhydrous copper sulphate is ______ .
3. The digestion of food in the body is an example of combination reaction. (True or False)
4. Write a balanced chemical equation of the following:
   (*a*) Silver bromide exposed to light.
   (*b*) Calcium carbonate is heated.
   (*c*) When electricity is passed through acidified water.
5. Name the catalyst which increases the rate of decomposition of H_2O_2.

5.1.3 Single Replacement Reactions

ACTIVITY 5.7: Reaction of Iron Nails with Copper Sulphate

- Take three iron nails and clean them by rubbing with sand paper.
- Take two test tubes marked as (A) and (B). In each test tube, take about 10 ml of copper sulphate solution.
- Tie two iron nails with a thread and immerse them carefully in the copper sulphate solution in test tube B for about 20 minutes [figure (*a*)]. (If the solution is diluted, leave it for 2 hours) keep one iron nail aside for comparison.
- After 20 minutes, take out the iron nails from the copper sulphate solution.
- Compare the intensity of the blue colour of copper sulphate solutions in test tubes (A) and (B), [figure (*b*)].
- Also, compare the colour of the iron nails dipped in the copper sulphate solution with the one kept aside [figure (*b*)].

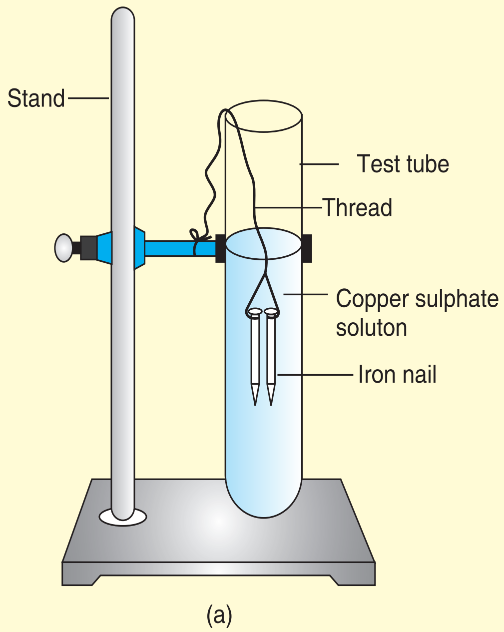
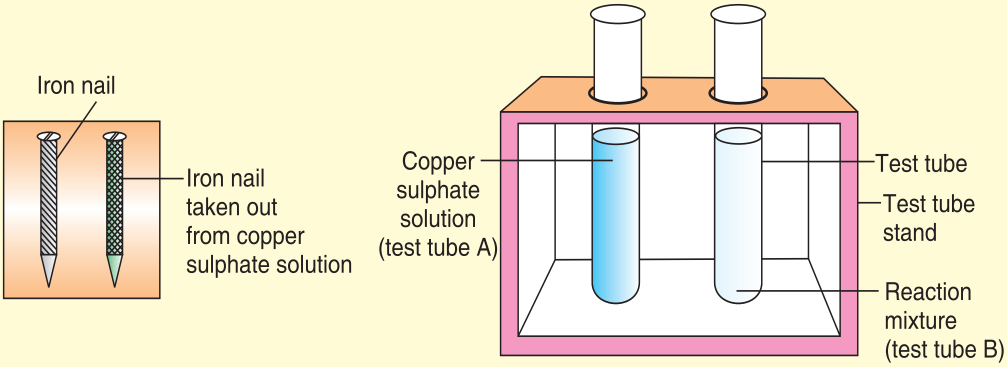


Figure. (a) Iron nails dipped in copper sulphate solution



(b)

Figure. (b) Iron nails and copper sulphate solutions compared before and after the experiment.

Single replacement reactions are those reactions in which one element displaces another element from a compound. The general equation for this type of reaction is

$$X + YZ \longrightarrow XZ + Y$$

In reactants, X is an element and YZ is a compound. In products, XZ is a compound and Y is an element.

In single replacement reactions, more active metals displace less active metals (or hydrogen) from their compounds. Some common elements are arranged in decreasing order of their ability to replace element (metal ion) in aqueous solution. This series is known as activity series.

K > Na > Ca > Mg > Al > Zn > Cr > Fe > Ni > Sn > Pb > H > Cu > Ag > **Au**

Note: Potassium (K) is the *most reactive* metal and gold (Au) is the *least reactive* metal.

Example 1

When a piece of iron metal (say, an iron nail) is placed in copper sulphate solution, then iron sulphate solution and copper metal are formed.

$$\underset{\text{(Blue solution)}}{\underset{\text{Copper sulphate}}{CuSO_4\,(aq)}} + \underset{\text{(Grey)}}{\underset{\text{Iron}}{Fe\,(s)}} \longrightarrow \underset{\text{(Greenish solution)}}{\underset{\text{Iron sulphate}}{FeSO_4\,(aq)}} + \underset{\text{(Red-brown)}}{\underset{\text{Copper}}{Cu\,(s)}}$$

In this reaction, iron displaces copper from copper sulphate solution. The deep blue colour of copper sulphate solution fades due to the formation of light green solution of iron sulphate. A red-brown coating (or layer) of copper metal is formed on the surface of iron metal (or iron nail). Note that ***this displacement reaction occurs because iron is more reactive than copper.***

Example 2

When a copper strip is placed in a solution of silver nitrate, then copper nitrate solution and silver metal are formed.

$$\underset{\substack{\text{Silver nitrate}\\ \text{(Colourless solution)}}}{2AgNO_3\,(aq)} + \underset{\substack{\text{Copper}\\ \text{(Red-brown)}}}{Cu\,(s)} \longrightarrow \underset{\substack{\text{Copper nitrate}\\ \text{(Blue solution)}}}{Cu(NO_3)_2\,(aq)} + \underset{\substack{\text{Silver}\\ \text{(Greyish white)}}}{2Ag\,(s)}$$

In this case, copper displaces silver from silver nitrate compound. **This displacement reaction occurs because copper is more reactive than silver.** A shining greyish white deposit of silver is formed on the copper strip and the solution becomes blue due to the formation of copper nitrate.

Example 3

When a strip of zinc metal is placed in copper sulphate solution, then zinc sulphate solution and copper are obtained.

$$\underset{\substack{\text{Copper sulphate}\\ \text{(Blue solution)}}}{CuSO_4\,(aq)} + \underset{\substack{\text{Zinc}\\ \text{(Silvery-white)}}}{Zn\,(s)} \longrightarrow \underset{\substack{\text{Zinc sulphate}\\ \text{(Colourless solution)}}}{ZnSO_4\,(aq)} + \underset{\substack{\text{Copper}\\ \text{(Red-brown)}}}{Cu\,(s)}$$

In this reaction, zinc displaces copper from copper sulphate compound so that copper is set free (or liberated). The blue colour of copper sulphate solution fades due to the formation of zinc sulphate (which is colourless). A red-brown deposit of copper metal is formed on the zinc strip. *Note that this displacement reaction takes place because zinc is more reactive than copper.*

Example 4

When a piece of magnesium metal is placed in copper sulphate solution, then magnesium sulphate solution and copper metal are formed.

$$\underset{\substack{\text{Copper sulphate}\\ \text{(Blue solution)}}}{CuSO_4\,(aq)} + \underset{\substack{\text{Magnesium}\\ \text{(Silvery white)}}}{Mg\,(s)} \longrightarrow \underset{\substack{\text{Magnesium sulphate}\\ \text{(Colourless solution)}}}{MgSO_4\,(aq)} + \underset{\substack{\text{Copper}\\ \text{(Red-brown)}}}{Cu\,(s)}$$

In this reaction, magnesium displaces copper from copper sulphate solution. The blue colour of copper sulphate solution fades due to the formation of colourless solution of magnesium sulphate. A red-brown deposit of copper metal is formed on the magnesium piece. Here, *magnesium is able to displace copper from copper sulphate solution because magnesium is more reactive than copper.*

Example 5

When a strip of lead metal is placed in a solution of copper nitrate, then lead nitrate solution and copper metal are formed.

$$\underset{\substack{\text{Copper nitrate}\\ \text{(Green solution)}}}{Cu(NO_3)_2\,(aq)} + \underset{\substack{\text{Lead}\\ \text{(Bluish grey)}}}{Pb\,(s)} \longrightarrow \underset{\substack{\text{Lead nitrate}\\ \text{(Colourless solution)}}}{Pb(NO_3)_2\,(aq)} + \underset{\substack{\text{Copper}\\ \text{(Red-brown)}}}{Cu\,(s)}$$

In this case, lead displaces copper from copper nitrate solution. The green colour of copper nitrate solution fades due to the formation of colourless solution of lead nitrate. A red-brown layer of copper metal is deposited on the lead strip. Please note that *lead is able to displace copper from copper nitrate solution because lead is more reactive than copper*. Another point to be noted is that copper used in this reaction is actually copper (II) nitrate.

Example 6

When copper oxide is heated with magnesium powder, then magnesium oxide and copper are formed.

$$\underset{\text{Copper oxide}}{CuO\,(s)} + \underset{\text{Magnesium}}{Mg\,(s)} \longrightarrow \underset{\text{Magnesium oxide}}{MgO\,(s)} + \underset{\text{Copper}}{Cu\,(s)}$$

This is a displacement reaction. *In this displacement reaction, a more reactive metal, magnesium, is displacing a less reactive metal, copper, from its oxide, copper oxide*.

Example 7

When iron (III) oxide is heated with aluminium powder, then aluminium oxide and iron metal are formed.

$$\underset{\substack{\text{Iron (III) oxide}\\\text{(Ferric oxide)}}}{Fe_2O_3\,(s)} + \underset{\text{Aluminium}}{2Al\,(s)} \longrightarrow \underset{\text{Aluminium oxide}}{Al_2O_3\,(s)} + \underset{\substack{\text{(Molten)}\\\text{Iron}}}{2Fe\,(l)}$$

Figure 5.1: The displacement reaction between iron (III) oxide and powdered aluminium produces so much heat that iron metal is obtained in molten form.

In this displacement reaction, a more reactive metal, (aluminium), is displacing a less reactive metal (iron), from its oxide, iron (III) oxide.

All the above examples of displacement reactions are **'single displacement reactions'**. This is because in all these reactions only 'one element' displaces 'another element' from its compound.

EXERCISE 5.3

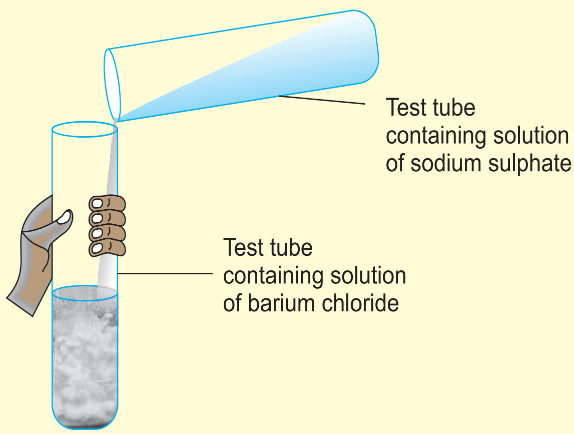
1. Copper is more reactive than silver. (True or False)
2. Why does blue colour of copper sulphate solution fade when a piece of iron object is placed into it?
3. When a copper strip is placed in a solution of zinc sulphate solution, copper sulphate is formed. (True or False)

4. Write a balanced chemical reaction of the following:
   (*a*) When copper oxide is heated with magnesium powder.
   (*b*) When iron oxide is heated with aluminium powder.
5. When a strip of lead metal is placed in a solution of ______ , lead nitrate and ______ are formed.

5.1.4 Double Displacement Reactions

ACTIVITY 5.8: Formation of Precipitate

- Take about 3 ml of sodium sulphate solution in a test tube.
- In another test tube, take about 3 ml of barium chloride solution.
- Mix the two solutions (Figure).



Mixing of sodium sulphate and barium chloride

- What do you observe?

In Activity 5.8, you will observe that a white substance, which is insoluble in water, is formed. This insoluble substance formed is known as a *precipitate*. Any reaction that produces a precipitate can be called a ***precipitation reaction*.**

Reaction

$$\underset{\text{Sodium sulphate}}{Na_2SO_4\,(aq)} + \underset{\text{Barium chloride}}{BaCl_2\,(aq)} \longrightarrow \underset{\text{Barium sulphate}}{BaSO_4\,(s)} + \underset{\text{Sodium chloride}}{2NaCl\,(aq)}$$

Double displacement reactions are those reactions in which two compounds react by exchange of ions to form two new compounds. These reactions are also known as **metathesis reactions** or **exchange reactions**.

The general equation for a double displacement reaction is

$$AX + BY \longrightarrow AY + BX$$

where A and B are positive ions (cations).

X and Y are negative ions (anions).

Another example of double displacement reaction is when we mix solutions of silver nitrate ($AgNO_3$) and sodium chloride, silver chloride and sodium nitrate is formed.

$$AgNO_3\,(aq) + NaCl\,(aq) \longrightarrow AgCl\,(s) + NaNO_3\,(aq)$$

Double displacement reactions result in the removal of ions from the solutions. The removal of ions can occur in the following two ways:

- By precipitation reaction.

Definition: When we mix two solutions both containing a soluble compound, the mixture immediately becomes cloudy by the formation of insoluble precipitate. The reaction which produces a precipitate is called precipitation reaction. Precipitation reactions can be used to test the presence of metals.

- By neutralisation reaction.

Definition: When solutions of acids and bases are mixed, the $H^+\,(aq)$ from the acid combines with the $OH^-\,(aq)$ from the base to form water. The reaction of an acid and a base is known as a **neutralisation reaction**.

Precipitation Reaction

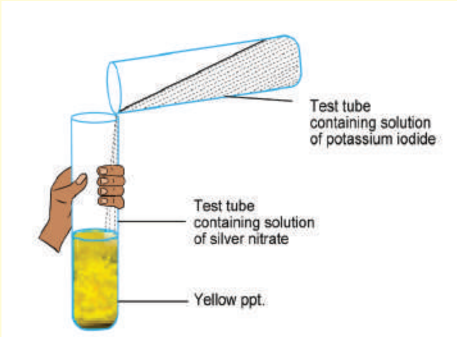
ACTIVITY 5.9: Formation of Silver Iodide

Take silver nitrate solution in a test tube and add potassium iodide solution.

What do you observe?

Explanation

You will observe that a **yellow precipitate** is formed at once. The yellow precipitate is of silver iodide which is formed as a result of double displacement reaction.



Formation of silver iodide

Reaction

$$AgNO_3\,(aq) + KI\,(aq) \longrightarrow \underset{\text{Yellow ppt.}}{AgI\,(s)} + KNO_3\,(aq)$$

Some examples of precipitation reaction:

Example 1

If barium chloride solution is added to copper sulphate solution, then a white precipitate of barium sulphate is produced along with copper chloride solution:

$$\underset{\text{Barium chloride}}{BaCl_2\,(aq)} + \underset{\text{Copper sulphate}}{CuSO_4\,(aq)} \longrightarrow \underset{\substack{\text{Barium sulphate}\\\text{(White ppt.)}}}{BaSO_4\,(s)} + \underset{\text{Copper chloride}}{CuCl_2\,(aq)}$$

In this double displacement reaction, two compounds, barium chloride and copper sulphate, react by an exchange of their ions to form two new compounds, barium sulphate and copper chloride.

Example 2

When hydrogen sulphide gas is passed through copper sulphate solution, then a black precipitate of copper sulphide is formed along with sulphuric acid solution:

$$\underset{\text{Copper sulphate}}{CuSO_4\,(aq)} + \underset{\text{Hydrogen sulphide}}{H_2S\,(g)} \longrightarrow \underset{\substack{\text{Copper sulphide}\\\text{(Black ppt.)}}}{CuS\,(s)} + \underset{\text{Sulphuric acid}}{H_2SO_4\,(aq)}$$

In this double displacement reaction, two compounds, copper sulphate and hydrogen sulphide react by an exchange of ions to form two new compounds, copper sulphide and sulphuric acid.

Example 3

When ammonium hydroxide solution is added to aluminium chloride solution, then a white precipitate of aluminium hydroxide is formed along with ammonium chloride solution:

$$\underset{\substack{\text{Aluminium}\\\text{chloride}}}{AlCl_3\,(aq)} + \underset{\substack{\text{Ammonium}\\\text{hydroxide}}}{3NH_4OH\,(aq)} \longrightarrow \underset{\substack{\text{Aluminium hydroxide}\\\text{(White ppt.)}}}{Al(OH)_3\,(s)} + \underset{\substack{\text{Ammonium}\\\text{chloride}}}{3NH_4Cl\,(aq)}$$

In this double displacement reaction, two compounds, aluminium chloride and ammonium hydroxide, react by an exchange of their ions to form two new compounds, aluminium hydroxide and ammonium chloride.

Example 4

When potassium iodide solution is added to lead nitrate solution, then a yellow precipitate of lead iodide is produced along with potassium nitrate solution:

$$\underset{\text{Lead nitrate}}{Pb(NO_3)_2\,(aq)} + \underset{\text{Potassium iodide}}{2KI\,(aq)} \longrightarrow \underset{\substack{\text{Lead iodide}\\\text{(Yellow ppt.)}}}{PbI_2\,(s)} + \underset{\substack{\text{Potassium}\\\text{nitrate}}}{2KNO_3\,(aq)}$$

In this double displacement reaction, two compounds, lead nitrate and potassium iodide, react by an exchange of ions to form two new compounds, lead iodide and potassium nitrate. Note that lead nitrate, $Pb(NO_3)_2$, is also written as lead (II) nitrate.

Neutralisation Reaction

ACTIVITY 5.10: Illustrating Neutralisation Reaction

- Take about 2 ml of dilute NaOH solution in a test tube.
- Add 2–3 drops of phenolphthalein solution.
  What is the colour of the solution?
- Add dilute HCl solution to the pink solution drop by drop.
  What do you observe?
- Now add a few drops of NaOH solution and again observe if there is any change in colour.

Observations

The colour of the solution becomes dark pink.

It is observed that when sufficient quantity of HCl solution has been added, the pink colour of the solution disappears.

In Activity 5.10, you will observe that the pink colour reappears.

Explanation

Phenolphthalein has pink colour in basic medium. When sufficiently large quantity of HCl solution is added to the NaOH solution, the medium becomes acidic due to the presence of excess acid. In acidic medium, the phenolphthalein is colourless and hence the solution becomes colourless. When we again add NaOH solution, it neutralises the excess acid and the solution again becomes alkaline. Therefore, the pink colour reappears.

In neutralisation reaction, a salt is formed from the cation of the base and the anion of the acid. For example, sodium hydroxide and hydrochloric acid react to form sodium chloride and water.

$$\underset{\text{Hydrochloric acid}}{HCl\,(aq)} + \underset{\text{Sodium hydroxide}}{NaOH\,(aq)} \longrightarrow \underset{\text{Sodium chloride}}{NaCl\,(aq)} + \underset{\text{Water}}{H_2O\,(l)}$$

In this double displacement reaction, two compounds, sodium hydroxide and hydrochloric acid, react by an exchange of ions to form two new compounds, sodium chloride and water. Note that no precipitate is formed in this double displacement reaction (This is because sodium chloride is soluble in water).

Some examples of neutralisation reactions are:

(*i*) $$\underset{\text{Base}}{NaOH\,(aq)} + \underset{\text{Acid}}{HCl\,(aq)} \longrightarrow \underset{\text{Salt}}{NaCl\,(aq)} + \underset{\text{Water}}{H_2O\,(l)}$$

(*ii*) $$\underset{\text{Base}}{KOH\,(aq)} + \underset{\text{Acid}}{HNO_3\,(aq)} \longrightarrow \underset{\text{Salt}}{KNO_3\,(aq)} + \underset{\text{Water}}{H_2O\,(l)}$$

(*iii*) $$\underset{\text{Base}}{Ba(OH)_2\,(aq)} + \underset{\text{Acid}}{2HCl\,(aq)} \longrightarrow \underset{\text{Salt}}{BaCl_2\,(aq)} + \underset{\text{Water}}{2H_2O\,(l)}$$

(*iv*) $$\underset{\text{Base}}{2NaOH\,(aq)} + \underset{\text{Acid}}{H_2SO_4\,(aq)} \longrightarrow \underset{\text{Salt}}{Na_2SO_4\,(aq)} + \underset{\text{Water}}{H_2O\,(l)}$$

Some applications of neutralisation reaction

1. A person suffering from hyper acidity is advised to take antacid tablets or antacid suspension. Antacid preparations contain magnesium hydroxide as the active component which neutralises the excess acid present in the stomach.

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

2. In acidic soils, slaked lime is added to reduce acidity.

$$\underset{\substack{\text{Slaked lime} \\ \text{(Base)}}}{Ca(OH)_2} + H_2SO_4 \longrightarrow CaSO_4 + 2H_2O$$

3. The sting of ants and bees contains formic acid. It is neutralised by rubbing soap or dilute ammonia solution.
4. The sting of yellow wasps contains an alkali. It is neutralised by rubbing dilute acetic acid (vinegar).

EXERCISE 5.4

1. The insoluble substance formed during a chemical reaction is called ______.
2. The reaction of an acid and a base is known as ______ .
3. Precipitation reaction can be used to test the presence of metals. (True or False)
4. What happens when barium chloride solution is added to copper sulphate solution?
5. Complete the following equation:

$$CuSO_4\,(aq) + H_2S\,(g) \longrightarrow ? + ?$$

6. Write the general equation for a double displacement reaction.

5.1.5 Combustion Reactions

ACTIVITY 5.11: Burning of Magnesium Ribbon

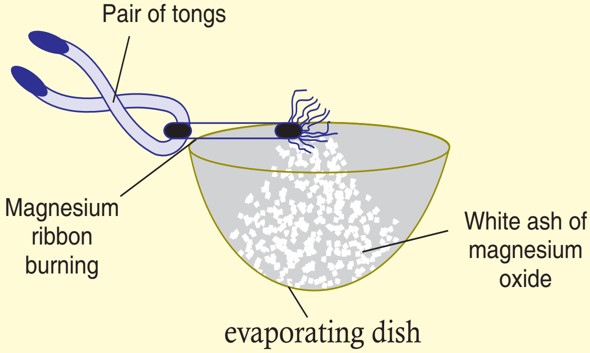
- Take a piece of magnesium ribbon and hold it with a pair of tongs. Burn magnesium ribbon.
- The magnesium ribbon starts burning with a dazzling flame.

Explanation

You would observe that magnesium ribbon soon changes into white powder. This white powdery substance is magnesium oxide which is formed as a result of combination reaction.

Reaction

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



Burning of magnesium ribbon in air

This type of reaction refers to the reaction of an element or compound with oxygen. Combustion usually releases a lot of heat energy. It is also referred to as *burning.*

Note: All combustion reactions are exothermic reactions.

Elements on Combustion

When elements undergo combustion, generally only one product is formed.

For example,

- Burning of coal

$$C\ (s) + O_2\ (g) \longrightarrow CO_2(g)$$

- Formation of water

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

- Burning of magnesium

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

Compounds on Combustion

When compounds undergo combustion, two or more products are formed. When carbon-hydrogen (hydrocarbon) or carbon-hydrogen-oxygen compounds undergo combustion in an excess of oxygen, the products are carbon dioxide and water. For example,

- Burning of natural gas

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

- Glucose reacts with oxygen

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

When any hydrocarbon burns in insufficient oxygen, the products are carbon monoxide and water. For example,

$$2C_8H_{18}(l) + 17O_2(g) \longrightarrow 16CO(g) + 18H_2O(l)$$

EXERCISE 5.5

1. All combustion reactions are exothermic. (True or False)
2. What happens when glucose reacts with oxygen?
3. When any hydrocarbon burns in insufficient oxygen, the products are ______ and _______ .
4. Combustion reaction is also referred to as ______ .
5. Complete the reaction
   $CH_4\,(g) + 2O_2\,(g) \longrightarrow ? + ? + ?$

5.2 CLASSIFICATION OF CHEMICAL REACTIONS AS ENDOTHERMIC AND EXOTHERMIC REACTIONS

During a chemical reaction, energy changes from one type to another. Energy is usually transferred either to or from the surroundings. Depending upon the evolution or absorption of energy, the chemical reactions can be classified into two types: *exothermic* and *endothermic.*

5.2.1 Exothermic Reactions

The chemical reactions which proceed with the evolution of heat energy are called exothermic reactions.

The heat energy produced during the reaction is indicated by writing $+q$ or more precisely by giving the actual numerical value along with products. In general, exothermic reactions may be represented as:

$$A + B \longrightarrow C + D + q \text{ (heat energy)}$$

The heat evolved is expressed in the units of *Joules* (J) or *kilo Joules* (kJ).

For example, when one mole of carbon (coal) is burnt in oxygen, there is an evolution of 393.5 kJ of energy. It can be expressed as:

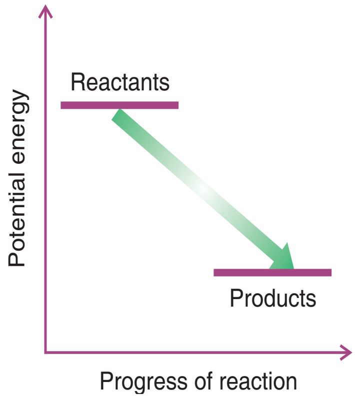
$$C(s) + O_2(g) \longrightarrow CO_2(g) + 393.5 \text{ kJ}$$

Some more examples of exothermic reactions are:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$
$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$
$$2Al\,(s) + Fe_2O_3(s) \longrightarrow Al_2O_3(s) + 2Fe(l)$$

It may be mentioned here that:

- All combustion reactions are exothermic.
- Most of the reactions are exothermic.
- During exothermic reactions, a part of the potential energy possessed by the reactants is released. Hence, products of an exothermic reaction have less potential energy than the reactants as illustrated.



Exothermic reaction

5.2.2 Endothermic Reactions

> ***The chemical reactions which proceed with the absorption of heat energy are called endothermic reactions.***

The heat energy absorbed during the reaction can be indicated by writing $+q$ (or the actual numerical value) with the reactants. It can be indicated by writing $-q$ (or the actual numerical value) with the products. In general, an endothermic reaction can be represented as:

$$A + B + q\,(\text{heat}) \longrightarrow C + D$$

$$A + B \longrightarrow C + D - q\,(\text{heat})$$

where q is the heat absorbed.

For example, formation of nitric oxide from nitrogen and oxygen proceeds with the absorption of 180.5 kJ of heat. It can be represented as:

$$N_2(g) + O_2(g) \longrightarrow 2NO(g) - 180.5\ \text{kJ}$$

Some more examples of endothermic reactions are:

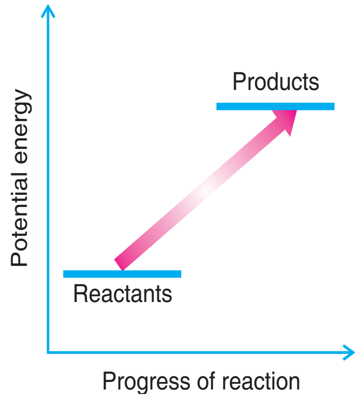
$$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$$

$$NH_4Cl(s) \longrightarrow NH_3(g) + HCl(g)$$

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

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

$$C(s) + H_2O(g) \longrightarrow CO(g) + H_2(g)$$



Endothermic reaction

It may be mentioned here that:

- Decomposition reactions are generally endothermic.
- The number of endothermic reactions is much less than the exothermic reactions.
- During endothermic reactions, reactants gain energy. Hence, products of an endothermic reaction have more potential energy than the reactants.

5.2.3 Explanation for the Energy Changes during Chemical Reactions

A chemical reaction involves the rearrangement of atoms. During the reactions certain bonds are broken while certain new bonds are formed between the atoms. Energy is absorbed for breaking the bonds and released during formation of bonds. If the energy required to break the bonds is more than the energy released during formation of bonds then there is net absorption of energy and the reaction is endothermic. On the other hand, if the energy released during formation of bonds is more than the energy absorbed for breaking the bonds, then there is net release of energy and the reaction is exothermic.

For example, let us consider the following reaction:

$$H_2 + Cl_2 \longrightarrow 2HCl$$

or $$H{-}H + Cl{-}Cl \longrightarrow 2H{-}Cl$$

In this reaction, energy is required to break H—H and Cl—Cl bonds and is released during formation of two H—Cl bonds.

Bond energy of H—H bond = 433 kJ/mole

Bond energy of Cl—Cl bond = 242 kJ/mole

Bond energy of H—Cl bond = 430 kJ/mole

Energy required to break two bonds

= Bond energy of H—H bond + Bond energy of Cl—Cl bond

= 433 + 242 = 675 kJ

Energy released

= 2 × Bond energy of H—Cl bond = 2 × 430 = 860 kJ

Since energy released during the reaction is more than the energy absorbed, there would be net release of energy and thus the reaction would be exothermic.

Net energy released = 860 – 675 = **185 kJ.**

EXPERIMENT 3

Aim

To differentiate between exothermic and endothermic reactions.

Apparatus and equipment (per group)

- Beaker
- Thermometer

Chemicals (per group)

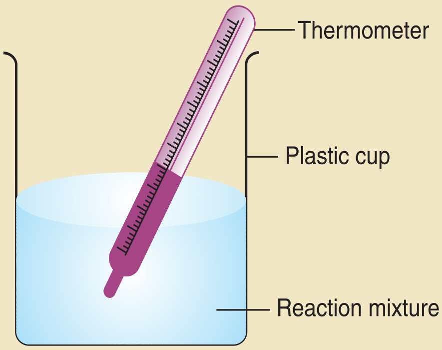
- Sodium hydroxide solution
- Dilute hydrochloric acid
- Sodium hydrogen carbonate solution
- Four spatula measures of citric acid
- Copper (II) sulphate solution
- Four spatula measures of magnesium powder (**Highly flammable**)
- 3 cm Magnesium ribbon (**Highly flammable**)
- Dilute sulphuric acid

Introduction

Some reactions give out heat and others take in heat. In exothermic reactions, the temperature goes up, in endothermic reactions the temperature goes down. In this experiment, various reactions are examined. Temperatures are measured to decide whether a particular reaction is exothermic or endothermic.

Procedure

1. Use the apparatus as shown in figure.
2. Put 10 cm^3 of sodium hydroxide solution in the beaker, record the temperature then add 10 cm^3 of dilute hydrochloric acid, stirring with the thermometer. Record the maximum or minimum temperature.
3. Repeat the procedure for the following reactions: (*a*) sodium hydrogen carbonate solution and citric acid; (*b*) copper(II) sulphate solution and magnesium powder; and (*c*) dilute sulphuric acid and magnesium ribbon.



What to record

Reaction	Temperature before mixing/°C	Temperature after mixing/°C	Exothermic or endothermic
Sodium hydroxide solution + dilute hydrochloric acid			
Sodium hydrogen carbonate solution + citric acid			
Copper(II) sulphate solution + Magnesium powder			
Dilute sulphuric acid + Magnesium ribbon			

Safety

1. Wear eye protection. Some of the solutions are irritant.
2. Use the thermometer properly to measure changes in temperature.

EXERCISE 5.6

1. Distinguish between exothermic and endothermic reactions.
2. Energy is absorbed for ______ the bonds and released during the ______ of bonds.
3. Decomposition reactions are generally exothermic. (True or False)
4. Give an example of exothermic reaction.
5. Give an example of endothermic reaction.

5.3 IONIC EQUATIONS

Most of the equations that we have studied thus far are condensed equations. In condensed equation, all reactants and products are written as electrically neutral compounds and molecules. The reactions which occur in aqueous solution can also be represented by ionic equations. Generally, all single-replacement reactions and double-displacement reactions occur in aqueous solution.

- In single replacement reaction, only cations are involved.
- In double displacement reaction, both cations and anions are involved.

Let us consider an example.

The reaction of aqueous barium chloride with aqueous sodium sulphate

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

In this reaction, three out of four components are strong electrolytes*. Therefore a better description of this reaction on the atomic or molecular level might be given by an ionic equation.

Ionic equations are written by assuming that strong electrolytes dissociate in aqueous solution into corresponding ions. When the cations and anions of a compound in solution are shown separately, the equation is known as **ionic equation**. The total ionic equation for the reaction between $BaCl_2$ and Na_2SO_4 would be written as

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

Notice that in this equation, Na^+ ions and Cl^- ions appear on both sides of the ionic equation. These ions remain unchanged during the chemical reaction. Ions that are in an identical state on both sides of ionic equations are called **spectator ions**. If spectator ions are subtracted from both sides of the equation, the remaining equation is known as the **net ionic equation**. The net ionic equation focuses only on the species that have undergone a change in the reaction. The net ionic equation for the reaction between $BaCl_2$ and Na_2SO_4 is

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

Net ionic equations contain all information necessary to understand the complete reaction.

5.4 RULES FOR WRITING IONIC EQUATIONS

The following rules should be followed in writing ionic equations

1. A well-balanced chemical molecular equation of the reaction with correct state symbols should be written.
2. A complete ionic equation is the written. This involves all ions in molten and aqueous substances.
3. Ensure that the number of individual atoms

* **Electrolytes: The compound which ionises completely in water.**

is equal on both sides of the equation.

4. The total electrical charges must also be equal on both sides of the equation.
5. Spectator ions should be cancelled out to get the net ionic equation.
6. Pure solids, liquids and gases do not form ions and their formulae appear as in molecular equation.

Example 1

Write ionic equation for the reaction of magnesium with dilute hydrochloric acid.

Solution

Rule 1:

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

Balance the equation:

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

Rule 2 and 3: Split dissolved ionic substances into separate ions:

$$Mg(s) + 2H^+(aq) + 2Cl^-(aq) \longrightarrow Mg^{2+}(aq) + 2Cl^-(aq) + H_2(g)$$

Rule 4: Cancel the spectator ions (in bold), which are the $2Cl^-$ ions on each side, to give:

$$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$$

Example 2

Write ionic equation for the reaction between copper(II) sulfate and sodium hydroxide.

Solution

Rule 1 gives:

$$CuSO_4(aq) + 2NaOH(aq) \longrightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

Write a complete ionic equation and put spectator ions in bold

$$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$$

The spectator ions are in bold type. Apply Rule 4 by crossing out the SO_4^{2-} ion on the left with the SO_4^{2-} ion on the right and the $2Na^+$ ions on the left with the $2Na^+$ ions on the right. The net ionic equation is:

$$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$$

The most difficult task in writing net ionic equations is determining which components of the reaction are strong electrolytes and should be written as ions. Strong electrolytes include soluble ionic compounds. Solubility rules for ionic compounds are given in Table 5.1. Some examples of strong acids are HCl, HBr, HI, HNO_3, $HClO_4$, and H_2SO_4, and strong bases are LiOH, NaOH, KOH, RbOH, CsOH, $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$.

Table 5.1: Solubility Rules for Ionic Compounds

Soluble Ionic Compounds	Insoluble Ionic Compounds
The Na^+, K^+ and NH_4^+ ions form *soluble ionic compounds.* Thus, NaCl, KNO_3 and $(NH_4)_2CO_3$ are *soluble ionic compounds.*	Sulphides (S^{2-}) are usually *insoluble.* Exceptions include Na_2S, K_2S, $(NH_4)_2S$, MgS, CaS, SrS and BaS.
The nitrate ion (NO_3^-) forms *soluble ionic compounds.* Thus, $Cu(NO_3)_2$ and $Fe(NO_3)_3$ are soluble.	Oxides Na_2O, K_2O, SrO and BaO react with water. CaO is slightly soluble.

The chloride (Cl^-), bromide (Br^-), and iodide (I^-) ions usually form *soluble ionic compounds.* Exceptions include ionic compounds of the Pb^{2+}, Hg_2^{2+}, Ag^+ and Cu^+ ions. $CuBr_2$ is soluble, but CuBr is not.

The sulphate ion (SO_4^{2-}) usually forms *soluble ionic compounds.* Exceptions include $BaSO_4$, $SrSO_4$ and $PbSO_4$, which are insoluble, and Ag_2SO_4, $CaSO_4$ and Hg_2SO_4, which are slightly soluble.

Hydroxides (OH^-) are usually *insoluble*. Exceptions include NaOH, KOH, $Sr(OH)_2$, and $Ba(OH)_2$, which are soluble, and $Ca(OH)_2$, which is slightly soluble.

Chromates (CrO_4^{2-}), phosphates (PO_4^{3-}), and carbonates (CO_3^{2-}) are usually *insoluble*. Exceptions include ionic compounds of the Na^+, K^+ and NH_4^+ ions, such as Na_2CrO_4, K_3PO_4 and $(NH_4)_2CO_3$.

Example 3

Write the condensed, ionic and net ionic equations when metallic copper is added to a solution of silver nitrate.

Solution

Condensed equation: $2AgNO_3(aq) + Cu(s) \longrightarrow 2Ag(s) + CuNO_3(aq)$

Ionic equation: $2Ag^+(aq) + 2NO_3^-(aq) + Cu(s) \longrightarrow 2Ag(s) + Cu^{2+}(aq) + 2NO_3^-(aq)$

Net ionic equation: $2Ag^+(aq) + Cu(s) \longrightarrow 2Ag(s) + Cu^{2+}(aq)$

Example 4

Write the condensed, ionic and net ionic equations when aqueous hydrochloric acid reacts with aqueous sodium hydroxide.

Solution

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

Ionic equation: $H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Na^+(aq) + Cl^-(aq)$

Net ionic equation: $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

Note: The net ionic equation for all reactions of strong acids with strong bases that form salt and water is $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

Example 5

Write the molecular equation (condensed), ionic equation and net ionic equation when a strip of zinc metal is dipped into copper sulphate solution.

Solution

Condensed equation: $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$

Ionic equation: $Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$

Net ionic equation: $Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Example 6

Write the condensed ionic, and net ionic equations that describe the following chemical reactions.

Aqueous solutions of sodium chromate and lead(II) nitrate react to form a yellow precipitate of lead(II) chromate and an aqueous solution of sodium nitrate.

Solution

Condensed equation: $Na_2CrO_4(aq) + Pb(NO_3)_2(aq) \longrightarrow PbCrO_4(s) + 2NaNO_3(aq)$

Ionic equation:

$$2Na^+(aq) + CrO_4^{2-}(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \longrightarrow PbCrO_4(s) + 2Na^+(aq) + 2NO_3^-(aq)$$

Net ionic equation: $CrO_4^{2-}(aq) + Pb^{2+}(aq) \longrightarrow PbCrO_4(s)$

Example 7

Write molecular, ionic and net ionic equations when aqueous silver nitrate reacts with copper(II) chloride.

Solution

Molecular equation: $2AgNO_3(aq) + CuCl_2(aq) \longrightarrow 2AgCl(s) + Cu(NO_3)_2(aq)$

Ionic equation: $2Ag^+(aq) + 2NO_3^-(aq) + Cu^{2+}(aq) + 2Cl^-(aq) \longrightarrow 2AgCl(s) + Cu^{2+}(aq) + 2NO_3^-(aq)$

Net ionic equation: $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

EXERCISE 5.7

1. The reactions which occur in aqueous solution can also be represented by ______.
2. What do you mean by net ionic equation?
3. In double displacement reaction, both cations and anions are involved. (True or False)
4. Write net ionic equation when metallic copper is added to a solution of silver nitrate.
5. Write the net ionic equation for all reactions of strong acids with strong bases that form salt and water.

5.5 SUMMARY

- A complete chemical equation represents the reactants, products and their physical states symbolically.
- In a combination reaction two or more substances combine to form a new single substance.
- Decomposition reactions are opposite to combination reactions. In a decomposition reaction, a single substance decomposes to give two or more substances.
- Decomposition reactions are usually endothermic because energy is required to break the bonds present in reactants.
- When an element displaces another element from its compound, a single displacement reaction occurs.

- Two different atoms or groups of atoms (ions) are exchanged in double displacement reactions.
- Precipitation reactions produce insoluble salts.
- The reaction between an acid and a base to give a salt and water is known as neutralisation reaction.
- Reactions in which heat is given out along with the products are called exothermic reactions.
- Reactions in which energy is absorbed are known as endothermic reactions.
- In single-replacement reaction only cations are involved.
- In double-displacement reaction both cations and anions are involved.
- When cations and anions of a compound are shown separately, in a chemical equation, the equation is called ionic equation.
- The ions that remain unchanged during the chemical reaction and present on both sides of an ionic equation is called spectator ions.
- If spectator ions are subtracted from both sides of ionic equation, the remaining equation is called net ionic equation.

5.6 GLOSSARY

- **Absorption:** a chemical or physical process by which one thing takes in or soaks up energy or a liquid or other substance from another
- **Alkaline:** having a pH greater than 7
- **Anhydrous:** containing no water especially a crystalline compound
- **Antacid:** a medicine that prevent or correct acidity; especially in the stomach
- **Burning:** on fire
- **Catalyst:** a substance that increases the rate of a chemical reaction without taking part in the reaction
- **Combustion:** the process of burning something
- **Electrolysis:** a chemical decomposition produced by passing an electric current through a liquid or solution containing ions.
- **Electrolytes:** a liquid which contains ions and can be decomposed by electrolysis
- **Endothermic:** a chemical reaction in which heat energy is absorbed
- **Exothermic:** a chemical reaction in which heat energy is released
- **Fermentation:** the chemical breakdown of a substance by bacteria or yeasts
- **Flammable:** easily set on fire
- **Hydrated:** combine chemically with water molecules, compounds containing water molecules

- **Hydrocarbon:** a compound of hydrogen and carbon
- **Immersed:** dipped or submerged in a liquid
- **Insufficient:** not enough; inadequate
- **Liberated:** : released (gas) during a chemical reaction
- **Neutralise:** to make an acidic or alkaline substance chemically neutral
- **Precipitate:** a solid substance deposited from a solution after a chemical reaction

5.7 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Which of the following is/are characteristics of chemical reactions?
   (*a*) Change in colour (*b*) Evolution of gas
   (*c*) Formation of precipitate (*d*) All of these
2. When a single product is produced from two or more reactants, the reaction is
   (*a*) Metathesis reaction (*b*) Decomposition reaction
   (*c*) Combination reaction (*d*) Displacement reaction
3. Combination reactions may involve
   (*a*) Combination of two elements (*b*) Combination of two compounds
   (*c*) Combination of one element and one compound
   (*d*) All of the above
4. Electrolysis of sodium chloride is an example of
   (*a*) Combination reactions (*b*) Decomposition reactions
   (*c*) Exchange reactions (*d*) None of these
5. Name the reaction in which energy in the form of heat, light and electricity is required to complete the reaction.
   (*a*) Combination (*b*) Decomposition
   (*c*) Single replacement (*d*) Double replacement
6. In which reaction, more active metals displace less active metals?
   (*a*) Combustion reaction (*b*) Exchange reaction
   (*c*) Single replacement reaction (*d*) Decomposition reaction
7. Which of the following statement(s) is/are not true?
   (*a*) Calcium is less reactive than copper
   (*b*) Aluminium is more reactive than sodium
   (*c*) Iron is more reactive than zinc
   (*d*) All of these

8. Choose the correct statement(s).
   (*a*) Magnesium is able to displace copper from copper sulphate
   (*b*) Zinc cannot displace copper form copper sulphate
   (*c*) Iron cannot displace copper from copper sulphate
   (*d*) Silver is able to displace copper from copper sulphate
9. When hydrocarbon compounds undergo combustion in excess of oxygen, the products are
   (*a*) Carbon dioxide and carbon monoxide (*b*) Carbon monoxide and water
   (*c*) Carbon dioxide and water (*d*) Either (*b*) or (*c*)
10. Which of the following is a net ionic equation?
   (*a*) $2AgNO_3 + Cu \longrightarrow 2Ag + Cu(NO_3)_2$
   (*b*) $2Ag^+ + 2NO_3^- + Cu \longrightarrow 2Ag + Cu^{2+} + 2NO_3^{2-}$
   (*c*) $2Ag+ + Cu \longrightarrow 2Ag + Cu^{2+}$
   (*d*) All of these

II. Open Ended Questions

1. Translate the following statements into chemical equations and balance them.
   (*a*) Hydrogen gas combines with chlorine gas to form hydrogen chloride.
   (*b*) Calcium oxide reacts with carbon dioxide to form calcium carbonate.
   (*c*) Ammonia reacts with hydrogen chloride to form ammonium chloride.
2. Identify the type of reaction in each case.
   (*a*) Hydrogen gas combines with nitrogen to give ammonia.
   (*b*) On heating, calcium carbonate breaks up into calcium oxide and carbon dioxide.
   (*c*) Hydrochloric acid reacts with sodium hydroxide to give salt and water.
3. Complete the following reactions:
   (*a*) $CuSO_4(aq) + H_2S(g) \longrightarrow ? + H_2SO_4(aq)$
   (*b*) $AgBr(s) \xrightarrow{\text{Sunlight}} Ag(s) + ?$
   (*c*) $H_2O(l) \xrightarrow{\text{Electricity}} ? + ?$
   (*d*) $3KClO_3 \xrightarrow{?} 2KCl(s) + ?$
   (*e*) $CuSO_4 \cdot xH_2O \xrightleftharpoons{\text{Heat}} ? + ?$
   (*f*) $2AgNO_3(aq) + ? \longrightarrow ? + 2Ag$
   (*g*) $CuSO_4(aq) + ? \longrightarrow MgSO_4(aq) + ?$
   (*h*) $? + NaCl(aq) \longrightarrow AgCl + ?$
   (*i*) $NaOH(aq) + HCl(aq) \longrightarrow ? + ?$
   (*j*) $NaOH(aq) + ? \longrightarrow Na_2SO_4(aq) + ?$
4. What do you mean by exothermic and endothermic reaction?
5. A shining brown coloured element X on heating in air becomes black. Name the element X and the black coloured compound.
6. State the important use of decomposition reaction.
7. What happens when
   (*a*) a piece of iron metal is placed in copper sulphate solution?
   (*b*) a strip of copper is dipped in a solution of silver nitrate?

(c) copper oxide is heated with magnesium?

(d) iron powder is heated with sulphur?

(e) silver metal is placed in copper sulphate solution?

8. A colourless lead salt produces a yellow residue and brown fumes when heated.

(a) Name the brown fumes (b) Name the yellow compound

(c) Name the lead salt

9. How can you say that "A decomposition reaction is opposite of a combination reaction"?

10. Distinguish between exothermic and endothermic reactions. Write a general equation for each of the reactions.

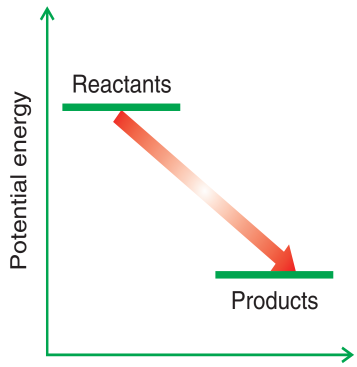
11. Write net ionic equations for the reaction of

(a) magnesium with dilute hydrochloric acid

(b) zinc with copper sulphate

(c) aqueous hydrochloric acid and aqueous sodium hydroxide

12.



The graph shows _______ reaction.

III. Practical-based Questions

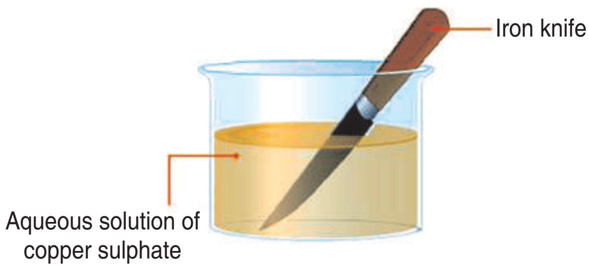
1. Which statement is true for the following reaction?

$$A + BC \rightarrow B + AC$$

(a) A is less reactive than B (b) A is more reactive than C

(c) A is more reactive than B (d) B is more reactive than A

2. What happens when an iron knife is dipped into aqueous solution of copper sulphate?



(*a*) No reaction takes place

(*b*) Fe^{2+} ions are produced in the solution and a brown solid of metallic copper formed

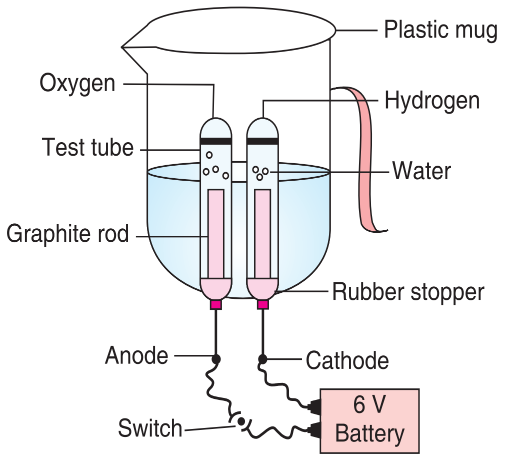
(*c*) All options are wrong

3. The following reaction is an example of:

$$AB + CD \rightarrow CB + AD$$

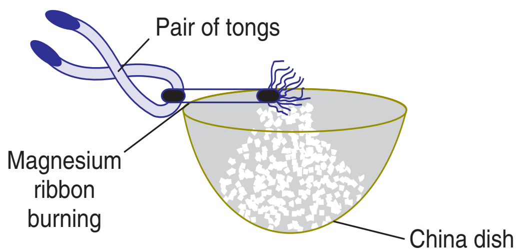
(*a*) Combination reaction
(*b*) Replacement reaction
(*c*) Double displacement reaction
(*d*) Decomposition reaction

4. The following figure illustrates



(*a*) Electrolysis of Sodium chloride
(*b*) Electrolysis of Water
(*c*) Electrolysis of Aluminium
(*d*) Electrolysis of Copper

5. When magnesium ribbon burns, the product formed is



(*a*) Acidic
(*b*) Basic
(*c*) Amphoteric
(*d*) Neutral

Unit 6

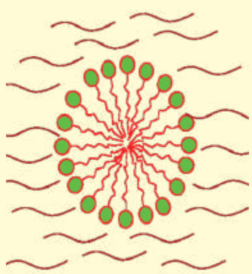
Preparation of Salts and Identification of Ions

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- define solubility.
- describe factors that affect solubility.
- explain the concept of unsaturated, saturated and supersaturated solutions.
- explain the solubility curves of different salt solutions.
- describe different methods of preparing soluble and insoluble salts.
- name the sources and uses of salts in daily life.

KNOWLEDGE GAIN



Liquid crystals are matter that has properties between those of liquid and solid crystal. Liquid crystals can be found naturally. For example, many proteins and cell membranes are liquid crystals. Other well-known example of liquid crystals is solution of soap.

ACTIVITY 6.1: Demonstrating Unsaturated, Saturated and Supersaturated Solutions

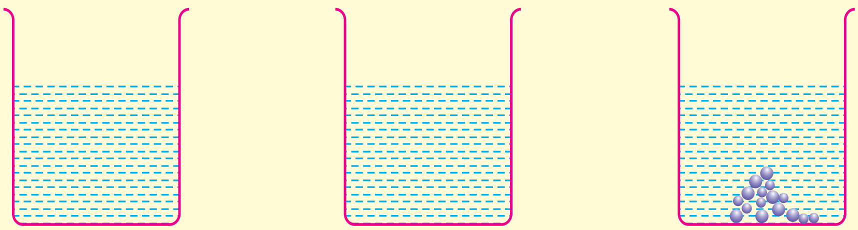
Materials Required

Three hard glass beakers (500 ml), sugar, water, and spoon.

Procedure

- Fill all three beakers with equal amount of water.
- Label them A, B and C.
- Add one teaspoon sugar in beaker A, two teaspoons sugar in beaker B, and three teaspoons in beaker C.
- Stir with spoon.

- Observe the beakers.

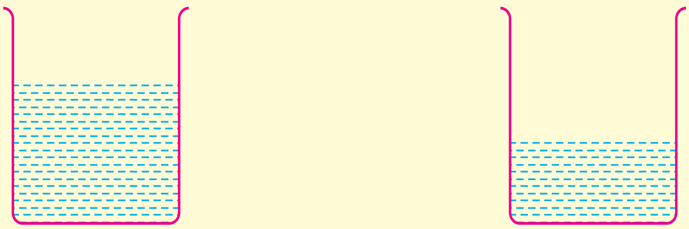


A — **After stirring,** 1 teaspoon sugar **dissolves** completely

B — **After stirring,** 2 teaspoons sugar **dissolves** completely

C — **After stirring,** 3 teaspoons sugar does not **dissolve** completely

- Heat the beaker C and observe the beaker.
- Let beaker C sit uncovered for some days to evaporate 1/2 of water.



C — **After heating,** the sugar dissolves completely.

After 4–5 days, beaker contains half-full water and 3-teaspoons sugar dissolved.

- Take beaker B and hang a string inside it.
- Leave beaker B for 45–60 minutes undisturbed.
- Observe the beaker B, after 45–60 minutes.
- Cover the beaker B to avoid further evaporation and let it sit for some days.
- Observe beaker B again.

  Did you see crystals?

In Activity 6.1,

- Solution in Beaker A is unsaturated.
- Solution in Beaker B is unsaturated.
- Solution in beaker C is saturated.

After heating, sugar left in beaker C also dissolves completely, so the solution in beaker C is unsaturated.

After evaporation ½ of water, beaker C contains 3 teaspoons of dissolved sugar. Now the solution in beaker C is supersaturated.

Definition

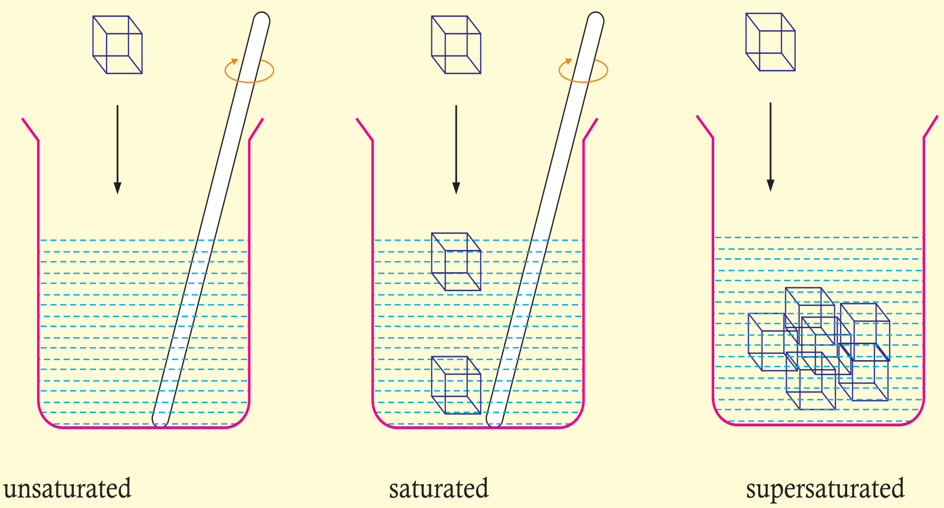
- **Solubility:** The amount of substance that can dissolve in specific amount of solvent (usually water) at a particular temperature.
- **Unsaturated solution:** A solution that contains less solute than that it can dissolve at a given temperature.
- **Saturated solution:** A solution that contains as much solute as can dissolve in the given solvent at a given temperature is a saturated solution.
- **Supersaturated solution:** A solution that contains the maximum amount of solute at an elevated temperature is a supersaturated.

6.1 SATURATED AND UNSATURATED SOLUTIONS

ACTIVITY 6.2: Testing Whether a Given Solution of a Substance is Unsaturated, Saturated, or Supersaturated, at a Particular Temperature

Prepare an unsaturated, saturated and a supersaturated solution of common salt separately in three beakers. Now, add a small amount of the crystal of common salt turn by turn to each of the solution.

- If it dissolves and the concentration of the solution increases, it is unsaturated.
- If it does not dissolve even on vigorous stirring and the concentration of the solution remains the same, it is saturated.
- If it grows in size and the concentration of the solution falls, it is supersaturated.



If a well-powdered solute, such as copper sulphate is added little by little to a definite volume of water at room temperature with constant stirring, it dissolves continuously, till a point comes, when no more of the copper sulphate will dissolve. The excess of the copper sulphate settles down to the bottom. The solution at this stage is said to be a saturated solution at that temperature. Thus, a saturated solution at a particular temperature contains as much solute as it can dissolve at that temperature.

At that temperature, solvent and solute are in a state of equilibrium. This state of equilibrium or saturation can be changed either by adding more water (solvent) to it or by increasing the temperature of solution. For example, the saturated solution at a given temperature becomes unsaturated when heated because more of solute will be required to make the solution saturated at higher temperature. Therefore, different amounts of the same solute are required to prepare saturated solutions at different temperatures. In other words, a solution saturated at one temperature may not be saturated at another temperature.

A solution which is unable to dissolve any more of the solute at a particular temperature is called a saturated solution at that temperature. At every step, before obtaining a saturated solution, a solution is always unsaturated with respect to the solute. An unsaturated solution contains less solute than it can dissolve at that temperature. It becomes more unsaturated with the rise of temperature.

A solution which can dissolve more solute at a given temperature is called unsaturated.

Table 6.1: Differences between Saturated and Unsaturated Solutions

Unsaturated Solution	Saturated Solution
1. An unsaturated solution contains as less solute as it can dissolve at that temperature. That is, it can dissolve more.	A saturated solution contains as much solute as it can dissolve at that temperature.
2. With the increase in temperature, it becomes more unsaturated.	With the increase in temperature it becomes unsaturated.

EXERCISE 6.1

1. How can state of equilibrium be changed?
2. A solution which can dissolve more solute at a given temperature is called unsaturated. (True or False)
3. In your own words, define saturated solution.
4. An unsaturated solution contains ____________ solute than it can dissolve at that temperature.
5. Distinguish between saturated and unsaturated solutions.

6.2 SUPERSATURATED SOLUTIONS

A saturated solution prepared at a higher temperature contains more of a given solute than does a saturated solution prepared at a lower temperature. So, when a saturated solution prepared at a higher temperature is cooled to room temperature, it will tend to throw out the excess of the solid from the solution in the form of crystals. It takes, however, sometime for the excess solute to come out. During this interval, solution holds in it more solute than is required to saturate it. Such a solution is called a supersaturated solution. Thus, a supersaturated solution is one which has more of the solute than a saturated solution requires.

The supersaturated solution in contact with the solid solute starts depositing the solute so as to reach the concentration of a saturated solution. Thus, in a supersaturated solution the concentration of the solute falls when it comes in contact with the solute.

A solution in which more solute is dissolved by increasing the temperature of a saturated solution is called supersaturated solution.

6.2.1 Dilute and Concentrated Solutions

ACTIVITY 6.3: Identifying Dilute and Concentrated Solution

Take equal amount of water in two beakers. Put a spoonful of sugar in one beaker and three teaspoons in the other. Stir with the glass rod.

Can you tell which solution is more sweet?

The solution containing three teaspoonful of sugar is more sweeter. This solution is concentrated whereas another one is dilute.

The quantity of solute dissolved in a certain quantity of solvent, by weight or volume denotes the concentration of the solution. Accordingly two types of solutions are known – dilute and concentrated.

A solution containing relatively small amount of solute in a fixed amount of solvent or compared to that of the solvent is a dilute solution. The solution made by mixing a teaspoon of sugar in a cup of water is a dilute solution.

Solution containing relatively more quantity or large amount of solute in the fixed amount of solvent is a concentrated solution. The solution made by mixing three teaspoons of sugar in a cup of water is a concentrated solution.

EXERCISE 6.2

1. What do you mean by supersaturated solution?
2. How can you distinguish between dilute and concentrated solutions?
3. Explain the formation of crystals.

6.3 FACTORS INFLUENCING SOLUBILITY OF DIFFERENT SALTS

The solubility of solid (solutes) in liquid (solvent) depends on

(*i*) **Nature of solute.** Dissolution of solid solutes in liquids can be summed up in a phrase "**like dissolves like**".

Some solutes such as NaCl, KCl, KNO_3, etc., have larger solubilities in water. On the other hand, solids such as I_2 and S_8 are not soluble in water they dissolve in CCl_4, CS_2.

(*ii*) **Temperature.** Temperature has a direct effect on solubility. The solubility of most of the ionic compounds increases with increase in temperature. On the other hand, some compounds dissolve better by decreasing the temperature.

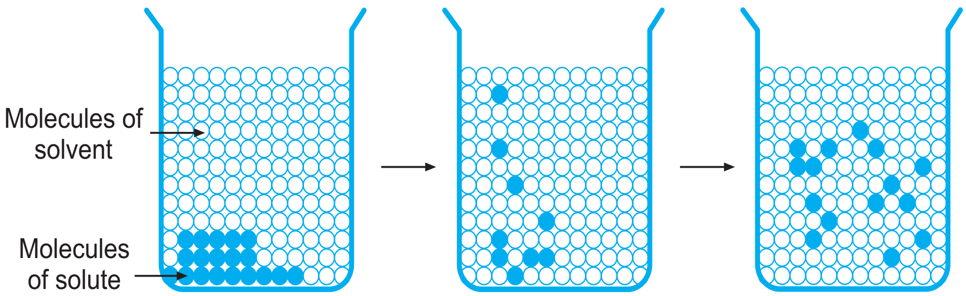
Note: The solubility of solids and liquids increases with increase in temperature.
The solubility of gases always decreases with increase in temperature.

Temperature can also increase the amount of solute that can be dissolved in a solvent. Generally speaking, as the temperature increases, more solute particles will dissolve.

For example, when you add potassium chloride (KCl) to water, a solution is easily made. When you heat this solution and keep adding KCl, you will find that a large amount of KCl can be dissolved as the temperature keeps rising. This occurs because as the temperature increases the attraction forces holding ions together in the KCl can be easily broken and allow more solute to dissolve.

6.3.1 How does Solubility Change with Temperature?

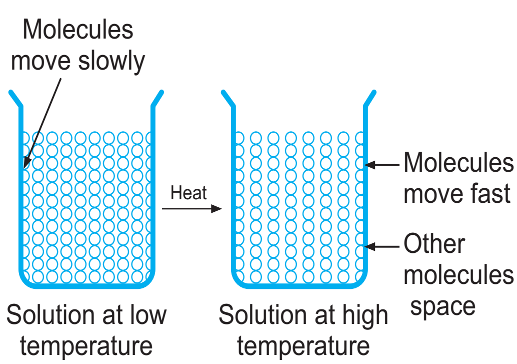
Molecules in liquid are less closely packed and the forces of attraction between the molecules are weaker than those in solids. The molecules of the liquids are constantly moving in different directions with different speeds. Thus, on heating, the kinetic energy of the molecules will increase more and move faster.



Molecules of solute get separated from each other on heating

Molecules of the solid solute are held together by strong forces of attraction. So the molecules in solids are closely packed. When the temperature is increased, the kinetic energy of the molecules will increase which makes the molecules to move fast.

Similarly, stirring the solution with a spoon or a glass rod helps to increase the kinetic energy of the molecules of solvent. Thus, on heating or stirring, the kinetic energy of the molecules of the solute increases. The molecules of the solute strike each other and get separated from each other. These separated solute molecules get mixed with the solvent molecules to form a solution.



When the molecules vibrate more at high temperature, the force of attraction between the molecules weakens and the space between them increases. That's why warm water dissolves more solute than cold water. Thus, the solubility of the substances increases with the rise of temperature.

Thus, the solubility of any solid in a solvent depends on temperature. The solubilities of most solutes like potassium nitrate, copper sulphate, ammonium chloride etc. increase with increase in temperature. The solubilities of sodium chloride and potassium chloride etc. increase slightly with an increase in temperature. Solubilities of some substances in water such as calcium sulphate, calcium hydroxide and sodium sulphate decrease with the rise of the temperature.

EXERCISE 6.3

1. Complete the phrase:
   Like dissolves __________.
2. I_2 and S_8 are not soluble in __________ but they are soluble in __________.
3. The solubility of gases always increases with increase in temperature. (True or False)
4. What are the factors that affect solubility?
5. The solubility of calcium sulphate in water

(*a*) decreases with rise of temperature
(*b*) increases with rise of temperature
(*c*) cannot be determined
(*d*) first increases and then decreases with rise of temperature.

6.4 SOLUBILITY CURVE

The variation in the solubility of any given substance with change of temperature is shown by solubility curve. *The graph showing relationship between temperature and solubility of the substance at different temperatures is called a solubility curve*.

To draw solubility curves, temperature is represented along the X-axis and solubility along the Y-axis. Various solubility points plotted are connected by a smooth curve which is a solubility curve.

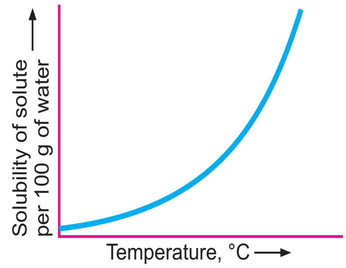


Figure 6.1: The solubility curve of a compound (say KNO_3) at various temperature.

The solubility of copper sulphate at different temperatures is given in Table 6.2.

Table 6.2: Solubility of Copper Sulphate at Different Temperatures

Temperature 0°C	Solubility
0	14
10	17
20	21
30	24
40	29
50	34
60	40
70	47

The solubility curve drawn for table 6.2 data is shown below.

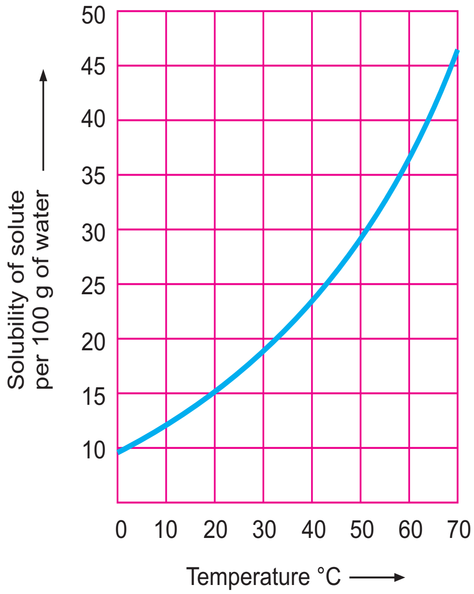


Figure 6.2: Solubility curve of copper sulphate.

The solubility curves of some substances is shown in Figure 6.3.

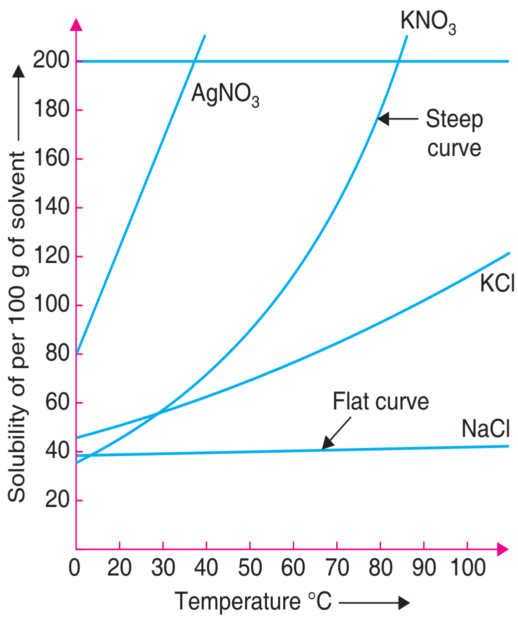


Figure 6.3: Solubility curve of some substances.

The solubility of sodium chloride rises a little with rise of temperature. The solubility of lead nitrate, potassium nitrate and sodium nitrate increases rapidly with increase of temperature.

The solubility of calcium sulphate and calcium hydroxide decreases with rise of temperature.

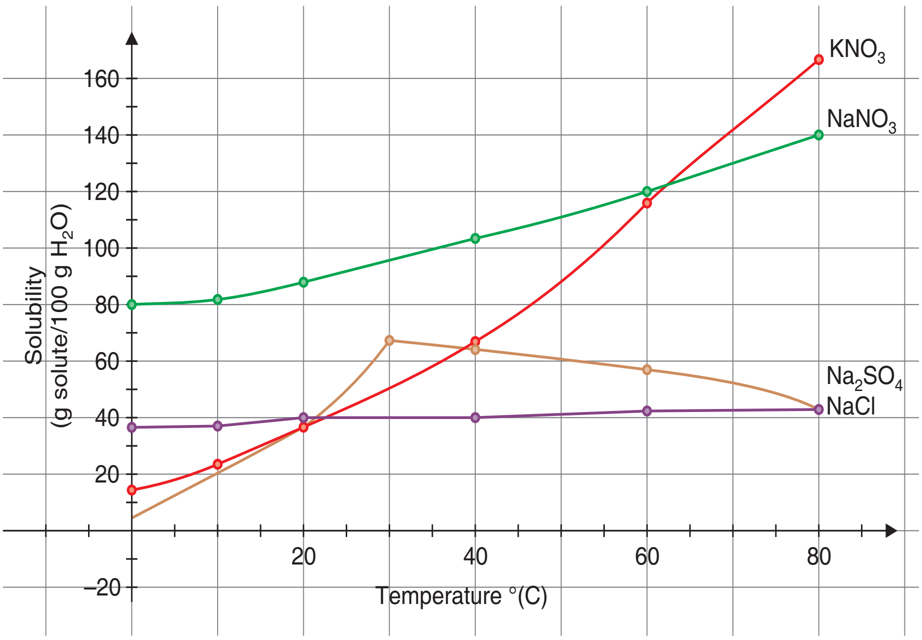
The general shape of the curve indicates the rate of change in the solubility with a rise in temperature. A steep curve, e.g., that of potassium nitrate shows that solubility increases rapidly with rise of temperature. A flat curve, e.g., that of sodium chloride indicates that solubility increases slowly with the rise of temperature.

After studying the solubility curve, the following information can be obtained.

1. The solubility of a substance at a particular temperature can be determined.
2. The solubility of a given substance at any temperature can be determined.
3. The solubility curve helps us predict which substance will crystallise out first from a solution containing two or more solutes.
4. The solubility curve helps us compare the solubilities of different substances at the same temperature.
5. It brings the change in the composition of a solute substance.
6. It gives a clear idea that solubility of substance changes with the temperature.

Example 1

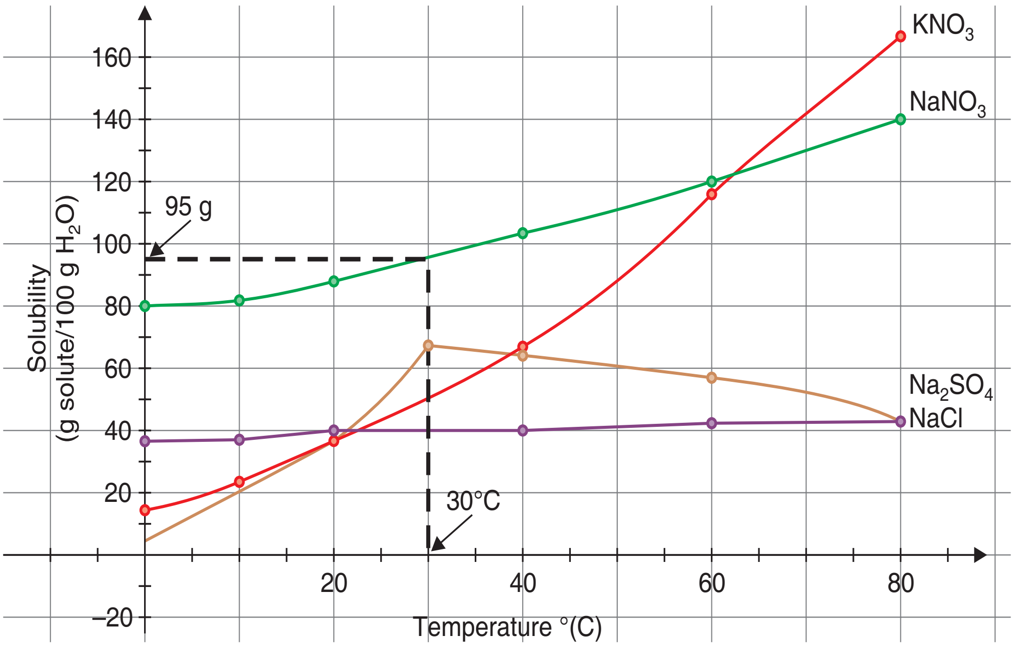
Answer the following questions using the solubility graph below.



1. *How much sodium nitrate will dissolve at 30°C?*
2. *Which solid is most soluble at 60°C?*
3. *Which solid is least soluble at 40°C?*
4. *At what temperature will 60 g of sodium sulphate dissolve in 100 g of water?*

Solution

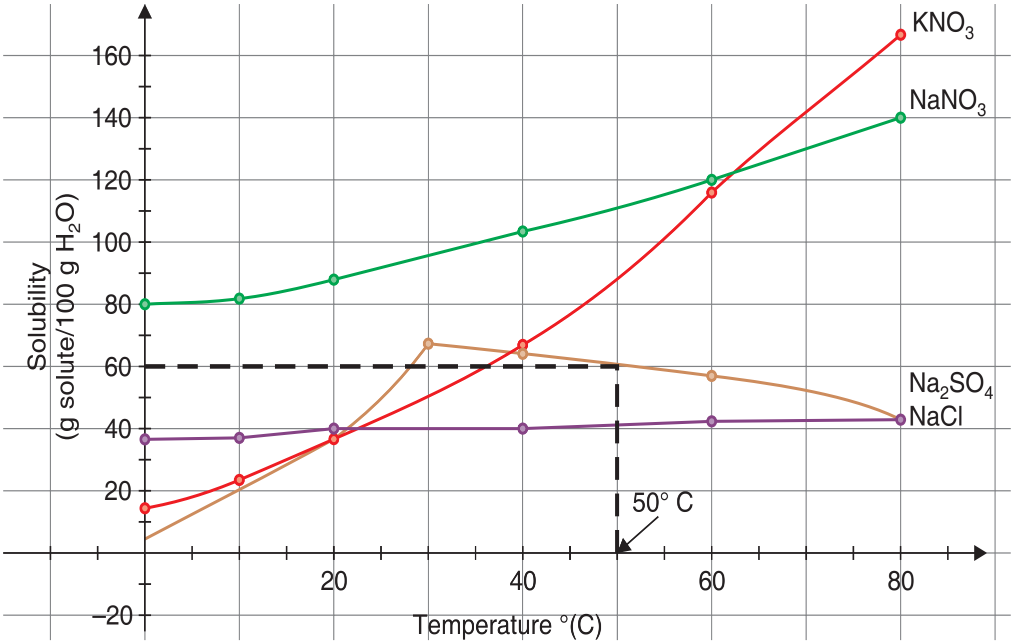
1. Looking at the following solubility graph, draw a line up (vertically) from 30°C until it touches the $NaNO_3$ line. Following this, locate the point on Y-axis to find the amount of $NaNO_3$ that dissolves. Therefore, approximately 95 g of $NaNO_3$ will dissolve in 100 g of water at 30°C.



2. The highest line at 60°C is the green line ($NaNO_3$), therefore it is the most soluble at 60°C.

3. The lowest line at 40°C is the purple line (NaCl), therefore NaCl is the least soluble at 40°C.

4. Looking at the solubility graph on page 137, draw a line over (horizontally) from 60 g until it touches the Na_2SO_4 line. Following this, locate the point on X-axis to find the temperature at which 60 g of Na_2SO_4 will dissolve. Therefore, 60 g of Na_2SO_4 will dissolve in 100 g of water at 50°C.



6.5 CALCULATION OF SOLUBILITY

ACTIVITY 6.4: Determining the Solubilities of Potassium Sulphate, Sodium Chloride and Sodium Nitrate at Room Temperature

- Take 100 grams of water in each of three separate beakers at room temperature.
- Take the weights of all the three beakers with water separately.
- Add potassium sulphate to the first, common salt to the second and potassium nitrate to the third beaker.
- Add and stir well to dissolve the required solute to make each solution saturated.
- Take the weight of each beaker containing saturated solution.
- Calculate the amount of solute added in each beaker containing 100 grams of water to make it a saturated solution at room temperature.

Calculation

Weight of the solvent in each beaker = x grams

Weight of the saturated solution in each beaker = y grams

$\therefore$ The weight of the solute = $(y - x)$ grams.

In other words, the difference in the weight of the two is the amount of a solute required to form a saturated solution at room temperature. If the above activity is done properly, following result will be obtained.

Solute	Amount of solvent (water)	Amount of solute	Solubility
Potassium sulphate	100 grams	12 grams	12
Sodium chloride	100 grams	35 grams	35
Potassium nitrate	100 grams	60 grams	60
	At room temperature		

The result of this activity shows that for different substances, the amount of solute needed to make a saturated solution in the same solvent is different at a particular temperature.

The amount of solute (in grams) required to form a saturated solution in 100 grams of solvent (water) at a particular temperature is called the solubility of the substance at that temperature.

The above activity shows that solubility of sodium chloride is 35 at 20°C, that of potassium nitrate is 60 at 20°C and that of potassium sulphate is 12 at 20°C.

Here,

$$\text{Solubility} = \frac{\text{Weight of solute (in gram)}}{\text{Weight of solvent (in gram)}} \times 100 \quad \text{(at a particular temperature)}$$

Different substances have different solubilities. Solubilities of some common substances in water at 20°C and 30°C are given in table below.

Solubilities of Some Common Substances

Solute	Solubility	
	at 20°C	at 30°C
Copper sulphate	21	25
Sodium nitrate	88	95
Sodium chloride	35	37

Above table shows clearly that the solubility of the solute increases with the rise in temperature, but the solubility of sodium chloride increases slightly with the rise in temperature.

6.5.1 Solved Numericals

Example 1

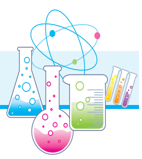
The solubility of a solute at 30°C is 40. What amount of water is required to make saturated solution of 80 grams of a solute?

Solution

Weight of a solute = 80 grams

Solubility at 30°C = 40

Weight of solvent (water) = ?



According to the formula,

Solubility

$$= \frac{\text{Weight of Solute (in grams)}}{\text{Weight of Solvent (in grams)}} \times 100$$

$\therefore$ Weight of solvent (grams)

$$= \frac{\text{Weight of Solute (in grams)}}{\text{Solubility}} \times 100$$

$$= \frac{80}{40} \times 100 = 200 \text{ grams}$$

$\therefore$ 80 grams of solute needs 200 grams of solvent (water) to form saturated solution at 30°C.

Example 2

7 grams of saturated solution of salt saturated at 60°C is evaporated to dryness; 2 grams of white residue is left behind. What is the solubility of salt at that temperature?

Solution

Weight of saturated solution = 7 grams

Weight of solute (Salt) = 2 grams

Solubility of salt at 60°C = ?

To find the weight of solvent (water),

Weight of Solution

= Weight of Solute + Weight of Solvent

Weight of Solvent

= Weight of Solution – Weight of Solute

Weight of Solvent

=7 grams – 2 grams = 5 grams

To find solubility, according to the formula,

Solubility

$$= \frac{\text{Weight of Solute (in grams)}}{\text{Weight of Solvent (in grams)}} \times 100$$

$$= \frac{2}{5} \times 100 = 40$$

$\therefore$ The solubility of salt at 60°C is 40.

Example 3

At 30°C, 7 grams of sugar dissolves in 5 grams of water to form a saturated solution. Find the solubility of sugar?

Solution

Given, Mass of solute (sugar) = 7 grams

Mass of solvent (water) = 5 grams

Solubility at 30°C = ?

According to the formula,

Solubility

$$= \frac{\text{Weight of Solute (in grams)}}{\text{Weight of Solvent (in grams)}} \times 100$$

$$= \frac{7 \, grams}{5 \, grams} \times 100$$

$$= 140$$

$\therefore$ The solubility of sugar at 30°C is 140.

Example 4

How much copper sulphate will be dissolved in 30 grams of its saturated solution. Solubility of the salt at 30°C is 35. Also find the weight of solvent.

Solution

Weight of saturated solution of copper sulphate = 30 grams

Solubility of copper sulphate at 30°C = 35

Let the amount of solute (copper sulphate) be x grams

Amount of solvent = $(30 - x)$ grams

Solubility of copper sulphate = 35

According to the formula,

$$\text{Solubility} = \frac{\text{Weight of Solute (in grams)}}{\text{Weight of Solvent (in grams)}} \times 100$$

or, $$35 = \frac{x}{(30 - x)} \times 100$$

or, $$35\,(30 - x) = 100x$$

or, $$1050 - 35x = 100x$$

or, $1050 = 135x$

or, $x = \frac{1050}{135} = 7.78$ grams

Therefore, weight of solvent in 30 grams of saturated solution at 30°C

= Weight of saturated solution – Weight of solute (x)

= 30 grams –7.78 grams

= 22.22 grams

6.6 DIFFERENT WAYS OF PREPARING NORMAL SALTS

6.6.1 Preparing Salts from Reaction of Acid with Active Metal

ACTIVITY 6.5: Preparation of Zinc sulphate $ZnSO_4.7H_2O$ (White Vitriol)

- Take about 40 cm^3 of dilute H_2SO_4 in a 250 cm^3 beaker and add about 10 g of zinc granules to the acid. Zinc reacts with dilute H_2SO_4 and hydrogen gas is evolved.
- When the evolution of gas ceases, it indicates that the reaction is over. Filter the solution to remove unreacted zinc.
- The solution thus obtained is heated in a china dish and concentrated to crystallisation point.
- Cool the solution. On cooling, needle shaped crystals of white vitriol ($ZnSO_4.7H_2O$) are obtained. The crystals are separated and dried between the folds of a paper.

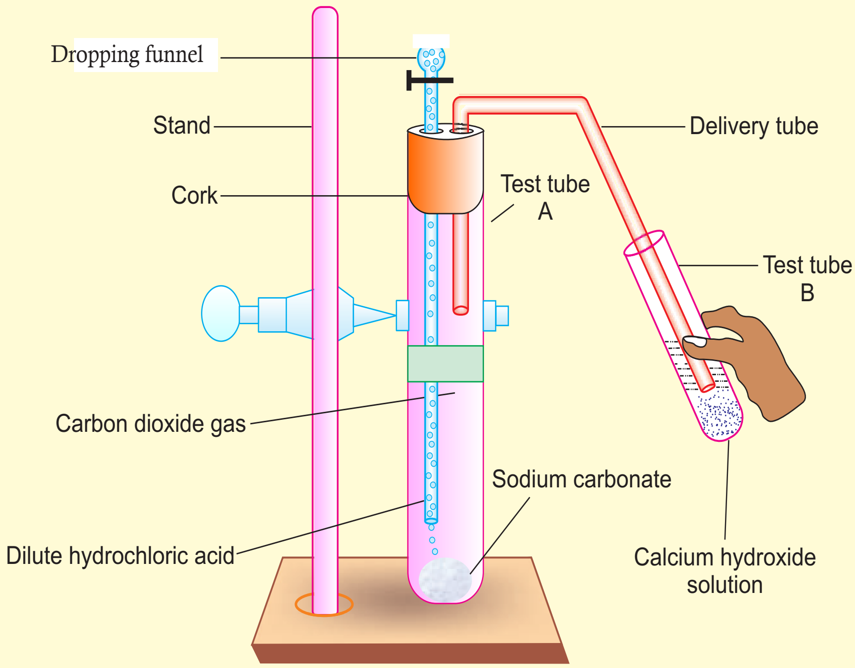
Zinc sulphate is a water soluble salt. It can be prepared in the laboratory by the action of dilute H_2SO_4 on zinc granules.

$$Zn + \text{dil. } H_2SO_4 \rightarrow ZnSO_4 + H_2 \uparrow$$

6.6.2 Preparing Salts from Reaction of Acid with Carbonates and Hydrogen Carbonate

ACTIVITY 6.6: Preparation of Sodium Chloride

- Take two test tubes, label them A and B.
- Take about 0.5 g sodium carbonate (Na_2CO_3) in test tube A and 10 ml of calcium hydroxide ($Ca(OH)_2$) in test tube B.
- Add about 2 ml of dilute HCl to the test tube A.
- What do you observe?
- Repeat the experiment with a hydrogen carbonate($NaHCO_3$)



Passing carbon dioxide gas through calcium hydroxide solution.

The reactions occurring in Activity 6.6 are written as:

Test tube A:

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

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

Test tube B:

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

EXPERIMENT 1

Aim

To prepare calcium chloride by the action of an acid on an insoluble carbonate

Materials Required

- measuring cylinder
- $250\ cm^3$ beaker (2)
- filter funnel
- conical flask
- hydrochloric acid ($50\ cm^3$)
- calcium carbonate powder (6 g)
- filter paper
- evaporating dish

- Bunsen burner
- glass spatula
- tripod stand
- wire gauze
- plastic vial
- electronic top pan balance

Procedure

- Take 50 cm^3 dilute hydrochloric acid in a beaker.
- Using an electronic top pan balance, weigh 6.0 g of calcium carbonate in a small plastic bottle.
- Add a spatula of calcium carbonate from the plastic bottle into the acid in the beaker. Stir the mixture until all the calcium carbonate has dissolved before adding more calcium carbonate.
- Continue to add more calcium carbonate to the acid until no more can dissolve and some of it is left at the bottom of the beaker.
- Filter off the excess calcium carbonate with a piece of filter paper and a funnel.
- Collect the filtrate and place it into an evaporating dish. Obtain crystals of calcium chloride from the solution by evaporation followed by crystallisation.

Conclusion

Calcium chloride is a salt of calcium and chlorine. Anhydrous calcium chloride is solid white powder at room temperature.

6.6.3 Preparation of Salts from Reaction of Acids with Metal Oxide

ACTIVITY 6.7: Preparation of Copper Sulphate $CuSO_4 . 5H_2O$ (Blue Vitriol)

- Take about 50 cm^3 of dilute H_2SO_4 in a 250 cm^3 beaker and warm it gently.
- Add black copper (II) oxide to the above acid in small portions. Initially, Copper (II) oxide reacts with the acid and dissolves. When no more of the oxide dissolves, the solution is filtered to remove unreacted copper (II) oxide.
- The bluish filtrate is concentrated by heating in a evaporation dish. When the hot saturated solution is obtained, the solution is allowed to cool undisturbed.
- On cooling, blue crystals of blue vitriol ($CuSO_4.5H_2O$) are formed. The crystals are separated and dried between the folds of a filter paper.

Copper sulphate can be prepared by the reaction of dilute H_2SO_4 with copper (II) oxide

$$CuO + H_2SO_4 \longrightarrow CuSO_4 + H_2O$$

In this activity, you will observe that black copper oxide (CuO) dissolves in dilute sulphuric acid (H_2SO_4) to form a blue copper sulphate solution.

On heating the copper sulphate solution, a saturated solution is obtained. On cooling the saturated solution, we get the crystals of copper sulphate.

> **Note:** All metal oxides react with acids but there is not always a colour change. Magnesium Oxide is a white powder. The salt produced when it reacts with acid is also white.

6.6.4 Preparing Salts from Reaction of Acids with Bases

ACTIVITY 6.8: Preparation of Sodium Sulphate

- Take about 2 ml of dilute NaOH solution in a test tube and add two drops of phenolphthalein solution.
- What is the colour of the solution?
- Add dilute sulphuric acid solution to the above solution drop by drop until the colour just changes.
- Is there any colour change for the reaction mixture?
- Why did the colour of phenolphthalein change after the addition of an acid?
- Now add a few drops of NaOH to the above mixture.
- Does the pink colour of phenolphthalein reappear?
- Why do you think this has happened?

In Activity 6.8, you will observe that the effect of a base is nullified by an acid and vice versa. The reaction that takes place is:

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

The reaction between an acid and a base to give a salt and water is known as a neutralisation reaction. In general, a *neutralisation reaction** can be written as:

$$\text{Base} + \text{Acid} \longrightarrow \text{Salt} + \text{Water}$$

ACTIVITY 6.9: Preparation of Lead Carbonate

Materials Required

Sodium carbonate, lead nitrate, test tubes, glass rod

Procedure

- Make a solution of sodium carbonate with water.
- Make a solution of lead nitrate with water.
- Take 20 ml of sodium carbonate solution in a test tube.
- Add 20 ml lead nitrate solution to the test tube.
- Observe the mixture (reaction).

Reaction

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

Observation

In this activity, you will observe that a solid white precipitate is produced. This is lead carbonate. The lead nitrate and sodium carbonate have exchanged their ions. This is an example of double replacement reaction.

* We have already studied the neutralisation reaction in Unit 5.

6.6.5 Preparation of Salt from Metal Oxides and Non-metal Oxide

Salts are also prepared from the reaction of metal oxides with non-metal oxides These oxides can combine because one is acidic (non-metal oxide) while the other is basic (metal oxide). Some examples are:

(*i*) Sodium oxide reacts with carbon dioxide to give sodium carbonate. The reaction takes place at a temperature of 500–550°C.

$$Na_2O(s) + CO_2(g) \longrightarrow Na_2CO_3(s)$$

(*ii*) Potassium oxide reacts with carbon dioxide to form potassium carbonate. The reaction takes place at a temperature of 400–450°C.

$$K_2O(s) + CO_2(g) \longrightarrow K_2CO_3(s)$$

(*iii*) Calcium Oxide reacts with sulphur trioxide to form calcium sulphate.

$$CaO(s) + SO_3(l) \longrightarrow CaSO_4(s)$$

(*iv*) Magnesium oxide reacts with carbon dioxide to form magnesium carbonate.

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

Note: Metal oxide reacts with carbon dioxide at very high temperature.

6.6.6 Preparing Salt with a Base and Non-metal Oxide

ACTIVITY 6.10: Preparation of Calcium Carbonate

- Take calcium hydroxide solution in a test tube.
- Pass carbon dioxide gas through lime water (calcium hydroxide solution) and record your observations.

Reaction

$$\underset{\text{(Lime water)}}{Ca(OH)_2(aq)} + CO_2(g) \longrightarrow \underset{\text{(White precipitate)}}{CaCO_3(s) + H_2O(l)}$$

Note: The salts obtained by reaction of *hydrochloric acid* are called *chlorides*.

The salts obtained by reaction of *sulphuric acid* are called *sulphates*.

The salts obtained by reaction of *nitric acid* are called *nitrates*.

The salts obtained by reaction of *carbonic acid* are called *carbonates*.

The salts obtained by reaction of acetic acid (CH_3COOH) are called acetates

EXERCISE 6.4

1. Copper oxide dissolves in dilute sulphuric acid to form __________.
2. Write the molecular formula of
   (*a*) White vitriol
   (*b*) Blue vitriol
3. Metal oxide reacts with carbon dioxide at very high temperature. (True or False)
4. Complete the following reactions:
   (*a*) $CaO(s) + SO_3(g) \longrightarrow ?$
   (*b*) $MgO(s) + CO_2(g) \longrightarrow ?$
5. The salts obtained by reaction of sulphuric acid are
   (*a*) chlorides
   (*b*) carbonates
   (*c*) acetates
   (*d*) none of these.

6.7 USES AND SOURCES OF SALT

Salt is all around us—underground and on the earth's surface in the dried up residues of ancient seas. Some salts have even arrived from outer space in meteors. But our biggest sources of salts are seas and oceans.

There are many different types and grades of salt and a number of different methods of production. White salt is produced by evaporating 'solution-mined' brine in pressure vessels. The rock salt we use for constructing roads comes from mining ancient deposits. In some countries the natural energy of the sun is used to evaporate brine produced from sea water.

6.7.1 Sodium Chloride (NaCl)

Sodium chloride (NaCl) is known as common salt. It is present in sea water along with many other salts. Sodium chloride is separated from these salts. It also exists as solid salt, in the form of rocks, in several parts of the world. The large crystals of sodium chloride obtained from rocks are generally brown in colour due to impurities present in them. This is known as Rock salt. Rock salt is mined just like any other mineral. Common salt is an important component of our food. It is also used for the extraction of sodium metal.

6.7.2 Washing Soda ($Na_2CO_3.10H_2O$)

Washing soda is prepared from sodium chloride by the **Ammonia-soda process** or **Solvay process.** In this process, CO_2 gas is bubbled through a brine solution saturated with ammonia. It results in the formation of sodium hydrogencarbonate.

$$NaCl + H_2O + CO_2 + NH_3 \longrightarrow \underset{\text{Ammonium chloride}}{NH_4Cl} + \underset{\text{Sodium hydrogen-carbonate}}{NaHCO_3}$$

$NaHCO_3$ so formed precipitates out in the presence of excess of sodium chloride. It is filtered off and then ignited to get sodium carbonate (Na_2CO_3).

$$\underset{\text{Sodium hydrogen carbonate}}{2NaHCO_3} \longrightarrow \underset{\text{Sodium carbonate}}{Na_2CO_3} + CO_2 + H_2O$$

Anhydrous sodium carbonate thus formed is called **soda ash.** When soda ash is dissolved in water and subjected to crystallisation, the crystals that separate out are of sodium carbonate decahydrate ($Na_2CO_3.10H_2O$) which is also known as **washing soda.**

$$\underset{\text{Soda ash}}{Na_2CO_3} + 10H_2O \longrightarrow \underset{\text{Washing soda}}{Na_2CO_3 \,.\, 10_2O}$$

Uses of Sodium Carbonate

- Large quantities of sodium carbonate are used in the manufacture of glass, borax, soap and caustic soda.
- It is used in the paper, paints and textile industries.
- It is used for softening hard water. It removes temporary as well as permanent hardness.
- It is used for washing purposes in the laundry.
- It is used as an important laboratory reagent both in qualitative and quantitative analysis.

6.7.3 Baking Soda, Sodium Hydrogen-carbonate ($NaHCO_3$)

Sodium hydrogencarbonate is obtained as a primary product in the Solvay's process for the manufacture of Na_2CO_3.

$$NaCl + NH_3 + CO_2 + H_2O \longrightarrow NH_4Cl + NaHCO_3$$

It can also be prepared by passing CO_2 through a solution of Na_2CO_3 whereby $NaHCO_3$ being less soluble crystallises out.

$$Na_2CO_3 + CO_2 + H_2O \longrightarrow 2NaHCO_3$$

Uses of Sodium Hydrogencarbonate

- It is used as a component of baking powder. In addition to sodium hydrogencarbonate, baking powder contains mild edible acid such as tartaric acid or some other similar acidic compounds. When mixed with water and heated, sodium hydrogencarbonate reacts with the acidic component of the baking powder producing CO_2 which causes bread or cake to swell and become light. Tartaric acid present in baking powder neutralises sodium carbonate which would otherwise impart bitter taste to the food item.

- It is used in soda-acid fire extinguishers. Soda-acid fire extinguishers contain a solution of $NaHCO_3$ and H_2SO_4. These two chemicals are brought into contact by pressing a knob or by inverting the extinguisher. CO_2 is produced which forces a stream of effervescing gas on the fire. CO_2 surrounds the combustible substance and cuts off the supply of oxygen in air. Thus, the fire gets extinguished.
- It is used in medicines. It acts as mild antiseptic for skin infections. It is also present as an ingredient in antacids. Being alkaline it neutralises excess acid in the stomach.

EXERCISE 6.5

1. Write two uses of sodium carbonate.
2. What is the molecular formula of washing soda?
3. The chemical name of baking soda is __________.
4. Sodium chloride is present as an ingredient in antacids. . (True or false)
5. The crystals of sodium chloride obtained from rocks are generally white in colour. (True or False)

6.8 IDENTIFICATION OF IONS

6.8.1 Identification of Cations

For identification of cations (positive ions), dilute sodium hydroxide (NaOH) solution and dilute ammonia solution are used.

> Most cations combine with hydroxide ions to form insoluble metal hydroxide.
>
> $$M^{n+}(aq) + nOH^-(aq) \longrightarrow M(OH)_n(s)$$

Note: For the identification of both cations and anions, precipitation reactions are used. The chemical reagent used for the precipitation of insoluble salt of the concerned ion is called precipitating reagent.

NH_4^+ ions do not form a precipitate with both aqueous sodium hydroxide and aqueous ammonia. When a mixture of a strong alkali (NaOH) and ammonium ions are heated, ammonia gas is given out.

$$NH_4^+ + NaOH \xrightarrow{\text{Heat}} NH_3$$

> Ammonia gas can be detected by its characteristic smell. It changes moist red litmus paper into blue.

Ca^{2+} ions form a *white* precipitate with aqueous sodium hydroxide but not with aqueous ammonia. The precipitate does not dissolve in excess sodium hydroxide.

Zn^{2+} ions form *white* precipitate with both aqueous sodium hydroxide and aqueous ammonia. The precipitate dissolves in excess of both alkalis.

Mg^{2+} ions form a *white* precipitate with both aqueous sodium hydroxide and aqueous ammonia but the precipitate does not dissolve in excess of both alkalis.

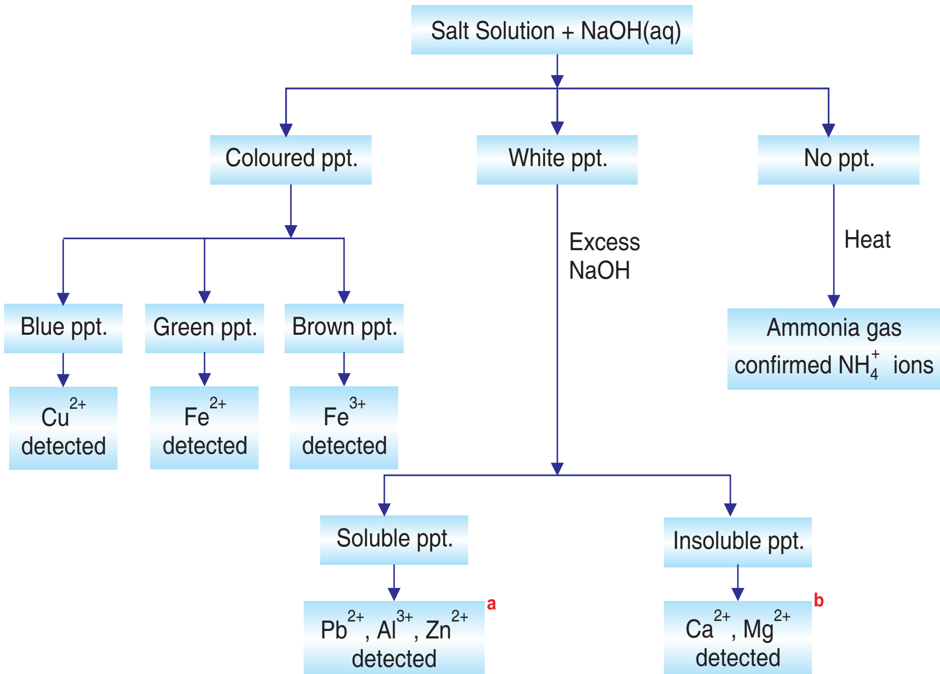
Cu^{2+} ions form a *blue* precipitate with both aqueous sodium hydroxide and aqueous ammonia. The precipitate dissolves in excess ammonia to form a dark blue solution.

Fe^{2+} ions form a *green* precipitate with both aqueous sodium hydroxide and aqueous ammonia. The precipitate does not dissolve in excess of both alkalis.

Fe^{3+} ions form a *brown* precipitate with both aqueous sodium hydroxide and aqueous ammonia. The precipitate does not dissolve in excess of both alkalis.

Al^{3+} ions form a *white* precipitate with both aqueous sodium hydroxide and aqueous ammonia. The precipitate dissolves in excess aqueous sodium hydroxide but not in excess aqueous ammonia.

Figures 6.4 and 6.5 summarise the reaction of cations with aqueous sodium hydroxide and aqueous ammonia respectively.



[a]To distinguish between Al^{3+}, Zn^{2+} and Pb^{2+}

If salt solution + KI(*aq*) $\longrightarrow$ Yellow ppt; the cation present in solution is Pb^{2+}

If salt solution + H_2S(*aq*) $\longrightarrow$ Black ppt, the cation present in solution is Zn^{2+}

If salt solution + KI(*aq*) $\longrightarrow$ No ppt, the cation present in solution may be Al^{3+}

or

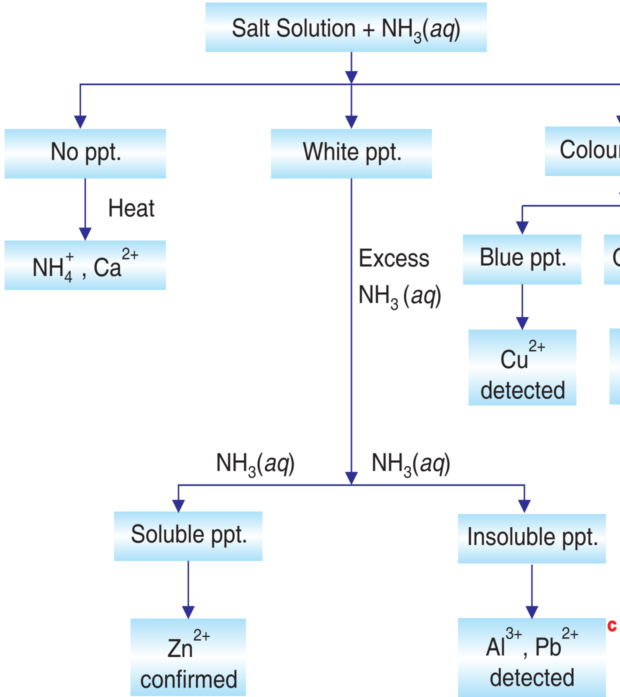
Salt solution + NH_4OH(*aq*) $\longrightarrow$ White ppt $\xrightarrow[NH_4OH]{\text{Excess}}$ Insoluble white ppt. confirms Al^{3+}

[b]To distinguish between Ca^{2+} and Mg^{2+}

If salt solution + NH_3(*aq*) $\longrightarrow$ White ppt $\xrightarrow[NH_3]{\text{Excess}}$ Insoluble white ppt. confirms Mg^{2+}

If salt solution + NH_3(*aq*) $\longrightarrow$ No ppt. confirms Ca^{2+}

Figure 6.4: Identification of cations using aqueous sodium hydroxide.



[c]To distinguish between Al^{3+} and Pb^{2+}

Distinction between Al^{3+} ion and Pb^{2+} ion can be done by adding aqueous potassium iodide to the solution. If yellow ppt. is obtained, it confirms the presence of Pb^{2+} ion; otherwise Al^{3+} ion detected.

Figure 6.5: Identification of cations using aqueous ammonia.

EXERCISE 6.6

1. NH_4^+ ions form precipitate with aqueous sodium hydroxide. (True or False)
2. Fe^{2+} ions form __________ precipitate with aqueous sodium hydroxide.
3. How will you identify Cu^{2+} ions in a sample of solution?
4. The precipitate formed by Al^{3+} ion does not dissolve in

(*a*) water (*b*) excess aqueous sodium hydroxide

(*c*) excess aqueous ammonia (*d*) All of these.

5. Fe^{3+} ions form a brown precipitate with __________ and __________.

6.8.2 Identification of Anions

The common chemical reagents used in the identification of anions are dilute hydrochloric acid (HCl), aqueous barium nitrate ($Ba(NO_3)_2$), aqueous silver nitrate ($AgNO_3$), dilute nitric acid (HNO_3) and dilute ammonia (NH_3) solution. Besides, red litmus is also used.

Cl^- ions form *white* precipitate when aqueous silver nitrate ($AgNO_3$) followed by excess dilute nitric acid (HNO_3) is added to salt solution. The precipitate does not dissolve in excess dilute nitric acid, but dissolves in excess aqueous ammonia. The white precipitate darkens if kept in sunlight for 10–15 minutes.

I^- ions form *yellow* precipitate when acidified with dilute nitric acid followed by aqueous lead nitrate [$Pb(NO_3)_2$] or silver nitrate($AgNo_3$) insoluble in access ammonia. The precipitate does not dissolve in excess ammonia, but dissolves in hot water forming a colourless solution.

NO_3^- ion does not form any precipitate with chemical reagent when a mixture of aqueous sodium hydroxide and nitrate ion along with aluminium powder is heated, ammonia gas is given out.

or

> Hydrogen gas is generated from the reaction between aluminium and sodium hydroxide.

Nitrate ion decomposes into nitrogen dioxide and oxygen when heated.

Nitrogen dioxide gas is brown in colour whereas oxygen is a colourless gas. Oxygen gas is tested by a glowing splint.

> Nitrates can be reduced by hydrogen to ammonia.

CO_3^{2-} ion turns moist red litmus into blue. Effervescence (bubbles) of colourless and odourless carbon dioxide gas forms when carbonates react with dilute hydrochloric acid.

> Carbon dioxide gas can be tested with lime water. CO_2 gas turns lime water milky. Excess carbon dioxide clears the milkiness.

SO_3^{2-} ion forms a *white* precipitate with both aqueous barium nitrate [$Ba(NO_3)_2$] and aqueous barium chloride ($BaCl_2$). The precipitate dissolves in both dilute hydrochloric acid and dilute nitric acid.

SO_4^{2-} ions form white precipitate with both aqueous barium chloride and aqueous barium nitrate. The precipitate neither dissolves in dilute nitric acid nor in hydrochloric acid. Figure 6.6 summarises the reaction of anions with chemical reagents in order to identify them.

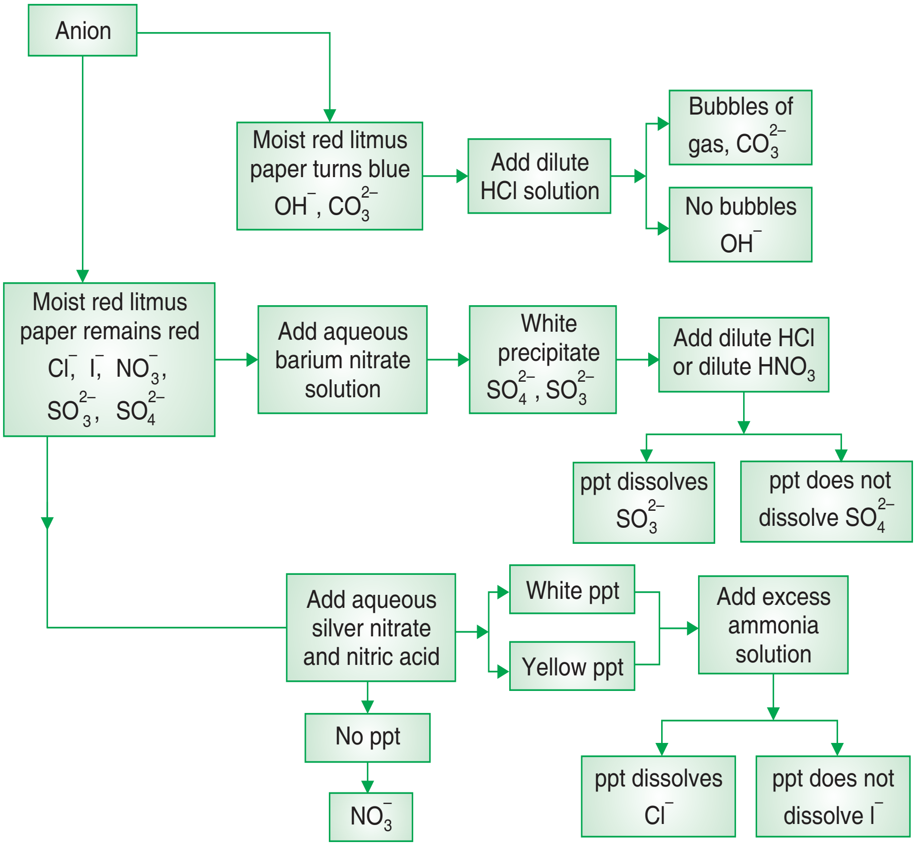


Figure 6.6: Identification of anions.

EXERCISE 6.7

1. How will you identify SO_3^{2-} ions in a solution?
2. I^- ions form __________ precipitate when acidified with dilute nitric acid followed by aqueous lead nitrate.
3. Name three reagents used in the identification of anions.
4. Nitrate ion decomposes into nitrogen dioxide and oxygen when heated. (True or False)
5. CO_3^{2-} ion turns moist __________ litmus into __________.

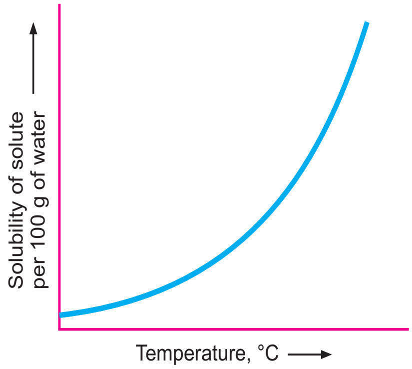
6.8.3 Identification of Gases

The following table gives the tests for identifying gases.

Gases	Colour and smell	Tests
Oxygen (O_2)	Colourless Odourless	Relights a glowing splint
Hydrogen (H_2)	Colourless Odourless	'Pops' with a lighted splint
Ammonia (NH_3)	Colourless Pungent smell	Turns moist red litmus paper blue. Produces dense white fumes of NH_4Cl with hydrogen chloride gas.
Carbon dioxide (CO_2)	Colourless Odourless	Turns limewater milky (excess CO_2 discharges the white ppt.).
Chlorine (Cl_2)	Greenish-Yellow Pungent smell (Poisonous)	Turns moist blue litmus paper red and then, bleaches it
Sulphur dioxide (SO_2)	Colourless Choking smell	Turns acidified $K_2Cr_2O_7(aq)$ from orange to green. (Using filter paper soaked in the solution.)

6.9 SUMMARY

- A solution which is unable to dissolve any more of the solute at a particular temperature is called a saturated solution at that temperature.
- A solution which can dissolve more solute at a given temperature is called an unsaturated.
- A solution containing relatively small amount of solute in a fixed amount of solvent or compared to that of the solvent is a dilute solution.
- Solution containing relatively more quantity or large amount of solute in the fixed amount of solvent is a concentrated solution.
- The solubility of any solute in a solvent depends on temperature.
- The graph showing relationship between temperature and solubility of the substance at different temperatures is called a solubility curve.



- The general shape of the curve indicates the rates of change in the solubility with a rise in temperature.
- Solubility = $\frac{\text{Wt. of solute (in gram)}}{\text{Wt. of solvent (in gram)}} \times 100$ (at a particular temperature)
- For the identification of cations and anions, precipitation reactions are used.
- For the identification of cations (positive ions), dilute sodium hydroxide (NaOH) solution and dilute ammonia solution are used.
- The common chemical reagents used in the identification of anions are dilute hydrochloric acid (HCl), aqueous barium nitrate ($Ba(NO_3)_2$), aqueous silver nitrate ($AgNO_3$), dilute nitric acid (HNO_3) and dilute ammonia (NH_3) solution. Besides, red litmus is also used.

6.10 GLOSSARY

- **Anhydrous:** substance containing no water.
- **Borax:** a white compound which occurs as a mineral in some alkaline salt deposits and is used in making glass and ceramics.
- **Crystallisation:** a separation technique that is used to separate a solid that has dissolved in a liquid.
- **Crystallise:** form or cause to form crystals.
- **Crystals:** a piece of a homogeneous solid substance having a natural geometrically regular form.
- **Excess:** an amount of something that is more than necessary.
- **Granules:** a small compact particle of a substance.
- **Phenolphthalein:** a colourless crystalline solid (pink in alkaline solution) used as an acid-base indicator.
- **Stirring:** mixing thoroughly.

6.11 UNIT ASSESSMENT

I. Multiple Choice Questions

1. A solution which can dissolve more solute at a given temperature is called ______ .

(*a*)Saturated (*b*) Unsaturated (*c*) Supersaturated (*d*) none of these

2. Choose the correct statement

(*a*)With increase in temperature, the saturated solution becomes unsaturated.

(*b*)The quantity of solute dissolved in a certain quantity of solvent denotes the concentration of solution.

(*c*)Both (*a*) and (*b*)

(*d*)None of these

3. Which of these is a dilute solution?

(*a*)A teaspoon of sugar in 200 ml water

(*b*)Three teaspoons of sugar in 200 ml water

(*c*)Both (*a*) and (*b*)

(*d*)neither (*a*) nor (*b*)

4. The temperature at which reversal of solubility occurs is called

(*a*)critical temperature (*b*) transition temperature

(*c*)both (*a*) and (*b*) (*d*) None of these

5. Which of the following is correct?

(*a*)Solubility $= \dfrac{\text{Weight of Solvent}}{\text{Weight of Solute}} \times 100$

(*b*)Solubility $= \dfrac{\text{Weight of Solute}}{\text{Weight of Solvent}} \times 100$

(*c*) $\dfrac{\text{Weight of Solvent}}{100} = \dfrac{\text{Weight of Solute}}{\text{Solubility}}$

(*d*)Both (*b*) and (*c*)

6. Sodium oxide reacts with carbon dioxide to give

(*a*)Carbonic acid (*b*) Sodium carbonate

(*c*)Sodium chloride (*d*) Sulphuric acid

7. Complete the reaction:

$$CaO\ (s) + SO_3\ (l) \longrightarrow ?$$

(*a*)$CaSO_4$ (*b*) CaS (*c*) SO_2 (*d*) $CaCO_3$

8. Metal oxide reacts with carbon dioxide at

(*a*) room temperature (*b*) low temperature

(*c*) very high temperature (*d*) does not react

9. Fe^{3+} ions form a ______ precipitate with both Na(OH) (*aq*) and NH_3 (*aq*)

(*a*) white (*b*) blue (*c*) green (*d*) brown

10. CO_2 gas turns lime water to

(*a*) milky white (*b*) red (*c*) blue (*d*) yellow

II. Open Ended Questions

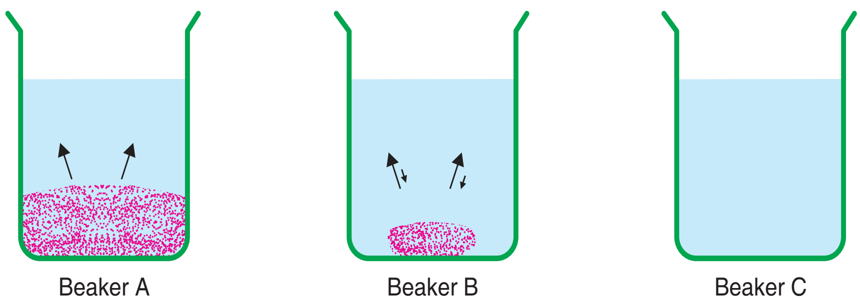
1. Distinguish between saturated and unsaturated solutions.
2. Explain supersaturated solution.
3. How can you prepare a supersaturated solution?
4. How can a saturated solution be changed to unsaturated?
5. Prepare zinc sulphate crystals from zinc metal and dilute sulphuric acid.
6. What are the factors that influence the solubility of salts?
7. Give two uses of solubility curves.
8. Write two uses of
   (*a*) Washing soda
   (*b*) Baking soda
9. How will you identify Ca^{2+} ions and Fe^{2+} ions in a sample of solution?
10. How will you identify CO_3^{2-} ions in a solution?

III. Numericals

1. The solubility of a solute at 30°C is 20. What amount of water is required to make saturated solution of 80 grams of a solute?
2. At 30°C, 14 grams of sugar dissolves in 10 grams of water to form a saturated solution. Find the solubility of sugar?

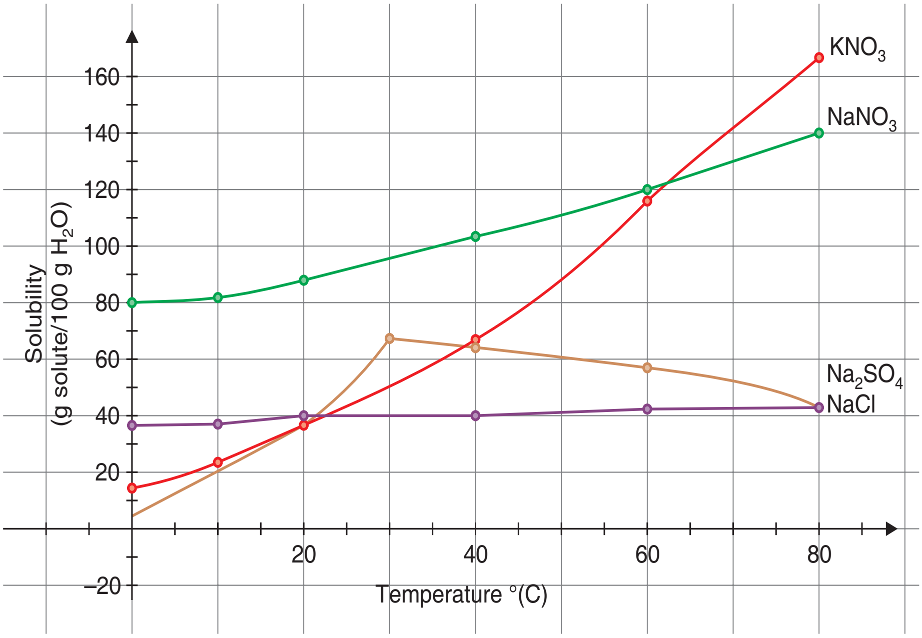
IV. Practical-based Questions

1. Which beaker contains unsaturated solution?



(*a*) Beaker A (*b*) Beaker B (*c*) Beaker C (*d*) None of these

2. Look at the solubility graph below and answer the questions that follow:



(i) How much sodium nitrate will dissolve at 60°C?

(*a*) 100 grams (*b*) 110 grams (*c*) 120 grams (*d*) 130 grams

(ii) Which solute is the most soluble at 80°C?

(*a*) Sodium Chloride (*b*) Sodium Sulphate

(*c*) Sodium Nitrate (*d*) Potassium Nitrate

(iii) Which solute is the least soluble at 10°C?

(*a*) Potassium Nitrate (*b*) Sodium Chloride

(*c*) Sodium Sulphate (*d*) Sodium Nitrate

(iv) 140 grams of solute is dissolved at 80°C is true for

(*a*) Sodium Nitrate (*b*) Potassium Nitrate

(*c*) Sodium Sulphate (*d*) Sodium Chloride

PROJECT

Carry out an experiment to show the method/process of crystallisation.

Unit 7

The Mole Concept and Gas Laws

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- explain the mole concept.
- explain the concepts of: relative atomic mass, relative formula mass, relative molecular mass, molar mass, limiting reactant, empirical and molecular formulae.
- state the gas laws: Gay-Lussac, Charles' law, Boyle's law and the ideal gas law, Grahams' law of diffusion.

KNOWLEDGE GAIN

Amedeo Avogardo Cerreto was an Italian scientist. He is most noted for his contributions to molecular theory, including what is known as Avogadro's law.

In tribute to him, the number of elementary entities (atoms, molecules, ions or other particles) in 1 mole of a substance, 6.022×10^{23}, is known as the Avogadro constant.

7.1 AVOGADRO NUMBER AND THE MOLE CONCEPT

ACTIVITY 7.1: Illustrating Link between Number of Particles and Mass of Particles

Materials Required

A paper bag, 5 handfuls peanuts, balance

Can you tell how many peanuts are there? (without counting)

- Using balance, measure the mass of the paper bag (say 50 g)
- Now measure the mass of a single peanut (say 0.5 g)
- Put 5 handfuls peanuts in a paper bag

  Note: Size of each peanut should be same.
- Now, weigh all the peanuts. Let it be 250 g.

Calculation

- Paper bag + 5 handfuls peanuts = 250 g

  5 handfuls peanuts = 250 g – 50 g (weight of paper bag)

  = 200 g.

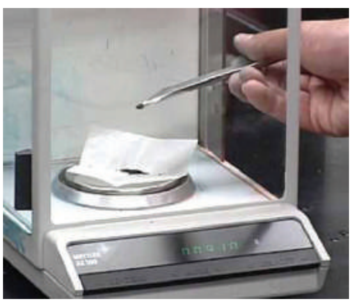
- Number of peanuts $= \dfrac{\text{Total weight of peanuts}}{\text{Weight of single peanuts}} = \dfrac{200\text{ g}}{0.5\text{ g}} = \dfrac{200 \times \cancel{10}^{2}}{\cancel{5}} = 400$

  Total number of peanuts = 400

In this way, you can count the number of objects by weighing the object.

In the same manner, bank people count a large number of coins. In chemistry also, scientists link the mass of an element or compound to the number of atoms or molecules present in them. This is done through the mole. Thus, mole is a link between the mass of atoms (or molecules) and the number of atoms (or molecules).

(a) Counting coins by weighing.



(b) Counting atoms by weighing chemicals.

Figure 7.1

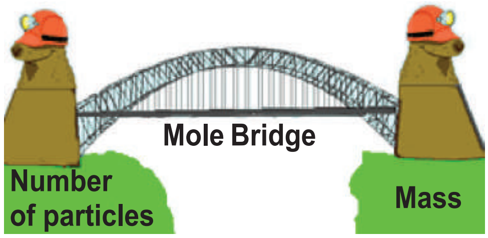


Figure 7.2: Mole links between number of particles and mass of atoms.

The word *mole* was introduced around 1896 by a German Physical Chemist Wilhem Ostwald who derived the term from the Latin word, '*mole*' which means a pile or a heap. A substance may be considered as a heap of atoms or molecules.

In everyday life, we use counting units like a dozen (12 objects) and a gross (144 objects) to deal with a large quantities. In Chemistry, the unit for dealing with the number of atoms, ions or molecules is mole. A mole is a group of 6.022×10^{23} particles (atoms, molecules or ions) of a substance. The SI symbol of mole is mol.

A mole is defined as the amount of matter that contains as many objects (atoms, molecules, ions or whatever objects we are considering) as the number of atoms in exactly 12 g of pure ^{12}C (carbon-12 isotope).

From numerous experiments, scientists have determined the number of atoms in 12 g of ^{12}C (carbon-12) and found it to be 6.022×10^{23}. This number is called **Avogadro's number**, in honour of an Italian scientist Amedeo Avogadro. In this book, we will use 6.02×10^{23} or 6.022×10^{23} for Avogadro's number. Avogadro's number is represented by N_A.

Take an example of the reaction of hydrogen and oxygen to form water:

$$2H_2 + O_2 \rightarrow 2H_2O.$$

The above reaction indicates that two molecules of hydrogen combine with one molecule of oxygen to form two molecules of water.

We can infer from the above equation that the quantity of a substance can be characterised by its mass or the number of molecules. But, a chemical reaction equation indicates directly the number of atoms or molecules taking part in the reaction. Therefore, it is more convenient to refer to the quantity of a substance in terms of moles (number of its molecules or atoms) rather than their masses. One mole of any particles (atoms, molecules, ions or particles) is that quantity in number having a mass equal to its atomic or molecular mass in grams.

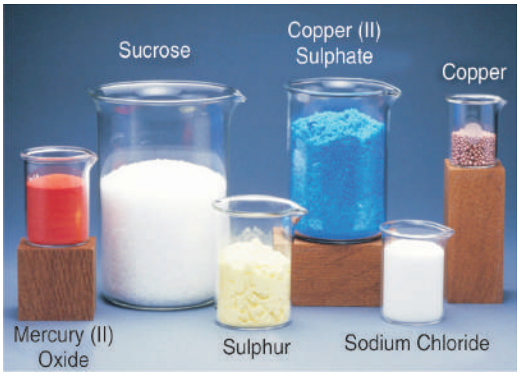


Figure 7.3: One mole of six chemicals.

The number of particles present in 1 mole of any substance is 6.022×10^{23}.

The mass of 1 mole of a substance is equal to its relative atomic or molecular mass in grams. The atomic mass of an element gives us the mass of one atom of that element in atomic mass units (u). To get the mass of 1 mole of atom of that element, we have to take the same numerical value but change the units from 'u' to 'g'. The mass of atoms in grams is also known as gram atomic mass. For example, atomic mass of hydrogen = 1 u. So, gram atomic mass of hydrogen = 1 g.

1 u hydrogen has only 1 atom of hydrogen, 1 g hydrogen has 1 mole atoms, that is, 6.022×10^{23} atoms of hydrogen.

Similarly,

16 u oxygen has only 1 atom of oxygen,

16 g oxygen has 1 mole atoms, that is,

6.022×10^{23} atoms of oxygen.

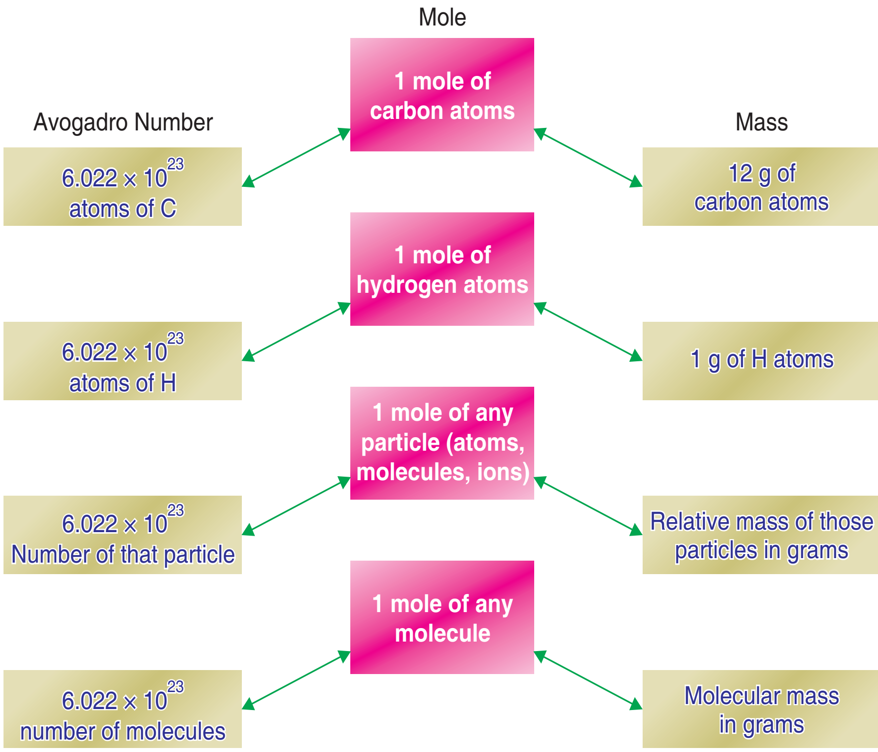


Figure 7.4: Relationship between Avogadro number, mole and mass.

EXERCISE 7.1

1. Avogadro number is the number of atoms in 1 gram-atom of an element. (True or False)
2. A group of 6.022×10^{23} particles (atoms, ions, or molecules) of a substance is called a ______ of that substance.
3. One mole of CO_2 contains _______ atoms of carbon.
4. The mass of one Avogadro number of nitrogen atoms is equal to:
   (*a*) 14 amu (*b*) 14 g (*c*) 28 g (*d*) 14 kg
5. Mole is a link between the _________ of atoms (or molecules) and the _____ of atoms (or molecules).

ACTIVITY 7.2: Illustrating Concept of Moles

In groups, discuss the concept of moles as a way of expressing the amount of substance and make a presentation.

7.2 CALCULATION OF THE NUMBER OF MOLES

(*i*) The number of moles

$$= \frac{\text{Given no. of particles}}{\text{Avogadro number}}$$

(*ii*) The number of moles

$$= \frac{\text{mass of element in grams}}{\text{molar mass of element}}$$

(*iii*) The number of moles

$$= \frac{\text{mass of molecule in grams}}{\text{molar mass of the molecule}}$$

Example 1

Calculate the number of moles of 12.044 × 10^{23} helium atoms.

Solution

$$\text{No. of moles} = \frac{\text{No. of particles}}{\text{Avogadro number}}$$

$$= \frac{12.044 \times 10^{23}}{6.022 \times 10^{23}} = \mathbf{2\ mol}$$

Example 2

Calculate the number of particles in 0.1 mole of carbon atoms.

Solution

No. of particles = No. of moles × Avogadro number

$= 0.1 \times 6.022 \times 10^{23}$

$= \mathbf{6.022 \times 10^{22}}$ **moles**

Example 3

Calculate the number of moles of 3.011 × 10^{23} hydrogen atoms.

Solution

Number of moles

$$= \frac{\text{No. of particles}}{\text{Avogadro number}}$$

$$= \frac{3.011 \times 100^{23}}{6.022 \times 10^{23}} = \mathbf{0.5\ mole}$$

Example 4

(*a*) *Distinguish between N_2 and 2N.*

Calculate the mass of particles in the following:

(*b*) *0.5 mole of nitrogen atom*

(*c*) *0.5 mole of nitrogen molecule*

(*Given: Atomic mass of nitrogen atom = 14 g*

Molecular mass of nitrogen molecule = 28 g)

Solution

(*a*) N_2 stands for 1 mole of nitrogen molecules and 2N stands for 2 moles of nitrogen atoms.

(*b*) Mass of nitrogen atoms = Atomic mass of nitrogen atom × number of moles

Mass of nitrogen atoms = ?

Atomic mass of nitrogen = 14 g

Molecular mass of N_2 = 28 g

No. of mole = 0.5

= 14 × 0.5 = 7 grams

The mass of 0.5 mole of nitrogen atom is 7 grams.

(*c*) Mass of nitrogen molecules (N_2)

= Molecular mass of molecules × no. of moles

= 28 × 0.5 = **14 grams**

Example 5

Calculate the no. of moles in

(*i*) *0.478 g of magnesium*

(*ii*) *12 g oxygen gas*

Given: Atomic mass of Mg = 24

Atomic mass of O = 16.

Solution

(*i*) No. of moles = $\dfrac{\text{Mass of Mg in grams}}{\text{Atomic mass}}$

$$= \frac{0.478}{24} = \mathbf{0.019\ mole}$$

(*ii*) In Oxygen gas, there are two atoms of Oxygen (O_2)

Molecular mass of $O_2 = 16 \times 2 = 32$ g

No. of moles in 12 g $O_2 = \dfrac{12}{32} = \mathbf{0.375\ mole}$

Example 6

Calculate the mass of 6.022×10^{23} nitrogen molecules.

Solution

The mass of nitrogen molecules (N_2) = $14 \times 2 = 28$ g

1 mole of N_2 molecules = 6.022×10^{23} molecules of nitrogen.

1 mole of N_2 molecule weighs = 28 g

Therefore, 6.022×10^{23} molecules will weigh 28 grams.

EXERCISE 7.2

1. Convert 12 g of oxygen gas into moles.
2. What is the mass of

(*i*)0.5 mole of water molecules? (*ii*) 0.2 mole of oxygen atoms?

7.3 DEFINITION OF RELATIVE ATOMIC MASS

Just like mole, relative atomic mass is based on Carbon-12 isotope as a standard measure. *The relative atomic mass (Ar) of an element is the average relative mass of an atom of an element as compared with an atom of ^{12}C taken as 12 atomic mass unit.* Hence,

Relative atomic mass of an element (A_r)

$$= \frac{\text{Average mass of one atom of an element}}{\frac{1}{12} \times (\text{Mass of one } ^{12}\text{C atom})}$$

The relative atomic mass (A_r) of an element is a pure number and it does not have a unit. The relative atomic masses of some elements are given in table 7.1.

Table 7.1: Relative Atomic Masses (A_r) of Some Elements

Element Name	Symbol	Relative Atomic Mass (A_r)	
		Exact Value	Common Value*
Aluminium	Al	26.982	27
Beryllium	Be	9.012	9
Bismuth	Bi	208.980	209
Bromine	Br	79.904	80
Calcium	Ca	40.080	40
Carbon	C	12.011	12
Chlorine	Cl	35.453	35.5
Chromium	Cr	51.996	52
Cobalt	Co	58.933	59
Copper	Cu	63.546	63.5
Fluorine	F	18.918	19
Gold	Au	196.966	197
Helium	He	4.003	4
Hydrogen	H	1.008	1
Indium	In	114.820	115
Iodine	I	126.904	127
Iron	Fe	55.847	56
Lead	Pb	207.200	207
Magnesium	Mg	24.305	24
Manganese	Mn	54.938	55
Mercury	Hg	200.59	200
Nitrogen	N	14.007	14
Oxygen	O	15.999	16
Phosphorus	P	30.974	31
Potassium	K	39.098	39
Silicon	Si	28.086	28
Silver	Ag	107.868	108
Sodium	Na	22.980	23
Sulphur	S	32.060	32
Thorium	Th	232.038	232
Tin	Sn	118.690	118.5
Titanium	Ti	47.900	48
Tungsten	W	183.85	184
Uranium	U	238.029	238
Zinc	Zn	65.38	65.5

*The common values are used in numerical problems.

7.4 DEFINITION AND CALCULATION OF RELATIVE MOLECULAR MASS

Relative Molecular Mass (M_r)

Like the atomic masses, the molecular masses of compounds are also very small and these cannot be measured directly. The molecular masses of compounds are also expressed as relative molecular masses (M_r). The relative molecular mass of a compound is the average relative mass of its molecule as compared with the mass of one ^{12}C atom taken as 12 u.

Relative molecular mass (M_r) =

$$\frac{\text{Average mass of one molecule of the compound}}{\frac{1}{12} \times (\text{Mass of an atom of } ^{12}C)}$$

The relative molecular mass (M_r) of a compound is a pure number and it does not have a unit.

Molecular Mass (M)

The average mass of one molecule of a compound in atomic mass unit is called **molecular mass** (M). Hence,

Molecular mass (M) = $M_r \times 1\ u = M_r\ u$

The molecular mass (M) has the unit of mass, i.e., g, kg or u. Note that the magnitudes of molecular mass (M) and relative molecular mass (M_r) are equal. They differ only in their units.

Calculation of molecular mass from atomic masses: The molecular mass of a compound is calculated by adding the atomic masses of all the atoms present in one molecule of the compound. This is illustrated on next page.

(a) Molecular mass of water: The molecular formula of water is H_2O. Hence,

Molecular mass of water = (2 × Atomic mass of hydrogen) + (1 × Atomic mass of oxygen) = (2 × 1) + (1 × 16) = 18 u.

(b) Molecular mass of nitric acid (HNO_3): The molecular formula of nitric acid is HNO_3. Hence, Molecular mass of nitric acid = (1 × Atomic mass of hydrogen) + (1 × Atomic mass of nitrogen) + (3 × Atomic mass of oxygen)

= (1 × 1) + (1 × 14) + (3 × 16)

= 1 + 14 + 48 = 63 u.

Example 7

Calculate the molecular mass of glucose ($C_6H_{12}O_6$).

Solution

Molecular mass of glucose

= (6 × atomic mass of C)

+ (12 × atomic mass of H)

+ (6 × atomic mass of O)

= 6 × 12 u + 12 × 1 u + 6 × 16 u

= 72 u + 12 u + 96 u = **180 u**

EXERCISE 7.3

1. Calculate the molecular mass of

(*a*)Phosphorus molecule, P_4

(*b*)Sulphur molecule, S_8

Given: Atomic masses S = 32 u and P = 31 u

7.5 DEFINITION AND CALCULATION OF RELATIVE FORMULA MASS

The sum of the atomic masses of all the atoms in a formula unit of a compound is called the **formula unit mass** of the compound. Always remember that the concept of the formula unit mass is used in case of ionic compounds. This is because in case of ionic compounds there are no separate individual molecules but they exist as aggregates, i.e., as cluster of ions. For example, solid sodium chloride is represented as $(Na^+Cl^-)_n$ and in the simplified form it is written as NaCl.

The formula unit mass of ionic compounds is calculated in the same manner as in the case of the molecular mass described earlier. For example, the formula unit mass of NaCl

= Atomic mass of Na + Atomic mass of Cl

= 23 u + 35.5 u = 58.5 u.

Example 8

Calculate the formula unit masses of ZnO, Na_2O and K_2CO_3. (Given atomic mass: Zn = 65 u, Na = 23 u, K = 39 u, C = 12 u and O = 16 u)

Solution

Formula unit mass of ZnO

= atomic mass of Zn + atomic mass of O

= 65 u + 16 u = 81 u.

Formula unit mass of Na_2O

= (2 × atomic mass of Na)

+ atomic mass of O

= 2 × 23 u + 16 u

= 46 u + 16 u = 62 u

Formula unit mass of K_2CO_3

= (2 × atomic mass of K)

+ atomic mass of C

+ (3 × atomic mass of O)

= 2 × 39 u + 12 u + 3 × 16 u

= 78 u + 12 u + 48 u

= **138 u**

EXERCISE 7.4

1. Calculate the formula unit mass of

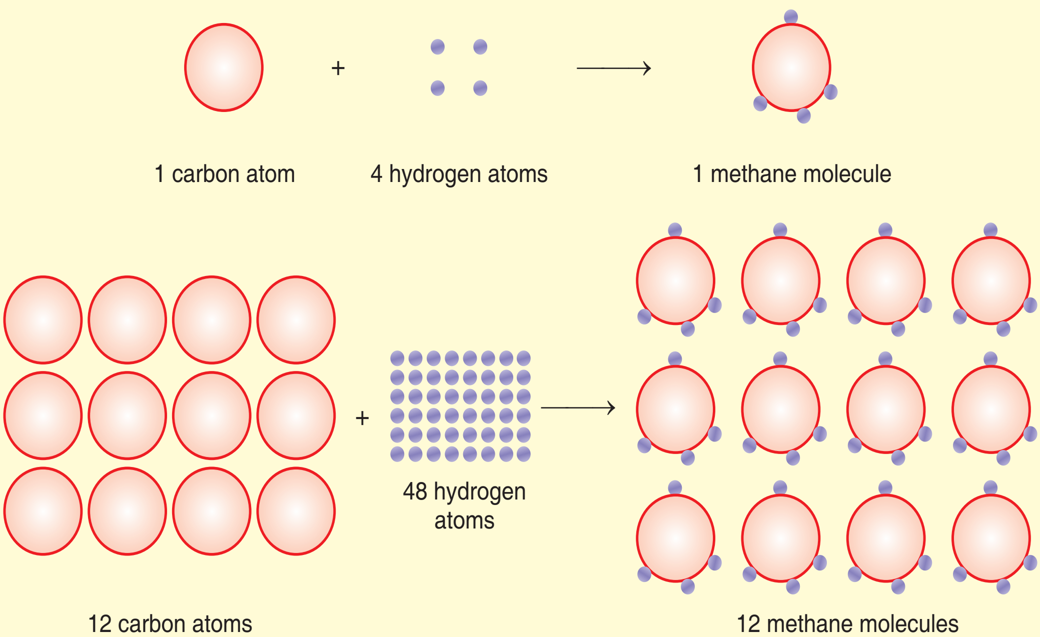
(*i*) ZnO (*ii*) Na_2CO_3 (*iii*) $C_6H_{12}O_6$

Given: Atomic mass : Zn = 65 u, Na = 23 u, C = 12 u, H = 1u, O = 16 u

7.6 CALCULATION OF MOLAR MASS

ACTIVITY 7.3: Understanding Molar Mass

Look at the following figure.



Can you calculate the mass of 1 mole of methane and 12 moles of methane?

What is the mass of 6.022×10^{23} molecules of methane?

We know that one methane molecule contains one carbon atom and four hydrogen atoms. 1 mole of methane molecules consists of 1 mole of carbon atoms and 4 moles of hydrogen atoms. The mass of 1 mole of methane can be found by adding the masses of carbon and hydrogen present.

Mass of 1 mole carbon

$$= 1 \times 12.01 \text{ g} = 12.01 \text{ g mol}^{-1}$$

Mass of 4 moles hydrogen

$= 4 \times 1.00 \text{ g} = 4.00 \text{ g mol}^{-1}$

Mass of 1 mole methane

$= 12.01 + 4.00 = 16.01 \text{ g mol}^{-1}$

The quantity 16.01 g mol^{-1} is called the molar mass of methane. 1 mole methane has 6.022×10^{23} molecules so the mass of 6.022×10^{23} molecules is also 16.01 g. Calculate the mass of 12 moles of methane yourself.

Mass of one mole of any substance is called its molar mass. Alternatively, the average mass of one mole of any substance is called its **molar mass**. Molar mass is represented by **M** and its unit is g mol^{-1}. It is given by,

Molar mass (M)

$$= \frac{\text{Mass of the substance}}{\text{Amount of the substance in moles}} = \frac{w}{n}$$

Molar mass of a substance

= Mass of 6.023×10^{23} chemical units of that substance

Molar mass of one hydrogen atom

= Mass of 6.023×10^{23} atoms of hydrogen

Molar mass of one hydrogen molecule

= Mass of 6.023×10^{23} molecules of hydrogen

Example 9

Calculate the molar mass of sulphur dioxide.

Solution

Mass of 1 mol S $= 1 \times 32.07 = 32.07 \text{ g mol}^{-1}$

Mass of 2 mol O $= 2 \times 16.00 = 32.00 \text{ g mol}^{-1}$

Mass of 1 mol SO_2 $= 64.07 \text{ g mol}^{-1}$

Molar mass = 64.07

The molar mass of SO_2 is 64.07 g mol^{-1}. It represents the mass of 1 mole of SO_2 molecules.

EXERCISE 7.5

1. Calculate the molar masses of the following substances :

(*a*)Ethyne, C_2H_2

(*b*)Hydrochloric acid, HCI

(*c*)Nitric acid, HNO_3

(Atomic masses : C = 12 u, H = 1 u, Cl = 35.5 u, N = 14 u, O = 16 u)

7.7 RELATIONSHIP BETWEEN NUMBER OF MOLES, MASS AND MOLAR MASS

How many moles are there in a certain mass of a substance?

The number of moles in a certain mass of a substance is calculated using the following formula:

No. of moles of a substance, X

$$n = \frac{\text{Mass of the substance in grams}}{\text{Molar mass of the substance in grams per mole}}$$

$$= \frac{w\text{g}}{\text{M g mol}^{-1}} = \frac{w}{\text{M}} \text{ mol}$$

where w is the mass of the substance and M is the molar mass of the substance.

How many molecules are there in a certain mass of a substance?

The number of molecules present in a certain mass of a substance is calculated using the following formula:

No. of molecules of a substance

$$= \frac{\text{Mass of the substance}}{\text{Molar mass of the substance}} \times \text{Avogadro number}$$

$$= \frac{m}{\text{M}} \times 6.023 \times 10^{23} \text{ molecules}$$

where m is the mass of the substance and M is the molar mass of the substance.

Example 10

How many atoms of S_8 molecules are present in 50 g of sulphur? The relative atomic mass of sulphur is 32.

Solution

32 g of sulphur contain 6.023×10^{23} atoms of sulphur.

50 g of sulphur contain $\frac{6.023 \times 10^{23} \times 50}{32}$ atoms

$$= 9.411 \times 10^{23} \text{ atoms.}$$

8 atoms of sulphur are present in 1 molecule of sulphur.

9.411×10^{23} atoms of sulphur are present in $\frac{9.411 \times 10^{23}}{8}$ molecules

$$= 1.176 \times 10^{23} \text{ molecules of sulphur.}$$

Example 11

The antibiotic penicillin has the molecular formula $C_{16}H_{18}N_2SO_4$. One injection of penicillin contains 500 mg of penicillin. When one intra-muscular injection of penicillin is administered:

(*a*) *How many moles of penicillin are administered?*

(*b*) *How many molecules of penicillin are administered?*

(*c*) *How many atoms of nitrogen are injected?*

Solution

(*a*) Molar mass of penicillin ($C_{16}H_{18}N_2SO_4$)

$$= (16 \times 12 + 18 \times 1 + 2 \times 14 + 1 \times 32 + 4 \times 16) \text{ g mol}^{-1}$$

$$= 334 \text{ g mol}^{-1}$$

334 g of penicillin = 1 mol

0.500 g of penicillin = $\frac{0.500}{334}$ mol = $\mathbf{1.497 \times 10^{-3}}$ **mol** (*b*) 1 mole of penicillin = 6.023×10^{23} molecules of penicillin.

1.497×10^{-3} moles of penicillin contain $6.023 \times 10^{23} \times 1.497 \times 10^{-3}$ molecules = 9.016×10^{20} molecules of penicillin.

(*c*) 1 molecule of penicillin contains 2 atoms of nitrogen.

9.016×10^{20} molecules of penicillin contain $2 \times 9.016 \times 10^{20} = 1.8032 \times 10^{21}$ molecules of nitrogen.

Example 12

Calculate the number of molecules of chloroform ($CHCl_3$) weighing 0.0239 g (H = 1, C = 12, Cl = 35.5).

Solution

Molar mass of $CHCl_3$

$$= 12 + 1 + 3 \times 35.5 = 119.5 \text{ g mol}^{-1}$$

119.5 g of $CHCl_3$ contain 1 mole of molecules of chloroform = 6.023×10^{23} molecules. 0.0239 g of $CHCl_3$ contains

$$\frac{6.023 \times 10^{23} \times 0.0239}{119.5} \text{ molecules}$$

$$= 1.2046 \times 10^{20} \text{ molecules}$$

Number of molecules in 0.0239 g of $CHCl_3$ is 1.2046×10^{20}.

Example 13

Find the number of atoms in the following:

(*a*) *52 mol of Ne*

(*b*) *52 g of Ne.*

Solution

(*a*) 1 mole of Ne contains 6.023×10^{23} atoms of Ne.

52 moles of Ne contain $6.023 \times 10^{23} \times 52$ atoms of Ne = 3.132×10^{25} atoms of Ne.

(*b*) 1 mole of Ne weighs 20 g and has 6.023×10^{23} atoms of Ne, *i.e.*, 20 g of Ne contain 6.023×10^{23} atoms of Ne.

52 g of Ne contain $\dfrac{6.023 \times 10^{23} \times 52}{20}$

atoms of Ne = 1.57×10^{24} atoms of Ne

Example 14

Calculate the mass in grams of

(*a*) *1 atom of nitrogen*

(*b*) *1 atom of silver*

(*c*) *1 molecule of benzene (C_6H_6).*

Solution

(*a*) Molecular mass of nitrogen = 14 u

Molar mass of nitrogen = 14.007 g mol^{-1}

Number of nitrogen atoms in 1 mole = 6.023×10^{23}

Mass of 1 atom of nitrogen

$$= \frac{14.007 \text{ g mol}^{-1}}{6.023 \times 10^{23} \text{ mol}^{-1}}$$

$$= \mathbf{2.33 \times 10^{-23} \text{ g}}$$

(*b*) Molecular mass of silver = 107.87 u

Molar mass of silver = 107.87 g mol^{-1}

Mass of 1 atom of silver

$$= \frac{\text{Molar mass}}{\text{Avogadro's number}}$$

$$= \frac{107.87 \text{ g mol}^{-1}}{6.023 \times 10^{23} \text{ mol}^{-1}}$$

$$= \mathbf{1.791 \times 10^{-23} \text{ g}}$$

(*c*) Molecular mass of benzene (C_6H_6)

$$= 12 \times 6 + 1 \times 6 = 72 + 6 = 78 \text{ u}$$

Mass of 1 molecule of benzene

$$= \frac{\text{Molar mass of benzene}}{\text{Avogadro's number}}$$

$$= \frac{78 \text{ g mol}^{-1}}{6.023 \times 10^{23} \text{ mol}^{-1}}$$

$$= \mathbf{12.95 \times 10^{-23} \text{ g}}$$

Example 15

Arrange the following in the decreasing order of masses:

(*a*) *one atom of gold*

(*b*) *one gram-atom of nitrogen*

(*c*) *one mole of calcium*

(*d*) *one gram of iron.*

Given:

Molar mass of gold = 196.97 u

Molar mass of calcium = 40 u

Solution

(*a*) Molecular mass of gold = 196.97 u

Molar mass of gold = 196.97 g mol^{-1}

Mass of one atom of gold

$$= \frac{\text{Molar mass}}{\text{Avogadro's number}}$$

$$= \frac{196.97 \text{ g mol}^{-1}}{6.023 \times 10^{23} \text{ mol}^{-1}}$$

$$= \mathbf{3.27 \times 10^{-22} \text{ g}}$$

(*b*) Mass of one gram-atom of nitrogen = **14.0 g**

(*c*) Mass of one mole of calcium = Molar mass of calcium in gram = 40 g

(*d*) Given: mass of iron = 1.0 g

Hence the order of decreasing mass is:

one mole of calcium > one gram-atom of nitrogen > one gram of iron > one atom of gold.

Example 16

Calculate

(*a*) *the number of molecules of sulphur* (S_8) *in 16 g of solid sulphur*

(*b*) *the number of aluminium ions in 0.051 g of aluminium oxide.*

Solution

(*a*) No. of moles of S_8 in 16 g of sulphur

$$= \frac{\text{Mass of sulphur}}{\text{Molar mass of } S_8} = \frac{16 \text{ g}}{32 \times 8 \text{ g mol}^{-1}}$$

$$= \frac{16}{256} \text{ mol}$$

No. of S_8 molecules

= No. of moles × Avogadro's number

$$= \frac{16}{256} \text{ mol} \times 6.023 \times 10^{23} \text{ molecules mol}^{-1}$$

$$= \mathbf{3.76 \times 10^{22} \text{ molecules}}$$

(*b*) Molar mass of Al_2O_3

= (2 × 27 + 3 × 16) g mol^{-1}

= 102 g mol^{-1}

No. of moles of Al_2O_3

$$= \frac{\text{Mass of } Al_2O_3}{\text{Molar mass of } Al_2O_3}$$

$$= \frac{0.051 \text{ g}}{102 \text{ g mol}^{-1}} = 5.00 \times 10^{-4} \text{ mol}$$

$$\underset{1 \text{ mol}}{Al_2O_3} \longrightarrow \underset{2 \text{ mol}}{2Al^{3+}}$$

1 mol of Al_2O_3 contains 2 × 6.023 × 10^{23} ions of aluminium.

5.00 × 10^{-4} mol of Al_2O_3 contains 2 × 6.023 × 10^{23} × 5.00 × 10^{-4} ions of aluminium = 6.0 × 10^{20} ions

Hence, 0.051 g of Al_2O_3 contains 6.0 × 10^{20} ions of aluminium.

Example 17

Chlorophyll, the green colouring matter in plants involved in photosynthesis, contains 2.7% magnesium by mass. Calculate the number of magnesium atoms in 1.0 g of chlorophyll.

Solution

100 g of chlorophyll contains 2.7 g of magnesium.

1 g of chlorophyll contains 0.027 g of magnesium.

1 mole of Mg = 24 g of magnesium

= 6.023×10^{23} atoms of magnesium.

2.7×10^{-2} g of magnesium contains

$$\frac{6.023 \times 10^{23} \times 2.7 \times 10^{-2}}{24}$$

= **6.78×10^{20} atoms of magnesium**

Example 18

Calculate the mass of carbon dioxide which contains the same number of molecules as are contained in 40 g of oxygen.

Solution

Molar mass of O_2 = 32 g mol^{-1}

32 g of O_2 contain 6.023×10^{23} molecules.

40 g of O_2 contain $\frac{6.023 \times 10^{23} \times 40}{32}$

= 7.529×10^{23} molecules

Mass of 6.023×10^{23} molecules of CO_2

= 44 g

Mass of 7.529×10^{23} molecules of CO_2

$$= \frac{44 \times 7.529 \times 10^{23}}{6.023 \times 10^{23}} = \mathbf{55\ g}$$

Example 19

Calculate the number of oxygen atoms present in 88 g of CO_2. What would be the mass of CO having the same number of oxygen atoms?

Solution

Number of moles of CO_2 in 88 g of CO_2

$$= \frac{\text{Mass of } CO_2}{\text{Molar mass of } CO_2}$$

$$= \frac{88\ \text{g}}{44\ \text{g mol}^{-1}} = 2\ \text{mol}$$

Since one mole of CO_2 contains two moles of oxygen atoms, two moles of CO_2 contain four moles of oxygen atoms. Hence, as

1 mole of oxygen atoms contains 6.023×10^{23} oxygen atoms.

4 moles of oxygen atoms contain $6.023 \times 10^{23} \times 4 = 2.4092 \times 10^{24}$ oxygen atoms.

Since 1 mole of oxygen atoms is present in 1 mole of CO, 4 moles of oxygen atoms are present in 4 moles of CO.

Mass of 4 moles of CO = Number of moles of CO × gram-molecular mass of CO

= 4 × (12 + 16) g = **112 g**

Example 20

How many molecules are present in

(*a*) *9 g of water*

(*b*) *17 g of ammonia?*

Solution

(*a*) Given: mass of water = 9 g

Molar mass of water (H_2O)

= (2 × 1 + 1 × 16) g mol^{-1} = 18 g mol^{-1}

18 g of water contain 6.023×10^{23} molecules

9 g of water contain $\frac{6.023 \times 10^{23} \times 9\ \text{g}}{18\ \text{g}}$

= **3.011×10^{23} molecules**

(*b*) Given: mass of ammonia = 17 g

Molar mass of ammonia (NH_3)

= (1 × 14 + 3 × 1) g mol^{-1}

= 17 g mol^{-1}

17 g of ammonia contain 6.023×10^{23} molecules.

Example 21

What is the mass of

(*a*) *0.2 mol of oxygen atoms*

(*b*) *0.5 mol of water molecules?*

Solution

(*a*) 0.2 mole of oxygen atoms

= 0.2 mol × molar mass of oxygen atom

= 0.2 mol × 16 g mol^{-1} = **3.2 g**

(*b*) 0.5 mole of water molecules

= 0.5 mol × molar mass of water

= 0.5 mol × 18 g mol^{-1} = **9 g**

Example 22

What is the mass of

(*a*) *1 mole of N atoms ?*

(*b*) *4 moles of Al atoms ?*

(*c*) *1 mole of* Na^+ *ions ?*

(*d*) *10 moles of* Na_2SO_3 *?*

Solution

(*a*) Mass of 1 mole of N atoms

= Mass of 6.023×10^{23} atoms of N

= Gram-atomic mass of N = **14 g**

(*b*) Mass of 4 moles of Al atoms

= Mass of $4 \times 6.023 \times 10^{23}$ atoms of Al

= 4 × gram-atomic mass of Al

= 4 × 27 g = **108 g**

(*c*) Mass of 1.5 mole of Na^+ ions

= Mass of 1.5x6.023 x 10 23ions of Na+

= 1.5x 23g = 34.5g

= Gram-atomic mass of Na = **23 g**

(*d*) Mass of 10 moles of Na_2SO_3

= 10 mol × molar mass of Na_2SO_3

= 10 mol × (2 × 23 g mol^{-1} + 1 × 32 g mol^{-1} + 3 × 16 g mol^{-1})

= 10 mol × (46 g mol^{-1} + 32 g mol^{-1} + 48 g mol^{-1})

= 10 mol × 126 g mol^{-1} = **1260 g**

Example 23

Convert into moles

(*a*) *12 g of oxygen gas*

(*b*) *20 g of water*

(*c*) *22 g of carbon dioxide.*

Solution

(*a*) Given mass of oxygen gas = 12 g

Molar mass of oxygen gas

= 2 × 16 g mol^{-1} = 32 g mol^{-1}

No. of moles of oxygen gas

$$= \frac{\text{Mass of oxygen gas}}{\text{Molar mass of oxygen gas}}$$

$$= \frac{12\text{ g}}{32\text{ g mol}^{-1}} = \frac{12}{32}\text{ mol} = 0.375\text{ mol}$$

Hence, 12 g of oxygen gas is equal to 0.375 mol of oxygen gas.

(*b*) Given: mass of water = 20 g

Molar mass of water (H_2O)

= (2 × 1 g mol^{-1} + 1 × 16 g mol^{-1})

= 18 g mol^{-1}

No. of moles of water

$$= \frac{\text{Mass of water}}{\text{Molar mass of water}}$$

$$= \frac{20\text{ g}}{18\text{ g mol}^{-1}} = 1.11\text{ mol}$$

Hence, 20 g of water is equal to 1.11 mol of water.

(*c*) Given: mass of carbon dioxide = 22 g

Molar mass of carbon dioxide (CO_2)

= (1 × 12 g mol^{-1} + 2 × 16 g mol^{-1})

= 44 g mol^{-1}

No. of moles of carbon dioxide

$$= \frac{\text{Mass of carbon dioxide}}{\text{Molar mass of carbon dioxide}}$$

$$= \frac{22 \text{ g}}{44 \text{ g mol}^{-1}} = 0.5 \text{ mol}$$

Hence, 22 g of carbon dioxide is equal to 0.5 mole of carbon dioxide.

Example 24

Calculate the number of moles in the following masses

(a) *7.85 g Fe (at. mass 56)*

(b) *65.5 μg of Carbon (at. mass 12)*

(c) *4.68 mg of Si (at. mass 28)*

(d) *1.46 metric tons of Al (at. mass 27)*

(e) *7.9 mg of Ca (at. mass 40).*

(Note: 1 metric ton = 10^3 kg).

Solution

Number of moles

$$= \frac{\text{mass in g}}{\text{Molar mass of the element}},$$

(*a*) $\frac{7.85}{56} = \mathbf{0.1401}$

(*b*) $\frac{65.5\times10^{-6}}{12} = \mathbf{5.46 \times 10^{-6}}$

(*c*) $\frac{4.68\times10^{-3}}{28} = \mathbf{1.67 \times 10^{-4}}$

(*d*) $\frac{1.46\times10^{6}}{27} = \mathbf{5.41 \times 10^{4}}$

(*e*) $\frac{7.9\times10^{-3}}{40} = \mathbf{1.975 \times 10^{-4}}$

Example 25

(*a*) *Calculate the mass of 2.5 moles of calcium. Atomic mass of calcium is 40.*

(*b*) *Calculate the mass of 1.5 mole of water (H_2O).*

Solution

(*a*) 1 mole of calcium

= molar mass of calcium atoms

= 40 g mol^{-1}

2.5 moles of calcium

= 40 × 2.5 = **100 g**

(*b*) Molecular mass of water (H_2O)

= 1 × 2 + 16 = 18 u

= Molar mass of H_2O

= 18 g mol^{-1}.

1.5 mole of H_2O

= 18 × 1.5 = **27 g**

Example 26

Calculate the number of gram-atom and gram-mole in 25.4 mg of iodine (I_2). Atomic mass of I = 127 u.

Solution

$$\text{Gram atom} = \frac{\text{Mass (g)}}{\text{Molar mass of } I_2}$$

$$= \frac{25.4\times10^{-3} \text{ g}}{127 \text{ g}}$$

= **2 × 10^{-4} g-atom**

$$\text{Gram mole} = \frac{\text{Mass (g)}}{\text{Molar mass}} = \frac{25.4\times10^{-3} \text{ g}}{254 \text{ (g mol}^{-1})}$$

= **1 × 10^{-4} g-mole**

Example 27

Calculate the molar mass of glucose ($C_6H_{12}O_6$) and the number of atoms of each kind in it.

Solution

Molecular mass of glucose ($C_6H_{12}O_6$)

= 6 × (12.011 *u*) + 12 × (1.008 *u*) + 6 × (16.00 *u*)

= 72.066 *u* + 12.096 *u* + 96.00 *u*

= **180.162 u**

Calculation of number of atoms of each kind

1 mole of glucose ($C_6H_{12}O_6$)

$\equiv$ 6 moles of carbon + 12 moles of hydrogen + 6 moles of oxygen

Hence, Atoms of carbon

$= 6 \times 6.02 \times 10^{23}$

$= \mathbf{36.12 \times 10^{23}}$

Atoms of hydrogen

$= 12 \times 6.02 \times 10^{23}$

$= \mathbf{72.24 \times 10^{23}}$

Atoms of oxygen

$= 6 \times 6.02 \times 10^{23}$

$= \mathbf{36.12 \times 10^{23}}$

Example 28

Calculate mass in grams of the following:

(*i*) *one atom of calcium*

(*ii*) *one molecule of sulphur dioxide* (SO_2).

Solution

(*i*) Mass of 6.022×10^{23} atoms of calcium

= gram atomic mass of calcium

= 40 g

$\therefore$ Mass of 1 atom of calcium

$$= \frac{40g}{6.022 \times 10^{23}}$$

$= \mathbf{6.6 \times 10^{-23}\ g}$

(*ii*) Mass of 6.022×10^{23} molecules of SO_2

= molar mass of SO_2

= 64 g

$\therefore$ Mass of 1 molecule of SO_2

$$= \frac{64g}{6.022 \times 10^{23}}$$

$= \mathbf{1.06 \times 10^{-22}\ g}$

Example 29

(*a*) *Calculate number of atoms in each of the following*:

(*i*) *0.5 mole of nitrogen atoms*

(*ii*) *0.2 mole of nitrogen molecules*

(*iii*) *3.2 g of sulphur.*

(*b*) *Calculate number of molecules in each of the following*:

(*i*) *14 g of nitrogen*

(*ii*) *3.4 g of hydrogen sulphide* (H_2S).

Solution

(*a*)(*i*) 1 mole of nitrogen atom

$= 6.022 \times 10^{23}$ atoms

$\therefore$ 0.5 mole of nitrogen atoms

$= 6.022 \times 10^{23} \times 0.5$

$= \mathbf{3.011 \times 10^{23}}$ **atoms**

(*ii*) 1 mole of nitrogen molecule

$= 6.022 \times 10^{23}$ molecules

$\therefore$ 0.2 mole of nitrogen molecule

$= 6.022 \times 10^{23} \times 0.2$

$= 1.2044 \times 10^{23}$ molecules.

1 molecule of nitrogen

= 2 atoms

1.2044×10^{23} molecules of nitrogen

$= 1.2044 \times 10^{23} \times 2$

$= \mathbf{2.409 \times 10^{23}}$ **atoms**

(*iii*) 32 g sulphur contain

$= 6.022 \times 10^{23}$ atoms

$\therefore$ 3.2 g of sulphur contain

$$= \frac{6.022 \times 10^{23} \times 3.2}{32}$$

$= \mathbf{6.022 \times 10^{22}}$ **atoms**

(*b*) (*i*) Molar mass of nitrogen

$= 28\ g\ mol^{-1}$

28 g of nitrogen contain

$= 6.022 \times 10^{23}$ molecules

$\therefore$ 14 g of nitrogen contain

$$= \frac{6.022 \times 10^{23}}{28} \times 14$$

$= \mathbf{3.01 \times 10^{23}}$ **molecules**

(*ii*) Molar mass of H_2S

$= (2 + 32)$ g $= 34$ g mol^{-1}

34 g of H_2S contain

$= 6.022 \times 10^{23}$ molecules

3.4 g of H_2S contain

$$= \frac{6.022 \times 10^{23}}{34} \times 3.4$$

$= \mathbf{6.023 \times 10^{22}}$ **molecules**

Example 30

How many atoms of oxygen are present in 300 g of $CaCO_3$?

Solution

Gram formula mass of $CaCO_3$ = 100 g

Now, 1 mole of $CaCO_3$ contains

= 3 mole of O atoms.

or 100 g of $CaCO_3$ contain

$= 3 \times 6.022 \times 10^{23}$ O atoms

$\therefore$ 300 g of $CaCO_3$ contain O atoms

$$= \frac{3 \times 6.022 \times 10^{23}}{100} \times 300$$

$= 54.198 \times 10^{23}$

$= \mathbf{5.4198 \times 10^{24}}$ **Oxygen atoms**

Example 31

How many molecules of water of hydration are present in 252 mg of oxalic acid, ($H_2C_2O_4.2H_2O$)?

Solution

Molar mass of $H_2C_2O_4.2\,H_2O$ = 126 g mol^{-1}

Now, water molecules in 1 mol of oxalic acid

= 2 mol.

or water molecules in 126 g of oxalic acid

$= 2 \times 6.022 \times 10^{23}$

$\therefore$ Water molecules in 252×10^{-3} g of oxalic acid

$$= \frac{2 \times 6.022 \times 10^{23} \times 252 \times 10^{-3}}{126}$$

$= \mathbf{2.4 \times 10^{21}}$

Example 32

Calculate mass of sodium which contains same number of atoms as are present in 4 g of calcium. Atomic masses of sodium and calcium are 23 and 40 respectively.

Solution

40 g of calcium contain $= 6.023 \times 10^{23}$ atoms

$\therefore$ 4 g of calcium contain

$$= \frac{6.022 \times 10^{23}}{40} \times 4$$

$= 6.022 \times 10^{22}$ atoms

Now, 6.022×10^{23} atoms of sodium have mass

= 23 g

$\therefore$ 6.022×10^{22} atoms of sodium have mass

$$= \frac{23}{6.022 \times 10^{23}} \times 6.022 \times 10^{22}$$

= **2.3 g**

Example 33

Chlorophyll, the green colouring matter of plants contains 2.68% of magnesium by mass. Calculate the number of magnesium atoms in 5.00 g of this complex.

Solution

Mass of magnesium in 5.00 g of complex

$$= \frac{2.68}{100} \times 5.00 = 0.134 \text{ g.}$$

Gram atomic mass of magnesium

= 24 g

24 g of magnesium contain

$= 6.022 \times 10^{23}$ atoms

0.134 g of magnesium would contain

$$= \frac{6.022 \times 10^{23}}{24} \times 0.134$$

$= 3.36 \times 10^{21}$ atoms

Therefore, 5.00 g of the given complex would contain **3.36×10^{21} atoms of magnesium.**

Example 34

Calculate number of atoms of each type in 3.42 g of sucrose ($C_{12}H_{22}O_{11}$).

Solution

Molecular mass of sucrose

$= (12 \times 12 + 1 \times 22 + 16 \times 11)$

$= 342$ u

342 g of sucrose contain

$= 6.022 \times 10^{23}$ molecules

$\therefore$ 3.42 g of sucrose contain

$$= \frac{6.022 \times 10^{23}}{342} \times 3.42$$

$= 6.022 \times 10^{21}$ molecules.

Number of atoms of carbon in 3.42 g of sucrose

1 molecule of sucrose contains

= 12 atoms of carbon

6.022×10^{21} molecules of sucrose contain

$= 12 \times 6.022 \times 10^{21}$ atoms of carbon

$= 7.226 \times 10^{22}$ atoms of carbon

Number of atoms of hydrogen in 3.42 g of sucrose

1 molecule of sucrose contains

= 22 atoms of hydrogen

6.022×10^{21} molecules of sucrose contain

$= 22 \times 6.022 \times 10^{21}$ atoms of hydrogen

$= 1.324 \times 10^{23}$ atoms of hydrogen

Number of atoms of oxygen in 3.42 g of sucrose

1 molecule of sucrose contains

= 11 atoms of oxygen

6.022×10^{21} molecules of sucrose contain

$= 11 \times 6.022 \times 10^{21}$ atoms of oxygen.

$= 6.624 \times 10^{22}$ atoms of oxygen

7.8 CALCULATION OF MASS PER CENT COMPOSITION OF AN ELEMENT IN A COMPOUND

A compound contains two or more elements combined in a certain fixed ratio. The percentage composition of a compound is the mass of each element of the compound, present in 100 g of that compound, *i.e.*, the mass percentage of each element present in the compound. The mass percentage of each element in a compound can be calculated using either of the following two equations.

When the masses of compound and each element are given

Mass percentage of an element can be obtained if we know the mass of that element in a known mass of the compound.

Mass percentage of an element X

$$= \frac{\text{Mass of element in the given mass of the compound}}{\text{Total mass of the compound}} \times 100$$

When the formula of the compound and the atomic masses of the elements are given

When the formula of the compound and the atomic masses of the elements are given, the molecular mass of the compound can be calculated by adding the masses of all the elements present in the compound, and then the mass percentage of each element can be calculated using the following formula:

Mass percentage of an element

$$= \frac{\text{Total mass of the element in one molecule of the compound}}{\text{Molecular mass of the compound}} \times 100$$

The above two methods of calculating the percentage composition of a compound are given in the following examples:

Example 35

0.24 g sample of a compound of oxygen and boron was found by analysis to contain 0.096 g of boron and 0.144 g of oxygen. Calculate the percentage composition of the compound by mass.

Solution

Percentage of boron

$$= \frac{\text{Mass of boron} \times 100\%}{\text{Mass of the compound}} = \frac{0.096\ \text{g} \times 100}{0.24\text{g}}$$

= 40.0%

Percentage of oxygen

$$= \frac{\text{Mass of oxygen} \times 100\%}{\text{Mass of the compound}} = \frac{0.144\ \text{g} \times 100}{0.24\ \text{g}}$$

= 60.0%

Hence, the mass percentages of boron and oxygen in the given compound are 40.0% and 60.0% respectively.

Example 36

Calculate the mass percentage of each element present in water.

Solution

Molecular mass of water (H_2O)

$$= 2 \times 1\ \text{u} + 16\ \text{u} = 18\ \text{u}$$

Mass percentage of hydrogen (H) in water

$$= \frac{2\ \text{u}}{18\ \text{u}} \times 100 = 11.11$$

Mass percentage of oxygen (O) in water

$$= \frac{16\ u}{18\ u} \times 100 = 88.89$$

Hence, the composition of water by mass is H = 11.11% and O = **88.89%.**

Example 37

Find the percentage composition of glucose whose formula is $C_6H_{12}O_6$.

Solution

Molecular mass of glucose ($C_6H_{12}O_6$)

$$= 6 \times 12\ \text{u} + 12 \times 1\ \text{u} + 6 \times 16\ \text{u}$$

$$= 72\ \text{u} + 12\ \text{u} + 96\ \text{u} = 180\ \text{u}$$

Mass percentage of carbon (C) in glucose

$$= \frac{72\ \text{u}}{180\ \text{u}} \times 100 = 40.0$$

Mass percentage of hydrogen (H) in glucose

$$= \frac{12\ \text{u}}{180\ \text{u}} \times 100 = 6.67$$

Mass percentage of oxygen (O) in glucose

$$= \frac{96\ \text{u}}{180\ \text{u}} \times 100 = 53.33$$

Hence, glucose contains 40.0% carbon, 6.67% hydrogen and 53.3% oxygen.

Example 38

Calculate the percentage of water of crystallisation in washing soda whose formula is $Na_2CO_3.10H_2O$.

Solution

Molecular mass of $Na_2CO_3.10H_2O$

$$= 2 \times 23\ \text{u} + 12\ \text{u} + 3 \times 16\ \text{u} + 10\ (2 \times 1\ \text{u} + 1 \times 16\ \text{u})$$

$$= 46\ \text{u} + 12\ \text{u} + 48\ \text{u} + 180\ \text{u} = 286\ \text{u}$$

286 g of washing soda contain 180 g of water of crystallisation.

100 g of washing soda will contain $\frac{180 \text{ u}}{286 \text{ u}} \times$

100 g water of crystallisation = 62.94 g

The amount of water of crystallisation in washing soda = 62.94% by mass.

Example 39

Find the percentage composition of potassium permanganate.

Solution

Molecular mass of $KMnO_4$

$= 39 + 55 + 16 \times 4$

$= 39 + 55 + 64 = 158$ u

Percentage of potassium

$$= \frac{39 \times 100}{158} = 24.68$$

Percentage of manganese

$$= \frac{55 \times 100}{158} = 34.81$$

Percentage of oxygen

$$= \frac{64 \times 100}{158} = 40.51$$

The percentage composition of potassium permanganate is as follows:

Potassium = **24.68%**

Manganese = **34.81%**

Oxygen = **40.51%**

Example 40

Ferric sulphate is a crystalline compound of iron. It is used in water and sewage treatment to help the removal of suspended impurities. Calculate the mass percentage of iron, sulphur and oxygen in the compound.

Solution

The formula of the compound is $Fe_2(SO_4)_3$.

The formula mass $= 2 \times 56 + 3(32 + 64)$

$= 400$ u

$\therefore$ Percentage of iron

$$= \frac{2 \times 56 \times 100}{400} = \mathbf{28\%}$$

Percentage of sulphur

$$= \frac{3 \times 32 \times 100}{400} = \mathbf{24\%}$$

Percentage of oxygen

$$= \frac{3 \times 64 \times 100}{400} = \mathbf{48\%}$$

Example 41

Calculate the mass per cent of different elements present in ethyl alcohol (C_2H_5OH).

Solution

Molar mass of ethanol

$= 2 \times 12.01 + 6 \times 1.008 + 16.00$

$= 46.068$ g

Mass per cent of carbon

$$= \frac{24.02}{46.068} \times 100 = \mathbf{52.14\%}$$

Mass per cent of hydrogen

$$= \frac{6.048}{46.068} \times 100 = \mathbf{13.13\%}$$

Mass per cent of oxygen

$$= \frac{16.00}{46.068} \times 100 = \mathbf{34.73\%}$$

Example 42

Calculate the percentage of hydration, total oxygen and copper in copper (II) sulphate pentahydrate.

Solution

The formula of the compound is $CuSO_4.5H_2O$

Formula mass $= 63.5 + 32 + 64 + 5 \times 18$

$= 249.5$ u

Percentage of hydration $= \frac{5 \times 18 \times 100}{249.5} = \mathbf{36.07\%}$

Percentage of total oxygen $= \frac{9 \times 16 \times 100}{249.5} = \mathbf{57.71\%}$

Percentage of copper $= \frac{63.3 \times 100}{249.5} = \mathbf{25.45\%}$

EXPERIMENT 1

Aim

Determining the Per cent Composition of a magnesium in magnesium oxide

Safety

- The burning of magnesium generates an intense white light. Mixed in with the white light there is ultraviolet light as well.
- Do not look at the magnesium as it is burning.
- Wear goggles

Apparatus and Material

Stand, ceramic tile hot plate, ring clamp, pipe clay triangle, crucible and lid, crucible tongs, Bunsen burner, 25 cm of Magnesium ribbon

Procedure

1. Obtain a 25 cm piece of magnesium ribbon. The exact length is not important.
2. Find the mass of the magnesium ribbon and record it in the data table.
3. Set up the apparatus with the stand, ring clamp, Bunsen burner sitting on the ceramic tile hot plate.
4. Find the mass of the empty crucible and record it in the data table.
5. Roll the ribbon of magnesium into a spiral or crush it into a compacted mass and place it in the crucible.
6. Cover the crucible and heat strongly for 5 minutes or until red hot.
7. With the crucible tongs gently lift the crucible lid to let air into the crucible.
8. With the lid left slightly open discontinue heating.
9. Remove the crucible lid and let the crucible cool before massing it again.
10. Mass the crucible and contents. The contents are your new compound of magnesium oxide. Record the mass on the data table.

Data Table

Mass of magnesium ribbon = __________

Mass of crucible = __________

Mass of crucible and contents = __________

Calculations

1. Find the mass of the contents of the crucible. This is your new compound of magnesium oxide.
2. The percentage of magnesium in the compound can be found like this:

$$\text{Percentage of magnesium} = \frac{\text{Mass of magnesium ribbon}}{\text{Mass of magnesium oxide compound}} \times 100$$

3. Find the percentage of oxygen in the compound.

Conclusions

The percentage composition of magnesium oxide is ......

Error Analysis

The above answers are your actual values.

If the formula for magnesium oxide is MgO use your periodic table to calculate the theoretical percentage composition of both the magnesium and the oxygen.

$$\%\ \text{error} = \frac{\text{Actual value} - \text{Theoretical value}}{\text{Theoretical value}} \times 100$$

Calculate the percentage composition for MgO: __________

Calculate the % error for O: __________

Calculate the % error for Mg:

EXERCISE 7.6

1. Calculate the mass percentage of each element present in water.
2. Find the percentage composition of each element in glucose whose formula is $C_6H_{12}O_6$.
3. Calculate the percentage of water of crystallisation in washing soda whose formula is $Na_2CO_3.10H_2O$.
4. Find the percentage composition of each element in potassium permanganate whose formula is $KMnO_4$(Atomic masses , H=1, C=12, O=16, K=39, Mn=55)

7.9 EMPIRICAL FORMULA AND MOLECULAR FORMULA

An **empirical formula** represents the simplest whole number ratio of various atoms present in a compound whereas the **molecular formula** shows the exact number of different types of atoms present in a molecule of a compound.

Relation between the Two Formulae

Molecular formula is whole number multiple of empirical formula. Thus,

Molecular formula = Empirical formula × n

where n = 1, 2, 3...

$$n = \frac{\text{Molecular formula}}{\text{Empirical formula}}$$

$$= \frac{\text{Molecular mass}}{\text{Empirical formula mass*}}$$

* **Empirical formula mass** of a substance is equal to the sum of atomic masses of all the atoms.

Note: In many compounds; empirical formula is same as molecular formula. For example, NH_3, H_2O, CH_4, etc.

If the mass per cent of various elements present in a compound is known, its empirical formula can be determined. Molecular formula can further be obtained if the molar mass is known. The following example illustrates this sequence.

EXAMPLE

A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine by mass. Its molar mass is 98.96 g. What are its empirical and molecular formulas?

Solution

Step 1: Conversion of mass per cent to grams.

Since we are having mass per cent, it is convenient to use 100 g of the compound as the starting material. Thus, in the 100 g sample of the above compound, 4.07 g hydrogen is present, 24.27 g carbon is present and 71.65 g chlorine is present.

Step 2: Convert into number moles of each element.

Divide the masses obtained above by respective atomic masses of various elements.

Moles of hydrogen $= \frac{4.07\ \text{g}}{1.008\ \text{g}} = 4.04$

Moles of carbon $= \frac{24.27\ \text{g}}{12.01\ \text{g}} = 2.021$

Moles of chlorine $= \frac{71.65\ \text{g}}{35.435\ \text{g}} = 2.021$

Step 3: Divide the mole value obtained above by the smallest number

Since 2.021 is smallest value, division by it gives a ratio of 2 : 1 : 1 for H : C : C : Cl.

In case the ratios are not whole numbers, then they may be converted into whole number by multiplying by the suitable co-efficient.

Step 4: Write empirical formula by mentioning the numbers after writing the symbols of respective elements.

CH_2Cl is, thus, the empirical formula of the above compound.

Step 5: Writing molecular formula

(*a*) Determine empirical formula mass

Add the atomic masses of various atoms present in the empirical formula.

For CH_2Cl, empirical formula mass is 12.01 + 2 × 1.008 + 35.453 = 49.48 g

(*b*) Divide Molar mass by empirical formula mass

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96\ \text{g}}{49.48\ \text{g}}$$

$$= 2 = (n)$$

(*c*) Multiply empirical formula by n obtained above to get the molecular formula

Empirical formula = CH_2Cl, n = 2. Hence molecular formula is $C_2H_4Cl_2$.

Steps for Writing the Empirical Formula

The percentage of the elements in the compound is determined by suitable methods and from the data collected, the empirical formula is determined by the following steps:

Step 1: Divide the percentage of each element by its atomic mass. This gives the relative number of moles of various elements present in the compound.

Step 2: Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.

Step 3: Multiply the figures, so obtained by a suitable integer, if necessary, in order to obtain whole number ratio.

Step 4: Finally, write down the symbols of the various elements side by side and put the above numbers as the subscripts to the lower right hand corner of each symbol. This will represent the **empirical formula** of the compound.

Steps for Writing the Molecular Formula

Step 1: Calculate the empirical formula as described above.

Step 2: Find out the empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula of the compound.

Step 3: Divide the molecular mass (determined experimentally by some suitable method) by the empirical formula mass and find out the value of *n*.

Step 4: Multiply the empirical formula of the compound with *n* so as to find out the **molecular formula** of the compound.

Example 43

Write the empirical formula of the compounds having molecular formulae:

(*i*) C_6H_6	(*ii*) C_6H_{12}	(*iii*) H_2O_2	(*iv*) H_2O	(*v*) Na_2CO_3
(*vi*) B_2H_6	(*vii*) N_2O_4	(*viii*) H_3PO_4	(*ix*) Fe_2O_3	(*x*) C_2H_2.

Solution

Empirical formula is the simplest whole number ratio of atoms in the molecule, therefore the empirical formula of given compounds are:

(*i*) CH	(*ii*) $\mathrm{CH_2}$	(*iii*) HO	(*iv*) $\mathrm{H_2O}$	(*v*) $\mathrm{Na_2CO_3}$
(*vi*) $\mathrm{BH_3}$	(*vii*) $\mathrm{NO_2}$	(*viii*) $\mathrm{H_3PO_4}$	(*ix*) $\mathrm{Fe_2O_3}$	(*x*) CH.

Example 44

A substance, on analysis, gave the following percentage composition: Na = 43.4%, C = 11.3%, O = 45.3%. Calculate its empirical formula. [*Na = 23, C = 12, O = 16*]

Solution

Element	Percentage	At. Mass	Relative No. of Moles	Simple Ratio	Simplest Whole No. Ratio
Sodium	43.4	23	$\frac{43.4}{23} = 1.88$	$\frac{1.88}{0.94} = 2$	2
Carbon	11.3	12	$\frac{11.3}{12} = 0.94$	$\frac{0.94}{0.94} = 1$	1
Oxygen	45.3	16	$\frac{45.3}{16} = 2.83$	$\frac{2.83}{0.94} = 3$	3

Therefore, the empirical formula is $\mathbf{Na_2CO_3}$.

Example 45

A compound has the following composition: Mg = 9.76%, S = 13.01%, O = 26.01%, H_2O = 51.22%. What is its empirical formula? [Mg = 24, S = 32, O = 16, H = 1]

Solution

Element	Percentage	At. Mass	Relative No. of Moles	Simple Ratio	Simplest Whole No. Ratio
Magnesium	9.76	24	$\frac{9.76}{24} = 0.406$	$\frac{0.406}{0.406} = 1$	1
Sulphur	13.01	32	$\frac{13.01}{32} = 0.406$	$\frac{0.406}{0.406} = 1$	1
Oxygen	26.01	16	$\frac{26.01}{16} = 1.625$	$\frac{1.625}{0.406} = 4$	4
Water	51.22	18 (mol. mass)	$\frac{51.22}{18} = 2.846$	$\frac{2.846}{0.406} = 7$	7

Hence, the empirical formula is $\mathbf{MgSO_4.7H_2O}$.

Example 46

What is the simplest formula of the compound which has the following percentage composition: Carbon 80%, Hydrogen 20%? If the molecular mass is 30, calculate its molecular formula.

Solution

Calculation of empirical formula:

Element	Percentage	At. Mass	Relative No. of Moles	Simple Ratio	Simplest Whole No. Ratio
C	80	12	$\frac{18}{12} = 6.66$	$\frac{6.66}{6.66} = 1$	1
H	20	1	$\frac{20}{1} = 20$	$\frac{20}{6.66} = 3$	3

$\therefore$ Empirical formula is CH_3

Calculation of molecular formula:

Empirical formula mass $= 12 \times 1 + 1 \times 3 = 15$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{30}{15} = 2$$

Molecular formula $=$ Empirical formula $\times 2$

$= CH_3 \times 2 = \mathbf{C_2H_6}$

Example 47

Butyric acid contains C, H, O elements. A 4.24 mg sample of butyric acid is completely burnt in oxygen. It gives 8.45 mg of carbon dioxide and 3.46 mg of water. What is the mass percentage of each element? Determine the empirical and molecular formula of butyric acid if molecular mass of butyric acid is determined to be 88 u.

Solution

Mass of carbon present in 8.45 mg of $CO_2 = \frac{8.45 \times 12}{44}$ mg $=$ **2.30 mg**

Percentage of carbon $= \frac{2.30 \times 100}{4.24} =$ **54.24%**

Mass of hydrogen in 3.46 mg of $H_2O = \frac{3.46 \times 2}{18}$ mg $=$ **0.384 mg**

Percentage of hydrogen $= \frac{0.384 \times 100}{4.24} =$ **9.05%**

Percentage of oxygen = 100 – 54.24 – 9.05 = **36.71%**

Calculation of empirical formula:

Element	Percentage	At. Mass	Relative No. of Moles	Simple Ratio	Simplest Whole No. Ratio
C	54.24	12	$\frac{54.24}{12} = 4.52$	$\frac{4.52}{2.29} = 1.97$	2
H	9.05	1	$\frac{9.05}{1} = 9.05$	$\frac{9.05}{2.29} = 3.95$	4
O	36.71	16	$\frac{36.71}{16} = 2.29$	$\frac{2.29}{2.29} = 1$	1

$\therefore$ Empirical formula is C_2H_4O.

Calculation of molecular formula:

Empirical formula mass $= 12 \times 2 + 1 \times 4 + 16 \times 1 = 44$

Molecular mass $= 88$ u

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{88}{44} = 2$$

Molecular formula = Empirical formula $\times\, n$

$= C_2H_4O \times 2 = \mathbf{C_4H_8O_2}$

Example 48

An organic compound on analysis gave the following data : C = 57.82%, H = 3.6%, and the rest is oxygen. Its vapour density is 83. Find its empirical and molecular formula.

Solution

Calculation of empirical formula:

Element	Percentage	At. Mass	Relative No. of Moles	Simple Ratio	Simplest Whole No. Ratio
C	57.82	12	$\frac{57.82}{12} = 4.818$	$\frac{4.818}{2.4} = 2$	4
H	3.60	1	$\frac{3.60}{1} = 3.60$	$\frac{3.6}{2.4} = 1.5$	3
O	38.58	16	$\frac{38.58}{16} = 2.40$	$\frac{2.4}{2.4} = 1$	2

$\therefore$ Empirical formula is $\mathbf{C_4H_3O_2}$.

Calculation of molecular formula:

Empirical formula mass $= 12 \times 4 + 1 \times 3 + 2 \times 16 = 83$

Molecular mass $= 2 \times \text{V.D.} = 2 \times 83 = 166$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = 2$$

Molecular formula $= \text{Empirical formula} \times n$

$= C_4H_3O_2 \times 2 = \mathbf{C_8H_6O_4}$

Example 49

2.746 g of a compound gave on analysis 1.94 g of silver, 0.268 g of sulphur and 0.538 g of oxygen. Find the empirical formula of the compound. (*At. masses: Ag = 108, S = 32, O = 16*)

Solution

To calculate empirical formula:

Element	Mass in (g)	At. Mass	No. of g-atom	Simple Ratio of g-atom	Simplest Whole No. Ratio
Ag	1.94 g	108	$\frac{1.94}{108} = 0.0179$	$\frac{0.0179}{8.375 \times 10^{-3}} = 2.12$	2
S	0.268 g	32	$\frac{0.268}{32} = 8.375 \times 10^{-3}$	$\frac{8.375 \times 10^{-3}}{8.375 \times 10^{-3}} = 1$	1
O	0.538 g	16	$\frac{0.538}{16} = 0.0336$	$\frac{0.0336}{8.375 \times 10^{-3}} = 4.01$	4

The empirical formula is $\mathbf{Ag_2SO_4}$.

Example 50

A compound on analysis gave the following percentage composition: Na = 14.31%, S = 9.97%, H = 6.22%, O = 69.5%.

Calculate the molecular formula of the compound on the assumption that all the hydrogen in the compound is present in combination with oxygen as water of crystallisation. Molecular mass of the compound is 322. [Na = 23, S = 32, H = 1 and O = 16]

Solution

Calculation of empirical formula:

Element	Percentage	At. Mass	Relative No. of Moles	Simple Ratio	Simplest Whole No. Ratio
Na	14.31	23	$\frac{14.31}{23} = 0.62$	$\frac{0.62}{0.31} = 2$	2
S	9.97	32	$\frac{9.97}{32} = 0.31$	$\frac{0.31}{0.31} = 1$	1
H	6.22	1	$\frac{6.22}{1} = 6.22$	$\frac{6.22}{0.31} = 20$	20
O	69.5	16	$\frac{69.5}{16} = 4.34$	$\frac{4.34}{0.31} = 14$	14

∴ The empirical formula is $Na_2SH_{20}O_{14}$.

Calculation of molecular formula:

Empirical formula mass $= 23 \times 2 + 32 + 20 \times 1 + 16 \times 14 = 322$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{322}{322} = 1$$

Hence, molecular formula $= Na_2SH_{20}O_{14}$

Since all hydrogen is present as H_2O in the compound, it means 20 hydrogen atoms must have combined with 10 atoms of oxygen to form 10 molecules of water of crystallisation. The remaining (14 – 10 = 4) atoms of oxygen should be present with the rest of the compound.

Hence, molecular formula = $\mathbf{Na_2SO_4.10H_2O}$

Example 51

A pure sample of compound is found to contain 2.04 g of sodium, 2.65 × 10^{22} atoms of carbon and 0.132 mole of oxygen atoms. Determine the empirical formula of the compound.

Solution

Let us calculate the molar ratio of each type of atoms.

$$\text{Mole of Na atoms} = \frac{\text{Mass (g)}}{\text{Atomic mass of sodium}} = \frac{2.04\text{ g}}{23\text{ g}} = 0.0887$$

$$\text{Mole of C atoms} = \frac{\text{No. of atoms}}{\text{Avogadro number}} = \frac{2.65 \times 10^{22}}{6.023 \times 10^{23}} = 0.044$$

Mole of O atoms $= 0.132$

The atomic ratio Na :C : O is: 0.0887 : 0.044 : 0.132

On dividing by the least, we get 2 : 1 : 3

Thus, the empirical formula is $\mathbf{Na_2CO_3}$.

EXERCISE 7.7

1. A substance on analysis gave the following composition : C = 40%, H = 6.67%, O = 53.33%. Calculate its empirical formula. [Atomic masses: C = 12, H = 1 and O = 16]
2. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass.
3. Calculate empirical formula of a compound which has 48% carbon, 8% hydrogen, 28% nitrogen and 16% oxygen.

7.10 STOICHIOMETRIC CALCULATIONS

The word 'stoichiometry' is derived from two Greek words – *stoicheion* (meaning *element*) and *metron* (meaning *measure*). Stoichiometry, thus, deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction. Before understanding how to calculate the amounts of reactants required or the products produced in a chemical reaction, let us study what information is available from the **balanced** chemical equation of a given reaction.

Balancing a Chemical Equation

According to the law of conservation of mass, *a balanced chemical equation has the same number of atoms of each element on both sides of the equation.* Many chemical equations can be balanced by *trial* and *error*. Let us take the reactions of a few metals and non-metals with oxygen to give oxides

$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$$

(*a*) balanced equation

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

(*b*) balanced equation

$$P_4(s) + O_2(g) \longrightarrow P_4O_{10}(s)$$

(*c*) unbalanced equation

Equations (*a*) and (*b*) are balanced since there are same number of metal and oxgyen atoms on each side of equations. However, equation (*c*) is not balanced. In this equation, phosphorus atoms are balanced but not the oxygen atoms. To balance it, we must place the coefficient 5 on the left of oxygen on the left side of the equation to balance the oxygen atoms appearing on the right side of the equation.

$$P_4(s) + 5O_2(g) \longrightarrow P_4O_{10}(s)$$

balanced equation

Now let us take combustion of propane, C_3H_8. This equation can be balanced in steps.

Step 1: Write down the correct formulas of reactants and products. Here propane and ogygen are reactants, and carbon dioxide and water are products.

$$C_3H_8(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$$

unbalanced equation

Step 2: *Balance the number of C atoms:* Since 3 carbon atoms are in the reactant, therefore, three CO_2 molecules are required on the right side.

$$C_3H_8(g) + O_2(g) \longrightarrow 3CO_2(g) + H_2O(l)$$

Step 3: *Balance the number of H atoms:* On the left there are 8 hydrogen atoms in the reactant, however each molecule of water has two hydrogen atoms, so four molecules of water will be required for eight hydrogen atoms on the right side.

$$C_3H_8(g) + O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$$

Step 4: *Balance the number of O atoms:* There are ten oxygen atoms on the right side ($3 \times 2 = 6$ in CO_2 and $4 \times 1 = 4$ in water). Therefore, five O_2 molecules are needed to supply the required ten oxygen atoms.

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$$

Step 5: *Verify that the number of atoms of each element is balanced in the final equation.* The equation shows three carbon atoms, eight hydrogen atoms, and ten oxygen atoms on each side.

All equations that have correct formulas for all reactants and products can be balanced. Always remember that subscripts in formulas of reactants and products cannot be changed to balance an equation.

Let us consider the combustion of methane. A balanced equation for this reaction is as given below:

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

Here, methane and oxygen are called *reactants* and carbon dioxide and water are called *products*. Note that all the reactants and the products are gases in the above reaction and this has been indicated by letter (g) in the brackets next to its formula. Similarly, in the case of solids and liquids, (s) and (l) are written respectively.

The coefficients 2 for O_2 and H_2O are called stoichiometric coefficients. Similarly, the coefficient for CH_4 and CO_2 is one in each case. They represent the number of molecules (and moles as well) taking part in the reaction or formed in the reaction.

Thus, according to the above chemical reaction,

- One **mole** of $CH_4(g)$ reacts with two **moles** of $O_2(g)$ to give one **mole** of $CO_2(g)$ and two **moles** of $H_2O(l)$
- One **molecule** of $CH_4(g)$ reacts with 2 **molecules** of $O_2(g)$ to give one **molecule** of $CO_2(g)$ and 2 molecules of $H_2O(g)$
- 22.7 *l* of $CH_4(g)$ reacts with 45.4 1 of $O_2(g)$ to give 22.7 *l* of $CO_2(g)$ and 45.4 *l* of $H_2O(g)$
- 16 g of $CH_4(g)$ reacts with 2×32 g of O_2 (g) to give 44 g of CO_2 (g) and 2×18 g of $H_2O(g)$.

From these relationships, the given data can be interconverted as follows:

mass $\rightleftharpoons$ moles $\rightleftharpoons$ no. of molecules

Example 52

How many grams of oxygen (O_2) are required to completely react with 0.200 g of hydrogen (H_2) to yield water (H_2O)? Also calculate the amount of water formed. (At. mass H = 2; O = 32).

Solution

The balanced equation for the reaction is:

$$\begin{array}{ccccc} 2H_2 & + & O_2 & \rightarrow & 2H_2O \\ 2 \text{ mol} & & 1 \text{ mol} & & 2 \text{ mol} \\ 4 \text{ g} & & 32 \text{ g} & & 36 \text{ g} \end{array}$$

Now, 4 g of H_2 require oxygen

$= 32$ g

0.200 g of H_2 requires oxygen

$$= \frac{32}{4} \times 0.200 = \mathbf{1.6\ g}$$

Again, 4 g of H_2 produce H_2O

$= 36$ g

0.200 g of H_2 produces H_2O

$$= \frac{36}{4} \times 0.200 = \mathbf{1.8\ g}$$

Example 53

What mass of zinc is required to produce hydrogen by reaction with HCl which is enough to produce 4 mol of ammonia according to the reactions

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$
$$3H_2 + N_2 \longrightarrow 2NH_3?$$

Solution

The given equations are:

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$
$$3H_2 + N_2 \longrightarrow 2NH_3$$

From the equations it is clear that

2 mol of NH_3 require = 3 mol of H_2,

and 1 mol of H_2 requires = 1 mol of Zn

or 3 mol of H_2 require = 3 mol of Zn

Thus, 2 mol of NH_3 require

= 3 mol of Zn = 3 × 65 g of Zn

∴ 4 mol of NH_3 require

$$= \frac{3 \times 65}{2} \times 4 = \mathbf{390\ g\ of\ Zn}$$

Example 54

Calculate the amount of water (g) produced by the combustion of 16 g of methane.

Solution

The balanced equation for combustion of methane is:

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

(*i*) 16 g of CH_4 corresponds to one mole.

(*ii*) From the above equation, 1 mol of CH_4 (*g*) gives 2 mol of $H_2O(g)$.

2 mol of water (H_2O)

= 2 × (2 + 16)

= 2 × 18 = 36 g

1 mol H_2O = 18 g H_2O

$$\Rightarrow \frac{18\text{ g } H_2O}{1\text{ mol } H_2O} = 18\text{g/mol}$$

$$\text{Hence, } 2\text{ mol } H_2O \times \frac{18\text{ g } H_2O}{1\text{ mol } H_2O}$$

$$= 2 \times 18\text{ g } H_2O = \mathbf{36\ g\ H_2O}$$

Example 55

How many moles of methane are required to produce 22 g $CO_2(g)$ after combustion?

Solution

According to the chemical equation,

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

44 g $CO_2(g)$ is obtained from 16 g $CH_4(g)$.

[∵ 1 mol $CO_2(g)$ is obtained from 1 mol of $CH_4(g)$]

mole of $CO_2(g)$

$$= 22\text{ g } CO_2(g) \times \frac{1\text{ mol } CO_2(g)}{44\text{ g } CO_2(g)}$$

$$= 0.5\text{ mol } CO_2(g)$$

Hence, 0.5 mol $CO_2(g)$ would be obtained from 0.5 mol $CH_4(g)$ or 0.5 mol of $CH_4(g)$ would be required to produce 22 g $CO_2(g)$.

EXERCISE 7.8

1. Balance the following chemical reactions:

(*a*) $CO + O_2 \rightarrow CO_2$

(*b*) $KNO_3 \rightarrow KNO_2 + O_2$

(*c*) $O_3 \rightarrow O_2$

(*d*) $NH_4NO_3 \rightarrow N_2O + H_2O$

(*e*) $CH_3NH_2 + O_2 \rightarrow CO_2 + H_2O + N_2$

(*f*) $Cr(OH)_3 + HClO_4 \rightarrow Cr(ClO_4)_3 + H_2O$

2. Write the balanced chemical equations of each reaction:

(*a*) Calcium carbide (CaC_2) reacts with water to form calcium hydroxide ($Ca(OH)_2$) and acetylene gas (C_2H_2).

(*b*) When potassium chlorate ($KClO_3$) is heated, it decomposes to form KCl and oxygen gas (O_2).

(*c*) C_6H_6 combusts in air.

(*d*) $C_5H_{12}O$ combusts in air.

7.11 LIMITING REACTANTS

ACTIVITY 7.4: Illustrating Limiting Reactants

Understanding Limiting Reagent

Imagine you are packing biscuits. In each biscuit packet, you put 10 biscuits.

For packing, you need 1 empty packet and 10 loose biscuits. This means

1 empty packet + 10 loose biscuits $\rightarrow$ 1 biscuit packet

Case 1: If you have 30 empty packets and 250 loose biscuits, how many packets of biscuits can you make? How many loose biscuits or empty packets will remain?

Case 2: If you have 15 empty packets and 200 loose biscuits, how many packets of biscuits can you make? How many loose biscuits or empty packet will remain?

In **Case 1**, you can make 25 packets of biscuits and 5 empty packets will be left over. Here, you do not have more biscuits to fill in the empty packets. So, biscuit is the limiting reagent.

In **Case 2**, you can make 15 packets of biscuits and 50 loose biscuits will be left over. Here, you do not have more empty packets to make biscuit packets. So, empty packet is the limiting reagent.

Many times, the reactions are carried out when the reactants are not present in the amounts as required by a balanced chemical reaction. In such situations, one reactant is in excess over the other. The reactant which is present in the lesser amount gets consumed after sometime and after that no further reaction takes place whatever be the amount of the other reactant present. Hence, the reactant which gets consumed, limits the amount of product formed and is therefore called the **limiting reagent.**

In performing stoichiometric calculations, this aspect is also to be kept in mind.

Example 56

How much magnesium sulphide can be obtained from 2.00 g of magnesium and 2.00 g of sulphur by the reaction $Mg + S \longrightarrow MgS$? Which is the limiting reagent? Calculate the amount of one of the reactants which remains unreacted.

Solution

First of all each of the masses are expressed in moles:

$$2.00 \text{ g of Mg} = \frac{2.00}{24.3}$$
$$= 0.0824 \text{ moles of Mg}$$

$$2.00 \text{ g of S} = \frac{2.00}{32.1}$$
$$= 0.0624 \text{ moles of S}$$

From the equation, $Mg + S \longrightarrow MgS$, it follows that one mole of Mg reacts with one mole of S. We are given more moles of Mg than of S. Therefore, Mg is in excess and some of it will remain unreacted when the reaction is over. **S is the limiting reagent** and will control the amount of product. From the equation we note that one mole of S gives one mole of MgS, so 0.0624 mole of S will react with 0.0624 mole of Mg to form 0.0624 mole of MgS.

Molar mass of MgS = 56.4 g

$\therefore$ Mass of MgS formed

$$= 0.0624 \times 56.4 \text{ g}$$
$$= \mathbf{3.52 \text{ g of MgS}}$$

Mole of Mg left unreacted

$= 0.0824 - 0.0624$ moles of Mg

$= 0.0200$ moles of Mg

Mass of Mg left unreacted

= moles of Mg × molar mass of Mg

$= 0.0200 \times 24.3$ g of Mg

= 0.486 g of Mg

Example 57

50.0 kg of $N_2(g)$ and 10.0 kg of $H_2(g)$ are mixed to produce $NH_3(g)$. Calculate the $NH_3(g)$ formed. Identify the limiting reagent in the production of NH_3 in this situation.

Solution

A balanced equation for the above reaction is written as follows:

Calculation of moles:

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

moles of N_2

$$= 50.0 \text{ kg } N_2 \times \frac{1000 \text{ g } N_2}{1 \text{ kg } N_2} \times \frac{1 \text{ mol } N_2}{28.0 \text{ g } N_2}$$

$= 17.86 \times 10^2$ mol

moles of H_2

$$= 10.00 \text{ kg } H_2 \times \frac{1000 \text{ g } H_2}{1 \text{ kg } H_2} \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2}$$

$= 4.96 \times 10^3$ mol

According to the above equation, 1 mol $N_2(g)$ requires 3 mol $H_2(g)$, for the reaction. Hence, for 17.86×10^2 mol of N_2, the moles of $H_2(g)$ required would be

$$17.86 \times 10^2 \text{ mol } N_2 \times \frac{3 \text{ mol } H_2(g)}{1 \text{ mol } N_2(g)}$$

$= 5.36 \times 10^3$ mol H_2

But we have only 4.96×10^3 mol H_2. Hence, hydrogen is the limiting reagent in this case. So $NH_3(g)$ would be formed only from that amount of available hydrogen, *i.e.*, 4.96×10^3 mol

Since 3 mol $H_2(g)$ gives 2 mol $NH_3(g)$

$$4.96 \times 10^3 \text{ mol } H_2(g) \times \frac{2 \text{ mol } NH_3(g)}{3 \text{ mol } H_2(g)}$$

$= 3.30 \times 10^3$ mol $NH_3(g)$

3.30×10^3 mol $NH_3(g)$ is obtained.

If they are to be converted to grams, it is done as follows:

1 mol $NH_3(g)$ = 17.0 g $NH_3(g)$

$$3.30 \times 10^3 \text{ mol } NH_3(g) \times \frac{17.0 \text{ g } NH_3(g)}{1 \text{ mol } NH_3(g)}$$

$= 3.30 \times 10^3 \times 17$ g $NH_3(g)$

$= 56.1 \times 10^3$ g NH_3

= 56.1 kg NH_3

Example 58

Copper reacts with silver nitrate solution according to the equation

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

If 0.50 mol of copper is added to 1.5 mol of silver nitrate, which is the limiting reagent and how many moles of silver are formed?

Solution

Decide which is the limiting reagent.

According to the equation, 1 mol Cu = 2 mol $AgNO_3$, so

0.50 mol Cu = 2×0.50 = 1.0 mol $AgNO_3$

but there are 0.50 mol Cu and 1.5 mol $AgNO_3$. Therefore, $AgNO_3$ is present in excess and Cu is the **limiting reagent**.

Calculate how many moles of silver are formed.

Use the amount of the limiting reagent to find the amount of product. According to the equation, 1 mol Cu = 2 mol Ag. Therefore, 0.50 mol Cu ≡ 2×0.50 = 1.0 mol Ag.

So, 1.0 mol Ag is formed.

EXERCISE 7.9

1. Identify the limiting reactant in the reaction of hydrogen and oxygen to form water if 61.0 g of O_2 and 8.40 g of H_2 are combined. Determine the amount (in grams) of excess reactant that remains after the reaction is complete.
2. Calculate the limiting reactant.
   Zinc metal reacts with hydrochloric acid by the following reaction:
   $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
   If 0.30 mol Zn is added to hydrochloric acid containing 0.52 mol HCl, how many moles of H_2 are produced?

7.12 THE GASEOUS STATE

This is the simplest state of matter. Throughout our life, we remain immersed in the ocean of air which is a mixture of gases. We spend our life in the lowermost layer of the atmosphere called troposphere, which is held to the surface of the earth by gravitational force. The thin layer of atmosphere is vital to our life. It shields us from harmful radiations and contains substances like oxygen, nitrogen, carbon dioxide, water vapour, etc.

Let us now focus our attention on the behaviour of substances which exist in the gaseous state under normal conditions of temperature and pressure. A look at the periodic table shows that only eleven elements exist as gases under normal conditions (Figure 7.5).

Group Number	1	15	16	17	18
	H				He
		N	O	F	Ne
				Cl	Ar
					Kr
					Xe
					Rn

Figure 7.5: Eleven elements that exist as gases.

The gaseous state is characterised by the following physical properties.

- Gases are highly compressible.
- Gases exert pressure equally in all directions.
- Gases have much lower density than the solids and liquids.
- The volume and the shape of gases are not fixed. These assume volume and shape of the container.
- Gases mix evenly and completely in all proportions without any mechanical aid.

Simplicity of gases is due to the fact that the forces of interaction between their molecules are negligible. Their behaviour is governed by same general laws, which were discovered as a result of their experimental studies. These laws are relationships between measurable properties of gases. Some of these properties like pressure, volume, temperature and mass are very important because relationships between these variables describe state of the gas. Interdependence of these variables leads to the formulation of gas laws.

7.13 THE GAS LAWS

The gas laws which we will study now are the result of research carried on for several centuries on the physical properties of gases. The first reliable measurement on properties of gases was made by Anglo-Irish scientist Robert Boyle in 1662. The law which he formulated is known as Boyle's Law. Later on, attempts to fly in air with the help of hot air balloons motivated Jacques Charles and Joseph Lewis Gay Lussac to discover additional gas laws. Contribution from Avogadro and others provided lot of information about gaseous state.

STANDARD TEMPERATURE AND PRESSURE (S.T.P.)

Since volume of a given mass of a gas depends on both temperature and pressure, it is necessary to specify the values of *p* and T when the value of V is stated.

In general, the comparison of the volumes of different gases is made with reference to standard temperature and pressure.

The standard temperature is taken as **0°C (or 273.15 K)**. The standard pressure according to the latest recommendations is taken as **1 bar (or 10^5 pascal).** The molar volume of ideal gas at S.T.P. conditions is **22.71098 L** mol^{-1} (≈ **22.7 L mol^{-1}**).

These conditions are quite often abbreviated as **S.T.P.** meaning *standard temperature and pressure.*

It is worth noting that previous S.T.P. conditions (**0°C or 273.15 K** temperature and **1 atm (= 1.01325 bar)** pressure are still used in many books quite often. At these conditions, the molar volume of ideal gas is 22.413996 L mol^{-1} (≈ 22.4 L mol^{-1}).

It may be out of place to mention here, that **Standard Ambient Temperature and Pressure (SATP)** conditions are also used in some scientific data. The **SATP** conditions are **298.15 K (25°C)** and **1 bar (10^5 Pa)** pressure the molar volume of ideal gas at SATP conditions is **24.789 L mol^{-1}**. At RTP (room temperature and pressure) with conditions of 1atm and 250C previous value of volume is 24L and it is still used.

7.13.1 Gay Lussac Law

ACTIVITY 7.5: Illustrating Gay Lussac Law

- Take a balloon.
- Fill the balloon with air.
- Put the air filled balloon in a hot summer day outside your classroom.
- Observe the balloon.

In Activity 7.5, you will observe that after some time the inflated balloon will burst. The balloon bursts because the pressure inside the balloon increases due to high temperature.

The pressure-temperature relationship.

Gay-Lussac's Law states that the pressure of a given mass of gas is directly proportional to the Kelvin temperature of that gas, when the volume is kept constant.

When the temperature of a sample of a gas in a non-expandable container increases, the pressure **of the gas increases too.**

Explanation of Gay-Lussac's law by kinetic theory

As the kinetic energy increases the molecules of the gas hit the walls of the container with more force, resulting in a greater pressure. This law shows

The mathematical expressions for Gay-Lussac's Law are given below:

P/T and P_1/T_1 =P_2/T_2

Worked example

The gas in an aerosol can is under a pressure of 3atm at a temperature of 25^0C. What would be the pressure in the aerosol can if the temperature rises to 845^0C?

Solution:

Known

P1=3atm

T1=25 ^{0}C=298K

T2=845 ^{0}C=1118K

Unknown

P2=?atm

$$P_2 = \frac{P_1T_2}{T_1} = \frac{3atmx1118K}{298K} = 11.3atm$$

7.13.2 Charles' Law

ACTIVITY 7.6: Illustrating Charles' Law

Step 1

- **Step 1:** Fill a balloon with air.
- **Step 2:** Fit the balloon between two steady objects such as stacks of books so that it just barely touches the objects on each side. Keep the objects exactly where they are. **Don't move the objects**.
- **Step 3:** Place the balloon in a freezer or ice box for thirty minutes to cool it.
- **Step 4:** Remove the balloon from the freezer or ice box.
- **Step 5:** Replace the balloon between the two objects of Step 2. **Don't move the objects.** What did you observe?

In this activity, you will observe that there should be extra space between the balloon and the two objects. The volume of the balloon has decreased because the air in the balloon is colder than it was initially.

Step 2

Step 5

Volume-temperature relationship

A French chemist **J .A .C Charles** made certain observations on the effect of temperature on the volumes of gases. He established the fact that a sample of any gas at 0°C would expand by $\frac{1}{273}$ of its volume for each degree it was heated, and would contract by $\frac{1}{273}$ of its volume for each degree it was cooled. This was true for all gases.

Based on Charles' results **Lord Kelvin**, a Scottish physicist, reasoned that if gases lost $\frac{1}{273}$ of their volume at 0°C for every degree they were cooled, there must be a temperature of – 273°C where gases had no volume at all.

He called – 273°C **absolute zero** at which molecules would have lost all their energy. As a result of this, a new scale of temperature, now known as **Kelvin scale** or **absolute scale** was developed. In this scale, 0 K corresponds to –273°C. Centigrade temperature is converted into Kelvin scale by adding 273 to the temperature in centigrade.

Thus, $T = t + 273$

where T = Temperature in Kelvin scale or the absolute scale

t = Temperature in centigrade scale

The relationship of volume and temperature of a gas is then expressed in absolute scale. This is known as **Charles' law**, which states:

Pressure remaining constant, the volume of a given mass of gas is directly proportional to its absolute temperature.

All gases obey Charles' law at very low pressures and high temperatures.

Let the volume of a given mass of gas be V at the absolute temperature T, then

$$V \propto T \text{ (pressure remaining constant)}$$

or, $\frac{V}{T} = k$ (a constant)

The value of k depends on the mass and the nature of gas. Now, let the volumes of a given gas be V_1 and V_2 at the absolute temperatures T_1 and T_2 respectively, pressure remaining constant, then

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Thus, using the absolute scale, we can now calculate the change in volume of a sample of a gas for any change in temperature. If the temperature is lowered, the volume will decrease; if the temperature is raised, the volume will increase.

Explanation of Charles' Law by Kinetic Theory

On heating the gas, the kinetic energy of molecules increases. This means the molecules will move faster. Hence, the gas will expand provided pressure remains constant.

Example 59

What will be the volume of a gas at 0°C which occupies 200 mL at 27°C? (assume no change in pressure).

Solution

First, change the temperature to absolute scale.

$$0°C = (0 + 273)\ K = 273\ K\ (\textit{i.e.}, T_2)$$

$$27°C = (27 + 273)\ K = 300\ K\ (\textit{i.e.}, T_1)$$

$V_1 = 200$ ml, $V_2 = ?$

By Charles' law,

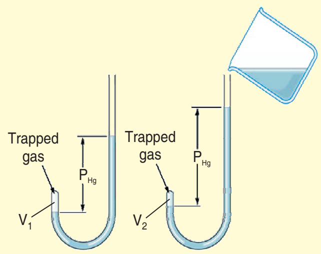
$$V_2 = \frac{V_1 \times T_2}{T_1}$$

$$V_2 = \frac{200 \times 273}{300} \text{ ml} = 182 \text{ ml}$$

7.13.3 Boyle's Law

ACTIVITY 7.7: Illustrating Boyle's Law

Take a J shape tube as shown in Figure. Tube is partially filled with mercury. Pressure is now increased by putting more mercury into the open limb. The volume of air enclosed in the space above mercury in shorter limb is noted each time. It is found that as pressure increases, the volume of enclosed air gradually decreases from V_1 to V_2.



Volume-pressure Relationship

In 1660, Robert Boyle formulated a quantitative relationship between pressure and volume of a given mass of air at constant temperature. Boyle's law states:

The temperature remaining constant, the volume of a given mass of gas is inversely proportional to the pressure applied to it.

This means, if the original pressure is increased five times, the original volume is reduced to its one-fifth.

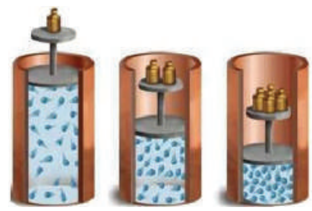


Figure 7.6: Increase in pressure brings the molecules close together.

The figure given above shows a sample of gas enclosed in a cylinder fitted with a moving piston. The pressure is increased by increasing the weight on the piston. With increasing pressure, the volume is reduced. When the weight is double, the new volume is one-half the original volume, and so on.

Mathematical Derivation of the Relationship

Let the given mass of gas occupy volume V at pressure P.

Then, according to Boyle's law,

$$P \propto \frac{1}{V}$$

(temperature remaining constant)

or, $PV = k$ (a constant)

The value of k depends on the nature and mass of gas.

Now, let the volumes occupied by the same amount of gas be V_1 and V_2 at P_1 and P_2 respectively, then at constant temperature

$$P_1V_1 = k$$

$$P_2V_2 = k$$

or, $P_1V_1 = P_2V_2 = k$

or, $P_1V_1 = P_2V_2$

Boyle's law is also manifested in the working of many devices that we use in daily life such as tyre pressure gauge, aneroid barometer and cycle pump, etc.

Boyle's law can also help to deduce relationship between ***density and pressure*** of the gas

$$\text{Density } (d) = \frac{\text{Mass}}{\text{Volume}} = \frac{M}{V}$$

Put the value of V as k_1/p from Boyle's law equation

$$\mathbf{d} = \left(\frac{\mathbf{m}}{\mathbf{k_1}}\right)\mathbf{p}$$

which shows that pressure of fixed mass of a gas is directly proportional to density.

Explanation of Boyle's Law by Kinetic Theory

- According to the kinetic theory of gases, the pressure exerted by a gas results from the combined bombardment of its molecules on the walls of the container.
- The number of bombardments depends on the concentration of molecules at a certain temperature.
- When the volume is reduced to one-fifth, the concentration of molecules is increased five times and hence the pressure is increased five times.
- Thus, the pressure of the gas becomes inversely proportional to the volume of a gas at constant temperature.

Under ordinary conditions, Boyle's law is entirely satisfactory. But at very low temperatures or under high pressures, deviations may occur. This is because the actual molecules are taking up a large part of the total volume of the gas at high pressures or at very low temperatures. Correction factors have been suggested to Boyle's law. A gas which obeys Boyle's law is said to show ideal behaviour.

Example 60

A certain mass of a gas occupies 48 ml, at a pressure of 720 mmHg. What is the volume when the pressure is increased to 960 mmHg? (temperature remains constant).

Solution

$$V_1 = 48 \text{ ml}, V_2 = ?$$
$$P_1 = 720 \text{ mm}, P_2 = 960 \text{ mmHg}$$

By Boyle's law,

$$P_1V_1 = P_2V_2$$
$$V_2 = \frac{V_1P_1}{P_2} = \frac{48 \times 720}{960} = \mathbf{36\ ml}$$

Example 61

A gas occupies 1200 litres at 2 atm pressure. To what pressure must it be compressed to occupy 60 litres at the same temperature?

Solution

$$V_1 = 1200 \text{ L}, V_2 = 60 \text{ L}$$
$$P_1 = 2 \text{ atm}, P_2 = ?$$

By Boyle's law,

$$P_1V_1 = P_2V_2$$
$$P_2 = \frac{V_1P_1}{V_2} = \frac{1200 \times 2}{60} = \mathbf{40\ atm}$$

The Gas Equation

Boyle's law and Charles' law can be combined to form one single relation. The equation is called combined gas law or general gas equation.

It can be deduced as follows:

$V \propto \frac{1}{P}$ at constant temperature (Boyle's law)

$V \propto T$ at constant pressure (Charles' law)

Then, according to mathematical law,

$$V \propto \frac{T}{P}$$

$$V \propto k\frac{T}{P} \quad (k \text{ is gas constant})$$

Because $\frac{PV}{T} = k$, to obtain equation relating the volume, pressure and temperature of a given sample of gas at one set of conditions to those under any other set of conditions, we may simply write

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{P_3V_3}{T_3}$$

where P_1, V_1 and T_1; P_2, V_2 and T_2; P_3, V_3 and T_3 respectively represent the pressure, volume and temperature values under given set of conditions.

Example 62

A gas measures 80 ml at 2.5 atm pressure and a temperature of 27°C. Calculate its volume at standard conditions of temperature and pressure.

Solution

$P_1 = 2.5$ atm, $\quad P_2 = 1$ atm,

$V_1 = 80$ ml, $\quad V_1 = ?$

$T_1 = (27 + 273) = 300$ K

$T_2 = 273$ K

From the gas equation,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

Substituting the values,

$$V_2 = \frac{2.5 \times 80 \times 273}{300 \times 1} = \mathbf{182\ ml}$$

Example 63

The volume of a gas is 150 ml. at 17°C and 700 mm of Hg. What will be its volume at standard temperature pressure?

Solution

$P_1 = 700$ mm of Hg,

$P_2 = 760$ mm of Hg

$V_1 = 150$ ml, initial, $\quad V_2 = ?$

$T_1 = (17 + 273) = 290$ K, $\quad T_2 = 273$K

From the gas equation,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{700 \times 150}{290} = \frac{760 \times V_2}{273}$$

$$V_2 = \frac{700 \times 150 \times 273}{290 \times 760} = \mathbf{130.1\ ml}$$

7.13.4 Avogadro's Law—Volume Amount Relationship

This law describes the ***volume-amount*** relationship of gases at constant temperature and pressure. It was given by Amedeo Avogadro in 1811. It states that ***equal volumes of all the gases under similar conditions of temperature and pressure contain equal number of molecules.***

For example, 1 mol of all the gases contains 6.023×10^{23} molecules. At the same time 1 mol of all the gases at 273.15 K (0°C) and 1 bar pressure occupy a volume of 22.7 *l* (22.7×10^{-3} m^3). This means that as long as the temperature and pressure remain constant, the volume of the gas is directly proportional to the number of molecules. In other words, the amount of the gas molecules is constant. Mathematically, we can write

$$V \propto N \qquad (\textit{T and p are constant})$$

The number of molecules N is directly proportional to number of moles (n)

$\therefore \quad V \propto n \quad$ (*T and p are constant*)

or $\quad V = k_4 n$

k_4 is constant of proportionality

Now, if m is the mass of the gas having molar mass equal to M, then the number of moles (n) are given as

$$n = \frac{m}{M}$$

$$\therefore \quad V = k_4 \frac{m}{M}$$

$$\text{or} \quad M = k_4 \frac{m}{V} \quad \text{or} \quad k_4.d$$

where d is density of gas.

The above relationship implies that density of the gas at a given temperature and pressure is directly proportional to its molar mass.

7.13.5 Ideal Gas Equation

A gas that follows Boyle's law, Charles' law and Avogadro's law strictly at all conditions, is called **Ideal gas.** It is assumed that inter-molecular forces are not present between the molecules of ideal gas.

The combination of various gas laws namely; *Boyle's law, Charles' law* and *Avogadro's law* leads to the development of the mathematical relation which relates four variables *pressure, volume, absolute temperature* and *number of moles of ideal gas*. The equation so formulated is called **ideal gas equation.** **PV=nRT** where P is pressure, V is volume, n is number of moles, R is the gas constant and T is temperature

Let us solve some numerical problems based on gas laws.

> **Note:** In many of these problems, the conversion of temperature in celsius scale to kelvin scale has been done by adding 273 instead of 273.15 in order to make the calculations simple.

Example 64

A sample of a gas is found to occupy a volume of 900 cm³ at 27°C. Calculate the temperature at which it will occupy a volume of 300 cm³, provided the pressure is kept constant.

Solution

Here, $\quad V_1 = 900 \text{ cm}^3 \quad V_2 = 300 \text{ cm}^3$

$T_1 = (27 + 273) \text{ K} = 300 \text{ K}, \quad T_2 = ?$

Applying Charle's law, $\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$

$$T_2 = \frac{V_2 \times T_1}{V_1} = \frac{300 \text{ cm}^3 \times 300 \text{ K}}{900 \text{ cm}^3}$$

$$= 100 \text{ K} = 100 - 273 = \mathbf{-173°C}$$

Example 65

It is desired to increase the volume of 80 cm³ of a gas by 20% without changing the pressure. To what temperature the gas be heated if its initial temperature is 25°C?

Solution

The desired increase in the volume of gas

$$= 20\% \text{ of } 80 \text{ cm}^3$$

$$= \frac{80}{100} \times 20 = 16 \text{ cm}^3$$

Thus, the final volume of the gas

$$= 80 + 16 = 96 \text{ cm}^3$$

Now, $\quad V_1 = 80 \text{ cm}^3$

$V_2 = 96 \text{ cm}^3$

$T_1 = 25°C = 298 \text{ K} \qquad T_2 = ?$

Applying Charle's law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$$\therefore \qquad T_2 = \frac{V_2 T_1}{V_1} = \frac{96\ cm^3 \times 298\ K}{80\ cm^3}$$

$$= 357.6\ K = 357.6 - 273$$

$$= \mathbf{84.6°C}$$

Example 66

An iron tank contains helium at a pressure of 2.5 atmospheres at 25°C. The tank can withstand a maximum pressure of 10 atmospheres. The building in which tank has been placed catches fire. Predict whether, the tank will blow up first or melt. (The melting point of iron = 1535°C).

Solution

Let us proceed to calculate the pressure build up in the tank at the melting point of iron.

Thus, $p_1 = 2.5$ atm. $\qquad p_2 = ?$

$T_1 = 25°C = 298\ K$

$T_2 = 1535°C = 1808\ K$

According to Charle's law equation,

$$\therefore \qquad \frac{p_1}{T_1} = \frac{p_2}{T_2}$$

$$\text{Thus,} \qquad p_2 = \frac{p_1 T_2}{T_1}$$

$$= \frac{2.5 \times 1806}{298} = \mathbf{15.15\ atm}$$

Since, pressure of the gas in the tank is much more than 10 atm at the melting point. Thus, the tank will **blow up before reaching the melting point.**

Example 67

When a ship is sailing in Pacific ocean where temperature is 23.4°C, a balloon is filled with 2.0 l of air. What will be the volume of the balloon when the ship reaches Indian ocean, where temperature is 26.1°C.

Solution

$T_1 = 23.4°C = 23.4 + 273.15 = 296.55\ K$

$T_2 = 26.1°C = 26.1 + 273.15 = 299.25\ K$

$V_1 = 2.0\ l$; $V_2 = ?$

Applying Charles' law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or}$$

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{2.0\ (l) \times 299.95\ (K)}{296.55\ (K)}$$

$$= \mathbf{2.022\ l}$$

Example 68

In a J tube partially filled with mercury the volume of air column is 4.2 ml and the mercury level in the two limbs is same. Some mercury is, now added to the tube so that the volume of air enclosed in shorter limb is now 2.8 ml. What is the difference in the levels of mercury in this situation. Atmospheric pressure is reported to be 1.0 bar.

Solution

Initial pressure $p_1 = 1$ bar $= 0.987$ atm

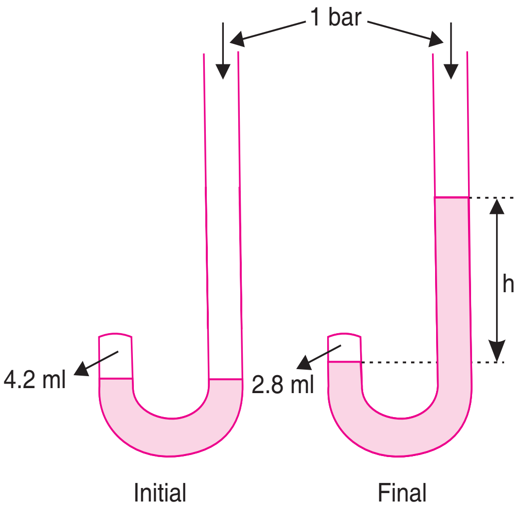
$= 0.987 \times 760 = 750.12$ mm Hg

Final Pressure $p_2 = (p_1 + h)$

$V_1 = 4.2$ ml; $V_2 = 2.8$ ml

$$p_2 V_2 = p_1 V_1 \quad \text{or} \quad p_2 = \frac{p_1 V_1}{V_2}$$

$$= \frac{750.12 \times 4.2}{2.8} = 1125.18 \text{ mm Hg}$$



Now, $1125.18 = 750.12 + h$

or $h = 1125.18 - 750.12$

$= 375.06$ mm Hg

or **37.50 cm Hg**

Example 69

A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 0.27 l Volume; up to what volume can the balloon be expanded by filling H_2?

Solution

Here, $p_1 = 1$ bar, $p_2 = 0.2$ bar

$V_1 = 0.27\ l$, $V_2 = ?$

From Boyle's law equation:

$$p_1V_1 = p_2V_2$$

$$\therefore \quad V_2 = \frac{p_1V_1}{p_2} = \frac{1 \times 0.27}{0.2} = 1.35\ l$$

Since the balloon bursts at 0.2 bar pressure. Hence, the volume of balloon should remain less than **1.35 *l***

Example 70

A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 2.27 l volume, up to what volume can the balloon be expanded?

Solution

According to Boyle's law, $p_1V_1 = p_2V_2$

If p_1 is 1 bar, V_1 will be 2.27 *l*

If $p_2 = 0.2$ bar, then $V_2 = \dfrac{p_1V_1}{p_2}$

$$\Rightarrow \quad V_2 = \frac{1 \text{ bar} \times 2.27\ l}{0.2 \text{ bar}} = 11.35\ l$$

Since balloon bursts at 0.2 bar pressure, the volume of balloon should be less than 11.35 *l*.

Type I. Solved Problems Based on Combined Gas Law: $\left(\dfrac{\mathbf{p_1V_1}}{\mathbf{T_1}} = \dfrac{\mathbf{p_2V_2}}{\mathbf{T_2}}\right)$

Example 71

A sample of nitrogen gas occupies a volume of 320 cm³ at S.T.P. Calculate its volume at 66°C and 0.825 bar pressure.

Solution

Here, $p_1 = 1.00$ bar $\quad p_2 = 0.825$ bar

$V_1 = 320\ cm^3 \quad V_2 = ?$

$T_1 = 273.15$ K

$T_2 = 66°C = (66 + 273.15)$ K

$= 339.15$ K

According to the gas equation,

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

$$\therefore \quad V_2 = \frac{p_1 V_1 T_2}{T_1 p_2} = \frac{1 \times 320 \times 339.15}{273.15 \times 0.825}$$

$$= \mathbf{482.09\ cm^3}$$

Example 72

1.0 mol of pure dinitrogen gas at SATP conditions was put into a vessel of volume 0.025 m³, maintained at the temperature of 50°C. What is the pressure of the gas in the vessel?

Solution

Volume of 1.0 mol of a gas at SATP (V_1)

$$= 24.8 \times 10^{-3}\ m^3$$

Initial condition (SATP)	*Final condition*
$V_1 = 24.8 \times 10^{-3}\ m^3$	$V_2 = 0.025\ m^3$
$p_1 = 1$ bar	$p_2 = ?$
$T_1 = 298.15$ K	$T_2 = 50°C$ $= 323.15$ K

According to general gas equation,

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

$$\text{or} \quad p_2 = \frac{p_1 V_1 T_2}{T_1 V_2}$$

$$= \frac{1.0 \times 24.8 \times 10^{-3} \times 323.15}{298.15 \times 0.025}$$

$$= \mathbf{1.075\ bar}$$

Example 73

At 25°C and 760 mm of Hg pressure, a gas occupies 600 ml volume. What will be its pressure at a height where temperature is 10°C and volume of the gas is 640 ml.

Solution

$p_1 = 760$ mm Hg, $V_1 = 600$ ml

$T_1 = 25 + 273 = 298$ K

$V_2 = 640$ ml

and $T_2 = 10 + 273 = 283$ K

According to Combined gas law,

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

$$\Rightarrow \quad p_2 = \frac{p_1 V_1 T_2}{T_1 V_2}$$

$$\Rightarrow p_2 = \frac{(760\ \text{mm Hg}) \times (600\ \text{ml}) \times (283\ \text{K})}{(640\ \text{ml}) \times (298\ \text{K})}$$

$$= \mathbf{676.6\ mm\ Hg}$$

Type II. Solved Problems Involving Molar Mass and Density $\mathbf{p = \frac{dRT}{M}}$

Example 74

The density of certain gaseous oxide at 1.5 bar pressure at 10°C is same as that of dioxygen at 20°C and 4.5 bar pressure. Calculate the molar mass of gaseous oxide.

Solution

Density of dioxygen (O_2) at 4.5 bar pressure and 10°C

$$d_{O_2} = \frac{pM}{RT} = \frac{4.5\ (\text{bar}) \times 32\ (\text{g mol}^{-1})}{R \times 283.15\ (\text{K})}$$

$$d_{\text{oxide}} = \frac{pM}{RT} = \frac{1.5\ (\text{bar}) \times M\ (\text{g mol}^{-1})}{R \times 293.15\ (\text{K})}$$

Now, $d_{O_2} = d_{\text{oxide}}$

$$\therefore \quad \frac{4.5\ (\text{bar}) \times 32\ (\text{g mol}^{-1})}{R \times 283.15\ (\text{K})}$$

$$= \frac{1.5\ (\text{bar}) \times M\ (\text{g mol}^{-1})}{R \times 293.15}$$

$$\text{or} \quad M = \frac{4.5 \times 293.15 \times 32}{283.15 \times 1.5} = \mathbf{99.39\ g\ mol^{-1}}$$

Type III. Solved Problems Involving Mass-Volume Relationship between Reactants and Products

Example 75

Isobutane (C_4H_{10}) undergoes combustion in oxygen according to the reaction:

$$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10\,H_2O(l)$$

When 10.00 L of isobutane is burnt at 27°C and 1 bar pressure, what volume of CO_2 is produced at 80°C and 1.5 bar pressure.

Solution

Step I. *Calculation of volume of CO_2 at 27°C and 1 bar.*

From the chemical equation it is clear that

2 *l* of C_4H_{10} produce CO_2 = 8 *l*

(Similar condition of T and P)

$$\therefore\ 10\ l \text{ of } C_4H_{10} \text{ produce } CO_2 = \frac{8}{2} \times 10 = 40\ l \text{ (at 27°C, 1 bar)}$$

Step II. *Conversion of volume of CO_2 at 80°C and 1.5 bar.*

Here, p_1 = 1 bar $\quad p_2$ = 1.5 bar

$V_1 = 40\ l \quad V_2 = ?$

$T_1 = 300$ K $\quad T_2 = 80°C = 80 + 273 = 353$ K

Using the gas equation,

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

$$V_2 = \frac{p_1V_1T_2}{T_1p_2} = \frac{1 \times 40 \times 353}{300 \times 1.5} = \mathbf{31.37\ l}$$

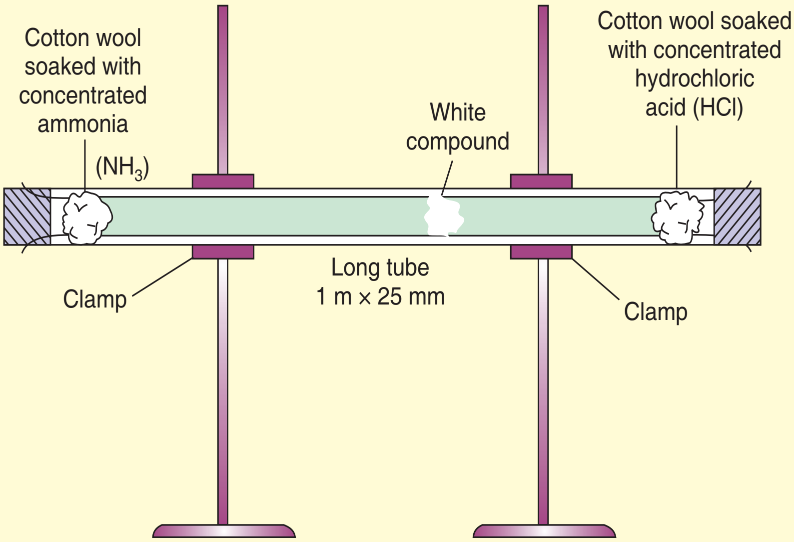
7.13.6 Graham's Law of Diffusion

ACTIVITY 7.8: To Study the Diffusion of Two Gases

1. Clamp a long glass tube about (1 m × 25 mm) dimensions in a horizontal position as shown in figure. Fix long strips of blue and red litmus papers with the help of a cello-tape. Insert cotton wool plugs at both ends.
2. With the help of a medicine dropper place 5 drops of concentrated hydrochloric acid on one cotton plug.

At the same time another student must place 5 drops of concentrated ammonia on the cotton plug at the other end. As quickly as possible start the stopwatch and cork at both ends of the tube.

3. When white ring of smoke is formed, stop the stopwatch and record the time. Mark the point where the ring appears first and measure the distance of this point from both ends.



Rates of diffusions of NH_3 and HCl.

Make the following record of observations

(*a*) Initial time (t_1) when drops are introduced ......

(*b*) Final time (t_2) when the white ring appears first ......

(*c*) Time taken for diffusion 't' = $(t_2 - t_1)$

(*d*) Distance travelled by HCl (x_1) =

(*e*) Distance travelled by NH_3 (x_2) =

(*f*) Rate of diffusion of HCl (r_1) = x_1/t

(*g*) Rate of diffusion of NH_3 (r_2) = x_2/t

(*h*) Find the ratio r_2/r_1

(*i*) Compare this ratio to another ratio

$$\sqrt{\frac{M_{HCl}}{M_{NH_3}}} \quad \text{or} \quad \sqrt{\frac{36.5}{17}}$$

(*j*) Verify Graham's Law.

Thomas Graham put forward a generalisation after studying the rates of diffusion of different gases, which is known after his name as *Graham's law of diffusion*. The law states:

Under similar conditions of temperature and pressure, the rates of diffusion of gases are inversely proportional to the square roots of their densities.

The law can mathematically be written as

$$r \propto \frac{1}{\sqrt{d}}$$

where r is the rate of diffusion and d is the density of the gas.

Now, if there are two gases A and B having r_1 and r_2 as their rates of diffusion and d_1 and d_2, and their densities respectively. Then

$$r_1 \propto \frac{1}{\sqrt{d_1}} \quad \text{and} \quad r_2 \propto \frac{1}{\sqrt{d_2}}$$

or $$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \qquad \text{(at same T and P)}$$

We know that molecular mass is twice the vapour density. Therefore, the above expression may be written as

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2/2}{M_1/2}} = \sqrt{\frac{M_2}{M_1}}$$

where M_1 and M_2 are the molecular masses of the gases having densities d_1 and d_2 respectively. Thus, Graham's law may also be stated as:

Under similar conditions of temperature and pressure, the rates of diffusion of gases are inversely proportional to the square root of their molecular masses.

Again, the rate of diffusion of the gas is equal to the volume of the gas which diffuses per unit time, *i.e.*,

Rate of diffusion

$$= \frac{\text{Volume of the gas diffused}}{\text{Time taken for diffusion}}$$

or $$r = \frac{V}{t}$$

If V_1 and V_2 are the volumes of the gases diffusing in the time t_1 and t_2 respectively under similar conditions, then

$$r_1 = \frac{V_1}{t_1} \quad \text{and} \quad r_2 = \frac{V_2}{t_2}$$

Thus, putting together the values of r_1 and r_2 we arrive at the following formula:

$$\boxed{\frac{r_1}{r_2} = \frac{V_1/t_1}{V_2/t_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}}$$

Now, if $V_1 = V_2 = V$, we get

$$\frac{t_2}{t_1} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

This implies that *time taken for the diffusion of equal volumes of two gases under similar conditions of temperature and pressure is directly proportional to the square root of their densities or molecular masses.*

Similarly, if $t_1 = t_2 = t$, then

$$\frac{V_1}{V_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

It means that *volumes of the two gases which diffuse in the same time under similar conditions are inversely proportional to the square roots of their densities or molecular masses.*

Let us now apply **Graham's law** to solve some **numerical problems.**

Example 77

Compare the rates of diffusion of $^{235}UF_6$ and $^{238}UF_6$.

Solution

Let the rate of diffusion of $^{235}UF_6$ be r_1 and rate of diffusion of $^{238}UF_6$ be r_2

Now, molecular mass of $^{235}UF_6(M_1)$

$$= 235 + 6 \times 19 = 349$$

Molecular mass of $^{238}UF_6(M_2)$

$$= 238 + 6 \times 19 = 352$$

According to Graham's law $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

$$\frac{r_1}{r_2} = \sqrt{\frac{352}{349}} = 1.004.$$

Thus, **$r(^{235}UF_6) : r(^{238}UF_6)$ is 1.004 : 1**

Example 78

Relative densities of oxygen and carbon (IV) oxide are 16 and 22 respectively. If 25 cm^3 of carbon (IV) oxide diffuses out in 75 seconds, what volume of oxygen will diffuse out in 96 seconds under similar conditions?

Solution

Here, volume of carbon (IV) oxide

$$V_{CO_2} = 25 \text{ cm}^3$$

Time taken $t_{CO_2} = 75$ seconds

Let the volume of oxygen diffused

$$V_{O_2} = V \text{ cm}^3$$

Time taken $t_{O_2} = 96$ seconds

Relative density of carbon (IV) oxide,

$$d_{CO_2} = 22$$

Relative density of oxygen $d_{O_2} = 16$

Now, $\frac{r_{O_2}}{r_{CO_2}} = \sqrt{\frac{d_{CO_2}}{d_{O_2}}}$ or $\frac{V_{O_2}}{t_{O_2}} \Big/ \frac{V_{CO_2}}{t_{CO_2}} = \sqrt{\frac{22}{16}}$

or $$\frac{V_{O_2}}{t_{O_2}} \times \frac{t_{CO_2}}{V_{CO_2}} = \sqrt{\frac{22}{16}}$$

Substituting the values

$$\frac{V_{O_2} \times 75}{96 \times 25} = \sqrt{\frac{22}{16}} = 1.172$$

or $$V_{O_2} = \frac{1.172 \times 96 \times 25}{75} = 37.5 \text{ cm}^3$$

Thus, volume of oxygen diffused = **37.5 cm^3**

Example 79

A certain gaseous fluoride of phosphorus has a formula PF_x. Under similar conditions, fluorine diffuses 1.82 times faster than the gaseous phosphorus fluoride. Find the value of x and formula of the phosphorus fluoride. (*Atomic masses, P = 31.0; F = 19.0*).

Solution

According to Graham's law

$$\frac{r_{F_2}}{r(PF_x)} = \frac{1.82}{1} = \sqrt{\frac{M_{(PF_x)}}{M_{(F_2)}}}$$

$$M_{(F_2)} = 2 \times 19 = 38$$

$$\frac{1.82}{1} = \sqrt{\frac{M_{(PF_x)}}{38}}$$

or $$M_{(PF_x)} = \frac{1.82 \times 1.82 \times 38}{1}$$

$$= 125.8 \qquad \ldots(i)$$

Molecular mass of PF_x from formula

$$= 31 + 19x \qquad \ldots(ii)$$

Equate (*i*) and (*ii*)

$$31 + 19x = 125.8$$

or $$x = \frac{125.8 - 31}{19} = 4.98 \approx 5$$

$\therefore$ The formula of phosphorus fluoride

= **PF_5**

7.14 CALCULATION OF MOLAR GAS VOLUME UNDER STANDARD CONDITIONS

Example 80

Calculate number of moles in each of the following:

(*i*) *11 g of CO_2*

(*ii*) *3.01×10^{22} molecules of CO_2*

(*iii*) *1.12 litre of CO_2 at S.T.P.*

Solution

(*i*) Molecular mass of CO_2

$= (12 + 2 \times 16) = 44$ u

44 g of CO_2 = 1 mole of CO_2

$$11 \text{ g of } CO_2 = \frac{1}{44} \times 11$$

$= \mathbf{0.25\ mol}$

(*ii*) 6.02×10^{23} molecules of CO_2

= 1 mole of CO_2

3.01×10^{22} molecules of CO_2

$$= \frac{1}{6.02 \times 10^{23}} \times 3.01 \times 10^{22}$$

$= \mathbf{0.05\ mol}$

(*iii*) 22.4 litres of CO_2 at S.T.P.

= 1 mole of CO_2

1.12 litres of CO_2 at S.T.P.

$$= \frac{1}{22.4} \times 1.12 = \mathbf{0.05\ mol}$$

Example 81

Find out volume of the following at S.T.P.:

(*i*) *14 g of nitrogen*

(*ii*) *6.022×10^{22} molecules of ammonia (NH_3)*

(*iii*) *0.1 mole of sulphur dioxide (SO_2).*

Solution

(*i*) Molar mass of N_2

$= 14 \times 2 = 28$ g mol^{-1}

28 g of N_2 occupy at S.T.P.

= 22.4 litres

$\therefore$ 14 g of N_2 occupy at S.T.P.

$$= \frac{22.4}{28} \times 14 = \mathbf{11.2}\ \boldsymbol{l}$$

(*ii*) 6.022×10^{23} molecules of NH_3 occupy at S.T.P.

= 22.4 litres

$\therefore$ 6.022×10^{22} molecules of NH_3 occupy at S.T.P.

$$= \frac{22.4}{6.022 \times 10^{23}} \times 6.022 \times 10^{22}$$

$= \mathbf{2.24}\ \boldsymbol{l}$

(*iii*) 1 mole of SO_2 occupies at S.T.P.

= 22.4 litres

$\therefore$ 0.1 mol of SO_2 occupies at S.T.P.

$$= \frac{22.4}{1} \times 0.1 = \mathbf{2.24}\ \boldsymbol{l}$$

EXERCISE 7.10

1. The relationship of volume and temperature of a gas is expressed in absolute scale. This is known as ________ .
2. Boyle's law can help to deduce relationship between ________ and ________ of the gas.
3. What do you mean by ideal gas ?
4. The rates of diffusion of gases are inversely proportional to the square root of their densities; under similar conditions temperature and pressure. (True or False)
5. Under similar conditions of temperature and pressure, the rate of diffusion of gases are inversely proportional to the square root of their molecular masses. (True or False)

7.15 SUMMARY

- **Avogadro's Number** is 6.022×10^{23}.
- The ratio of the average **mass** per **atom** of the naturally occurring form of an element to one-twelfth the **mass** of an **atom** of carbon-12 is known as **relative atomic mass**.
- The ratio of the average **mass** of one **molecule** of an element or compound to one twelfth of the **mass** of an atom of carbon-12 is known as **relative molecular mass**.
- **Formula mass** is the sum of the atomic masses of atoms in a molecule.
- **Molar mass** is the mass of one mole of a substance, usually expressed in grams. The molar mass of a substance in grams is numerically equal to the atomic, molecular, or formula mass of the substance expressed in amu. Molar mass and Avogadro's number can be used to interconvert among mass, moles, and number of particles (atoms, molecules, ions, formula units, etc.).
- The amount of substance present in a given volume of a solution is expressed in mol per volume.
- The **empirical formula** can be used to calculate per cent composition. The empirical formula and molar mass can be used to determine the molecular formula.
- A chemical **formula** that gives the total number of atoms of each element in each **molecule** of a substance is called molecular formula.
- The quantitative study of the reactants required or the products formed is called stoichiometry. Using stoichiometric calculations, the amounts of one or more reactant(s) required to produce a particular amount of product can be determined and vice versa.
- The **limiting reactant** is the reactant that is consumed completely in a chemical reaction. An excess reactant is the reactant that is not consumed completely. The maximum amount of product that can form depends on the amount of limiting reactant.
- **Gay-Lussac's Law** states that the pressure of a given mass of gas is directly proportional to the Kelvin temperature of that gas, when the volume is kept constant.
- **Charles' law** is a relationship between volume and absolute temperature under isobaric condition. It states that volume of a fixed amount of gas is directly proportional to its absolute temperature.
- **Boyle's law** states that under isothermal condition, pressure of a fixed amount of a gas is inversely proportional to its volume.
- **Avogadro law** states that equal volumes of all gases under same conditions of temperature and pressure contain equal number of molecules.
- **Graham's law** states that the rate of **diffusion** of a gas is inversely proportional to the square root of its molecular mass.

7.16 GLOSSARY

- **Diffusion:** the spreading of something more widely.
- **Limiting reactant:** the reactant that is consumed completely in a chemical reaction.
- **Mass:** the amount of matter in a particle or object.
- **Reactant:** a substance that takes part in and undergoes change during a reaction.
- **Stoichiometry:** is the quantitative study and calculation of relationships between the reactants and products in chemical reactions.

7.17 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Mole is the link between the number of .... and .... of atoms.

(*a*)Particles, mass (*b*) Electrons, mass

(*c*)Protons, mass (*d*) Particle, protons

2. The number of particles present in 1 mole of any substance is

(a)6.022×10^{23} (*b*) 6022×10^{22} (*c*) 60.22×10^{22} (*d*) 6.022×10^{20}

3. The molecular mass of nitric acid is

(*a*) 36 u (*b*) 48 u (*c*) 63 u (*d*) 84 u

4. The formula unit mass of NaCl is

(*a*)55.8 u (*b*) 85.5 u (*c*) 58.5 u (*d*) 55.5 u

5. Empirical formula of H_2O_2 is

(*a*)H_2O_2 (*b*) H_2O (*c*) HO (*d*) None of these

6. The mathematical relationship between pressure and temperature was given by

(*a*)Robert Boyle (*b*) Gay Laussac

(*c*)Amedeo Avogadro (*d*) A.C. Charles

7. The universal gas constant is denoted by

(*a*) R (*b*) V (*c*) N (*d*) T

8. Charles's law states that

(*a*)$V \propto T$(*b*) $V \propto P$ (*c*) $V \propto 1/p$ (*d*) $V \propto 1/t$

9. Graham's law is the relation between

(*a*)Rate of diffusion and density (*b*) Rate of diffusion and volume

(*c*)Rate of diffusion and mass (*d*) Both (*a*) and (*c*)

II. Open Ended Questions

1. What does the term mole mean?

2. Describe relative atomic mass and relative molecular mass.

3. Give the relationship between number of moles, mass and molar mass.

4. Define empirical formula.

5. Give an example in which molecular formula is same as empirical formula.

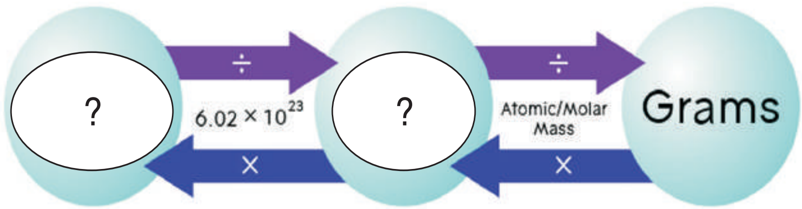
6. Explain limiting reagent.
7. With the help of an activity; explain Gay Lussac law.
8. Illustrate Charles' law and Boyle's law.
9. Explain ideal gas equation.
10. State Graham's law of diffusion.

III. Numericals

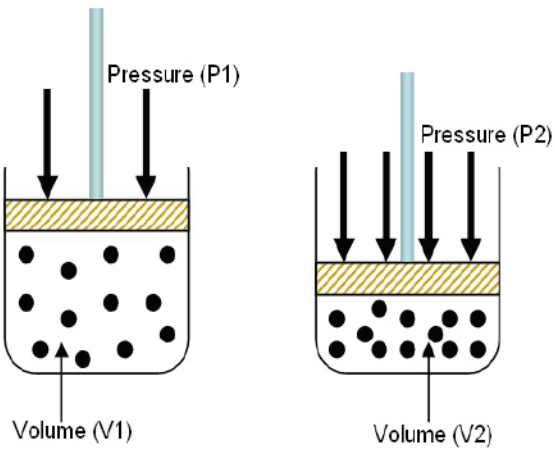
1. A vessel of 120 ml capacity contains a certain amount of gas at 35°C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 ml at 35°C. What would be its pressure?
2. A sample of helium has volume of 520 ml at 100°C. Calculate the temperature at which the volume will become 260 ml. Assume that pressure is constant.
3. Calculate the density of ammonia (NH_3) at 30°C and 5 bar pressure.
4. Calculate the volume occupied by 8.8 g of CO_2 at 31.1°C and 1 bar pressure. ($R = 0.083$ bar $dm^3 k^{-1}$ mol^{-1})

IV. Practical-based Questions

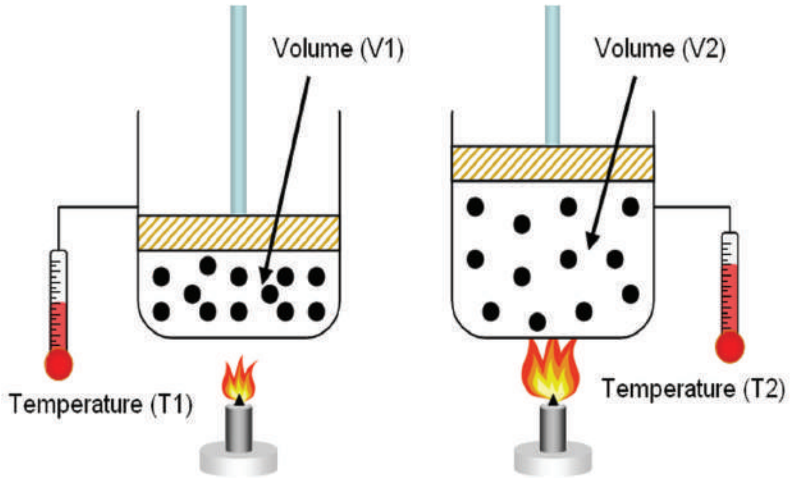
1. Complete the following representation.



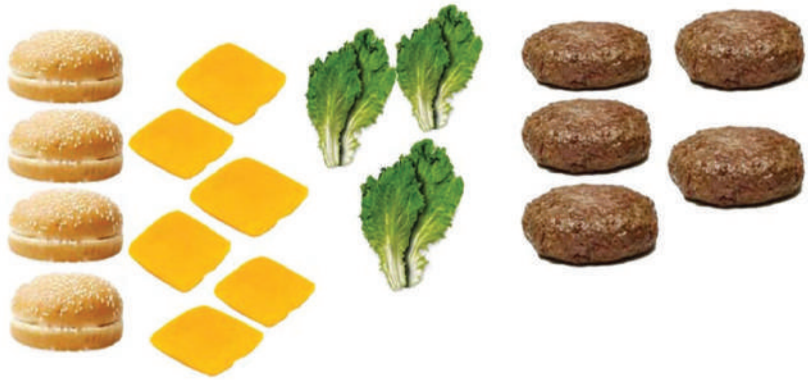
2. The following figure depicts ………………. law.



3. The following figure illustrates Charles' law. (True or False)



4. How many burgers can be made? Which part is limiting reagent?



PROJECT

Carry out an experiment of burning magnesium ribbon in air in order to determine the per cent composition of magnesium in magnesium oxide, and write a report on the findings.

Unit 8

Preparation and Classification of Oxides

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- categorise different oxides.
- explain how different oxides are formed.
- state the uses of different oxides.
- describe the reaction of oxides with other substances.
- describe the process of producing slaked lime.

KNOWLEDGE GAIN

Corrosion is a natural process which converts a refined metal to a more stable form such as its oxide or hydroxide.

ACTIVITY 8.1: Illustrating Types of Oxides

- Take a sample of magnesium, iron filings, sulphur and carbon.
- Burn each sample one by one.
- Observe the product formed and gas produced.

Was any residue left in case of non-metals?

Can you name the gas produced when carbon burns in air?

Write the chemical equation for each reaction.

8.1 PREPARATION OF OXIDES

We have already studied in Unit 2 that elements (metals and non-metals) react with oxygen to produce oxides. Oxides can be prepared by the following methods:

- Direct combination of an element with oxygen.
- Thermal decomposition of hydroxides, carbonates and nitrates.

Let us discuss these methods one by one.

8.1.1 Direct Combination of an Element with Oxygen

Combination of Metals with Oxygen

ACTIVITY 8.2: Illustrating Nature of Metal Oxide

Repeat Activity 8.1 with Calcium.

Make a report on the products obtained from the reaction of metals with oxygen.

Metals react with oxygen to produce metal oxides. Metal oxides are generally *basic* in nature with a few exceptions of amphoteric oxides such as Al_2O_3, ZnO and PbO.

$$\text{Metal} + \text{Oxygen} \longrightarrow \text{Metal Oxide}$$

Some elements react with oxygen at room temperature, some react on heating whereas others react only on strong heating. The reactions of oxygen with some elements is given below:

(*i*) Sodium combines with oxygen at room temperature with yellow flame to form sodium oxide.

$$\underset{\text{Sodium}}{4Na(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Sodium oxide}}{2Na_2O(s)}$$

Potassium also combines with oxygen at room temperature to produce potassium oxide. Try to write equation for this reaction.

(*ii*) Magnesium reacts with oxygen on heating to form magnesium oxide. It does not react with oxygen at room temperature.

$$\underset{\text{Magnesium}}{2Mg(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Magnesium Oxide}}{2MgO}$$

(*iii*) Aluminium reacts with oxygen on heating to form aluminium oxide

$$\underset{\text{Aluminium}}{4Al(s)} + \underset{\text{Oxygen}}{3O_2(g)} \longrightarrow \underset{\text{Aluminium oxide}}{2Al_2O_3(s)}$$

(*iv*) Zinc reacts with oxygen on strong heating to form zinc oxide

$$\underset{\text{Zinc}}{2Zn(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Zinc Oxide}}{2ZnO}$$

(*v*) Copper does not burn in air (oxygen) even on strong heating. Copper reacts with oxygen on prolonged heating to form a black substance (copper oxide).

$$2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$$

(*vi*) Iron reacts with oxygen on heating to form iron oxide.

$$\underset{\text{Iron}}{3Fe(s)} + \underset{\text{Oxygen}}{2O_2(g)} \longrightarrow \underset{\text{Iron oxide}}{Fe_3O_4(s)}$$

Note: Iron metal does not burn in air but iron filings (small particles of iron) burn vigorously when sprinkled on flame.

(*vii*) Calcium reacts with oxygen to form calcium oxide

$$\underset{\text{Calcium}}{2Ca(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\substack{\text{Calcium oxide}\\\text{(Alkali)}}}{2CaO(s)}$$

Calcium is a silvery white metal. The surface of calcium metal is covered with a thin layer of oxide that helps protect the metal from attack by air.

Figure 8.1: Calcium.

Calcium is quite reluctant to start burning, but once ignited, calcium burns in (reacts with) oxygen to give white calcium oxide.

EXERCISE 8.1

1. Metals react with oxygen to produce _______ .
2. Sodium combines with oxygen at room temperature to form sodium oxide.

   (True or False)
3. Write the molecular formula of iron oxide.
4. Name the metal which does not react with oxygen at room temperature?
5. Which of these formula are correct:

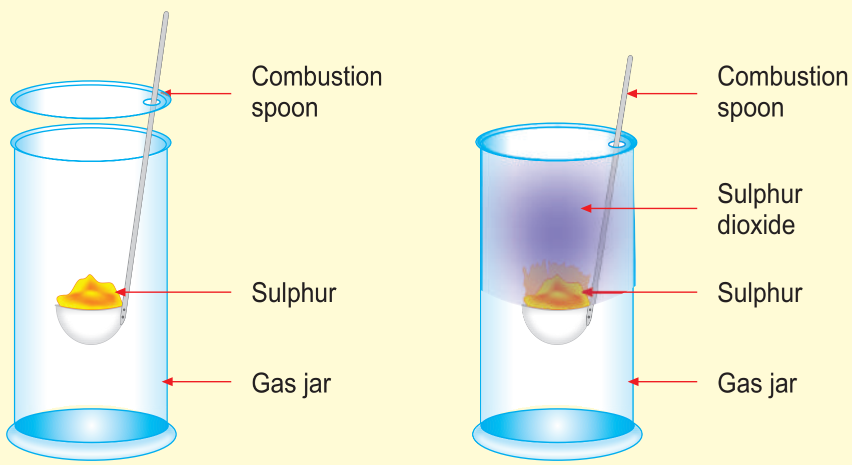
(*a*)Al_2O_3(*b*) Na_2O (*c*) Only (*a*) (*d*) both (*a*) and (*b*)

Combination of a Non-metal with Oxygen

ACTIVITY 8.3: Illustrating Nature of Non-metal Oxide

Take about 1 g of powdered sulphur in a deflagrating (combustion) spoon.

Heat the sulphur over the Bunsen burner, till it catches fire and starts burning with a blue flame. Lower the deflagrating spoon in a glass gas jar. Wait till the sulphur stops burning. Take out the deflagrating spoon from the gas jar and cover it with a glass disc.



Burning of sulphur

With the help of a dropper pour about 4–5 ml water in the gas jar and shake well. Test the water with the strips of red litmus paper and blue litmus paper.

Note your observations.

Make a report on the products obtained from the reaction of non-metals with oxygen.

In Activity 8.3, you will observe that blue litmus turns red, but there is no change in colour in case of red litmus paper. From this observation, it is evident that the product formed by the burning of sulphur dissolves in water and is acidic in nature.

Oxides of sulphur and nitrogen react with water to form acids. Sulphur forms sulphuric acid and nitrogen forms nitric acid. These acids can cause severe burns. It is thus important to handle them with care.

Non-metals react with oxygen to produce non-metal oxide. Non-metal oxides are generally *acidic* in nature.

$$\text{Non-metal} + \text{Oxygen} \longrightarrow \text{Non-metal oxide}$$

Similarly, other non-metals are also burnt in air to produce acidic oxides.

The equations for reaction of some non-metals with oxygen are:

(*i*) Sulphur combines with oxygen to give sulphur dioxide.

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

(*ii*) Carbon combines with oxygen to give carbon dioxide.

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

(*iii*) Hydrogen combines with oxygen to form neutral oxide water.

$$\underset{\text{Hydrogen}}{2H_2(g)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Water}}{H_2O(l)}$$

(*iv*) Phosphorus burns with dazzling white flame in oxygen forming phosphorus pentoxide.

$$4P + 5O_2 \longrightarrow 2P_2O_5$$

EXERCISE 8.2

1. Write the molecular formula of phosphorus pentoxide.
2. Complete the following chemical equation.
   $S(s) + O_2(g) \longrightarrow ?$
3. Non-metal oxides are generally ______ in nature.
4. Oxides of sulphur and nitrogen react with water to form acids. (True or False)
5. Which of these is/are not an example of non-metal oxides?
   (*a*) CO_2 (*b*) H_2O
   (*c*) SO_3 (*d*) none of these

8.1.2 Thermal Decomposition of Hydroxide, Carbonates and Nitrate

ACTIVITY 8.4: Illustrating Thermal Decomposition

Make a report on the products obtained from thermal decomposition of hydroxides, carbonates and nitrates.

When a decomposition reaction is carried out by heating, it is called thermal decomposition. Thermal means related to heat. In the following sections, we will study the formation of oxide by thermal decomposition.

Thermal Decomposition of Hydroxide

Metal hydroxide decomposes to form metal oxide and water. For example,

(*i*) Magnesium hydroxide on heating at a high temperature decomposes into magnesium oxide and water.

$$Mg(OH)_2(s) \xrightarrow{\text{Heat}} MgO(s) + H_2O(g)$$

(*ii*) Calcium hydroxide on heating decomposes into calcium oxide and water.

$$Ca(OH)_2(s) \xrightarrow{\text{Heat}} CaO(s) + H_2O(g)$$

(*iii*) Aluminium hydroxide decomposes into aluminium oxide and water.

$$2Al(OH)_3(s) \xrightarrow{\text{Heat}} Al_2O_3(s) + 3H_2O(g)$$

Thermal Decomposition of Carbonates

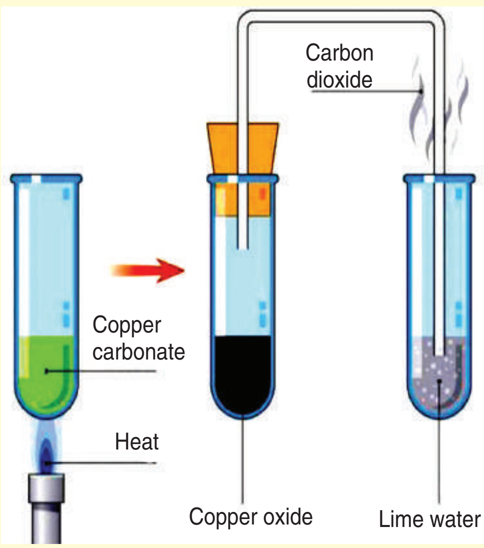
ACTIVITY 8.5: Thermal Decomposition of Copper Carbonate

Materials Required

Copper carbonate, Bunsen burner, boiling test tube.

Procedure

- Take a little amount of copper carbonate in a boiling test tube.
- Arrange all apparatuses as shown in the following figure.



- Heat the test tube on Bunsen burner.
- Observe the change.

In Activity 8.5, you will notice that copper carbonate (greenish) decomposes into copper oxide (black) and carbon dioxide. The carbon dioxide gas can be detected using lime water. Lime water turns cloudy white when carbon dioxide is bubbled through it.

Reaction

$$CuCO_3 \xrightarrow{\text{Heat}} CuO + CO_2$$

All the metal carbonates undergo thermal decomposition to give metal oxide and carbon dioxide. For example,

(*i*) When calcium carbonate is heated, it decomposes (breaks) to give calcium oxide and carbon dioxide.

$$\underset{\text{Calcium carbonate}}{CaCO_3(s)} \xrightarrow{\text{Heat}} \underset{\text{Calcium oxide}}{CaO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

(*ii*) Magnesium carbonate decomposes into magnesium oxide and carbon dioxide.

$$\underset{\text{Magnesium carbonate}}{MgCO_3(s)} \xrightarrow{\text{Heat}} \underset{\text{Magnesium oxide}}{MgO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

Note: The process of breaking up of magnesium carbonate at high temperature is known as **calcining**.

(*iii*) Zinc carbonate decomposes into zinc oxide and carbon dioxide.

$$\underset{\text{Zinc carbonate}}{ZnCO_3(s)} \xrightarrow{\text{Heat}} \underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

(*iv*) Lithium carbonate decomposes into lithium oxide and carbon dioxide.

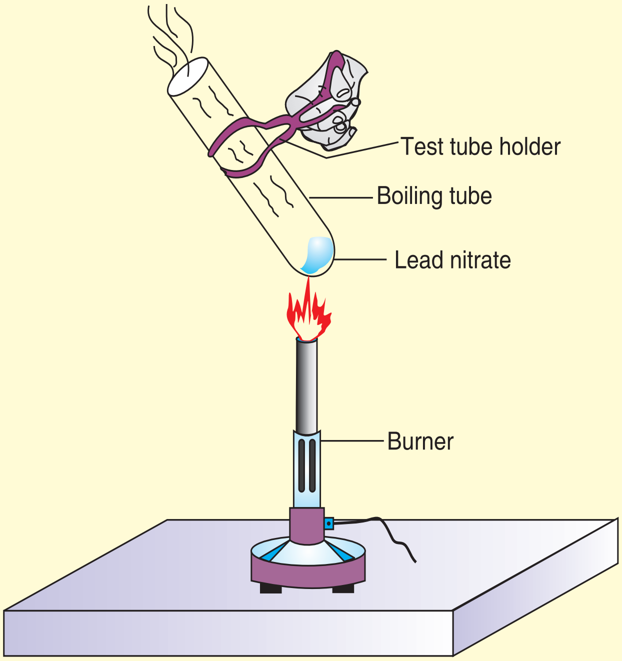
$$\underset{\text{Lithium carbonate}}{Li_2CO_3(s)} \xrightarrow{\text{Heat}} \underset{\text{Lithium oxide}}{Li_2O(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

Note: Carbonates of sodium and potassium decompose at very high temperatures. They do not decompose at Bunsen temperature.

Thermal Decomposition of Nitrates

ACTIVITY 8.6: Thermal Decomposition of Lead Nitrate

- Take about 0.5 g lead nitrate powder in a boiling tube.
- Hold the boiling tube with a pair of tongs.
- Heat it over a flame, as shown in figure.



- What do you observe? Note down the change, if any.

Safety

- Wear goggles, do not look directly into the boiling test tube.
- Do not touch the boiling test tube.
- Hold the boiling test tube with pair of tongs properly.

In Activity 8.6, you will observe the release of brown fumes. These fumes are of nitrogen dioxide (NO_2).

When lead nitrate is heated strongly, it breaks down to give lead monoxide, nitrogen dioxide and oxygen.

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

Most metal nitrates decompose on heating to give the metal oxide, brown fumes of nitrogen dioxide, and oxygen. Group I nitrates except $LiNO_3$ decompose to give metal nitrites and oxygen. Examples:

(*i*) Copper nitrate decomposes on gentle heating to give copper oxide, nitrogen dioxide and oxygen.

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

(*ii*) Lithium nitrate decomposes on heating to produce lithium oxide, nitrogen dioxide and oxygen.

$$4LiNO_3(s) \xrightarrow{\text{Heat}} 2Li_2O(s) + 4NO_2(g) + O_2(g)$$

Note: All nitrates of group 1 (from sodium to caesium) decompose to give metal nitrites and oxygen. As you go down the group you have to use higher temperature.

$2KNO_3 \longrightarrow 2KNO_2 + O_2$

EXERCISE 8.3

1. Metal hydroxide decomposes to form ______ and _______.

2. Complete the following reaction:

(i) $Ca(OH)_2(s) \xrightarrow{\text{Heat}} ? + ?$

(ii) $ZnCO_3\ (s) \xrightarrow{\text{Heat}} ? + ?$

3. Lithium nitrate decomposes on heating to produce ______, _______ and _______ .

4. What do you mean by thermal decomposition?

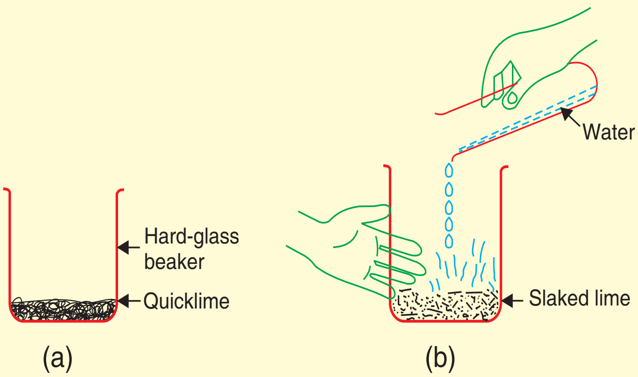
5. Carbonates of sodium and potassium decompose at very low temperature.

(True or False)

8.2 REACTION OF OXIDES WITH WATER

ACTIVITY 8.7: Illustrating Combination of Oxides with Water

- Take a little quicklime (CaO) in a hard beaker.
- Add water to it slowly as shown in the figure.
- Observe the change in temperature by touching the beaker.



Action of water on Quicklime

Can you tell whether the reaction is exothermic or endothermic?

In Activity 8.7, you will observe that calcium oxide reacts vigorously with water to form calcium hydroxide.

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2$$

The reaction between quicklime (calcium oxide) and water to form slaked lime is an exothermic reaction because it releases a lot of heat.

Most of the metal oxides are insoluble in water. Only a few metal oxides dissolve (react) in water to form alkalis*. Sodium oxide and potassium oxide are two metal oxides which are soluble in water. They dissolve in water to form alkali. The equations for reaction of metal oxides with water are:

(*i*) Sodium oxide reacts with water to form sodium hydroxide.

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

(*ii*) Potassium oxide reacts with water to form potassium hydroxide.

$$K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$$

(*iii*) Magnesium oxide dissolves partially in water to form magnesium hydroxide.

$$MgO(s) + H_2O(l) \longrightarrow Mg(OH)_2(aq)$$

* **Alkali:** Water soluble bases are called alkalis. For example, sodium hydroxide and potassium hydroxide.

ACTIVITY 8.8: Non-metal Oxides Dissolve in Water to form Acid

- Collect rainwater in a plastic container.
- Measure the pH of rainwater.
- Compare the colour of pH paper with the standard pH-colour chart.

What is the pH of collected rainwater?

Make a note on your observation.

In Activity 8.8, you will observe that the pH of rainwater is between 4.5 and 5.5 which shows it is slightly acidic. This is because of non-metal oxides. The non-metal oxides such as carbon dioxide, nitrogen dioxide, and sulphur dioxide are released in environment by industrial activities. These gases (non-metal oxides) combine with rainwater to form dilute carbonic acid, nitric acid, and sulphuric acid.

The non-metal oxides dissolve in water to form acids. For example,

(*i*) Carbon dioxide dissolves in water to form carbonic acid.

$$\underset{\text{Carbon dioxide}}{CO_2(g)} + \underset{\text{Water}}{H_2O(l)} \longrightarrow \underset{\text{Carbonic acid}}{H_2CO_3(aq)}$$

A solution of carbon dioxide gas in water turns blue litmus to red showing it is acidic in nature.

(*ii*) Sulphur dioxide reacts with water to give sulphurous acid.

$$\underset{\text{Sulphur dioxide}}{SO_2(g)} + \underset{\text{Water}}{H_2O(l)} \longrightarrow \underset{\text{Sulphurous acid}}{H_2SO_3(aq)}$$

A solution of sulphur dioxide and water turns blue litmus to red showing it is acidic.

EXERCISE 8.4

1. Name the product formed, when calcium oxide reacts with water.
2. Write the molecular formula of magnesium hydroxide.
3. Which of the following is correct?

(*a*) H_2SO_4 (*b*) H_2SO_3 (*c*) H_2CO_3 (*d*) all of these

4. Sulphur dioxide is acidic in nature. (True or False)
5. Name two metal oxides which are soluble in water.

8.3 REACTION OF OXIDES WITH ACIDS AND BASES

ACTIVITY 8.9: Illustrating Combination of Oxides with Acid and Bases

- Take a small amount of copper oxide in a beaker.
- Note the colour of copper oxide.
- Add dilute hydrochloric acid slowly while stirring with a glass rod.
- Observe the change:
  – Did copper oxide dissolve in dilute hydrochloric acid?
  – Did the colour of mixture change?

In Activity 8.9, you will observe that black colour of copper oxide changes to blue-green due to the formation of copper chloride.

Most of the metal oxides are basic** in nature but some metal oxides show both acidic and basic nature. The metal oxides which show both acidic*** and basic nature are *called amphoteric oxides*. Amphoteric oxides react with both acids and bases to form salt and water. Examples of amphoteric oxides are: aluminium oxide and zinc oxide.

The chemical equations for these oxides with acid and bases are:

(*i*) Aluminium oxide reacts with hydrochloric acid to form aluminium chloride (salt) and water.

$$\underset{\text{Aluminium oxide}}{Al_2O_3(s)} + \underset{\text{Hydrochloric acid}}{6HCl(aq)} \longrightarrow \underset{\text{Aluminium chloride (salt)}}{2AlCl_3(aq)} + \underset{\text{Water}}{3H_2O(l)}$$

In this reaction, aluminium oxide behaves as a basic oxide because it reacts with an acid to form salt and water.

(*ii*) Aluminium oxide reacts with sodium hydroxide to form sodium aluminate (salt) and water.

$$\underset{\text{Aluminium oxide}}{Al_2O_3(s)} + \underset{\text{Sodium hydroxide}}{2NaOH(aq)} \longrightarrow \underset{\text{Sodium aluminate (salt)}}{2NaAlO_2(aq)} + \underset{\text{Water}}{H_2O(l)}$$

In this reaction, aluminium oxide behaves as an acidic oxide because it reacts with a base to form salt and water.

(*iii*) Zinc oxide reacts with hydrochloric acid to form zinc chloride (salt) and water.

$$\underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \longrightarrow \underset{\text{Zinc chloride (salt)}}{ZnCl_2(aq)} + \underset{\text{Water}}{H_2O(l)}$$

In this reaction, zinc oxide behaves as a basic oxide. Do you know why?

** **Basic:** It turns red litmus solution into blue.
*** **Acidic:** It turns blue litmus solution to red.

(*iv*) Zinc oxide reacts with sodium hydroxide to form sodium zincate (salt) and water.

$$\underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Sodium hydroxide}}{2NaOH(aq)} \longrightarrow \underset{\substack{\text{Sodium zincate}\\ \text{(salt)}}}{Na_2ZnO_2(aq)} + \underset{\text{Water}}{H_2O(l)}$$

In this reaction, zinc oxide behaves as an acidic oxide. Do you know why?

EXERCISE 8.5

1. The colour of copper oxide is

(*a*) blue (*b*) black (*c*) green (*d*) yellow

2. Most of the metal oxides are ______ in nature.

3. Complete the following reactions:

(*i*) $Al_2O_3\ (s) + HCl\ (aq) \longrightarrow ? + ?$

(*ii*) $ZnO\ (s) + NaOH\ (aq) \longrightarrow ? + ?$

Also balance the both reactions.

4. Write balanced chemical equation for the following:

(*i*) When aluminium oxide reacts with sodium hydroxide.

(*ii*) When zinc oxide reacts with hydrochloric acid.

8.4 CLASSIFICATION OF OXIDES

ACTIVITY 8.10: Illustrating Acidity/Basicity of Oxides

Analyse the position of non-metals in the Periodic Table and observe the relationship between position of elements in the Periodic Table and acidity/basicity of oxides.

1	H																	He
2		Be											B	C	N	O	F	Ne
3													Al	Si	P	S	Cl	Ar
4												Zn		Ge	As	Se	Br	Kr
5				Zr										Sn	Sb	Te	I	Xe
6														Pb	Bi	Po	At	Rn

Forms no oxides; Forms amphoteric oxides; Forms acidic oxides

Forms basic oxides; Forms neutral oxides; Forms both acidic and neutral oxides

ACTIVITY 8.11: Illustrating Chemical Properties of MgO

- Take 2 g magnesium oxide in a test tube.
- Add 10 ml hydrochloric acid.
- Gently shake the test tube.
- Observe the change.
  – Did magnesium oxide dissolve in hydrochloric acid?
  – Did the temperature of test tube change?

Reaction

In Activity 8.11, you will notice that magnesium oxide dissolves to form a clear solution. Also the test tube becomes warm.

This implies that magnesium oxide reacts with hydrochloric acid. Therefore, magnesium oxide is basic in nature.

$$MgO(s) + 2HCl(aq) \longrightarrow MgCl_2 + H_2O$$

Figure 8.2: Magnesium chloride.

ACTIVITY 8.12: Illustrating Effect of Litmus Solution to Different Oxides

Safety: Wear goggles.

- Make a solution of different oxides such as sodium oxide, potassium oxide, magnesium oxide, etc. with water.
- Add 2–3 ml of each sample in a test tube.
- Put all test tubes in a test tube stand as shown in figure (see right).
- Add 2–3 drops of red litmus solution to each test tube.
- Observe and note if there is change in colour.
- If red litmus does not change to blue, add 2–3 drops of blue litmus solution.
  Did not any oxide change one of the litmus solutions?
- Record your result in tabular form.

Solutions of different oxides

Oxides are compounds of oxygen with another element. For example, magnesium oxide, sodium oxide, aluminium oxide, nitrogen oxide and calcium oxide. Depending upon the nature and the properties exhibited by these oxides, they are classified into:

- Acidic oxides
- Basic oxides
- Neutral oxides
- Amphoteric oxides

Metal/non-metal line

	1	2	13	14	15	16	17
2	Li_2O	BeO	B_2O_3	CO_2	N_2O_3 N_2O_5	O	OF_2
3	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_3 P_2O_5	SO_2 SO_3	Cl_2O_7
4	K_2O	CaO	Ga_2O_3	GeO_2	As_2O_3 As_2O_5	SeO_2 SeO_3	Br_2O
5	Rb_2O	SrO	In_2O_3	SnO_2	Sb_2O_5	TeO_3	I_2O_5
6	Cs_2O	BaO	Tl_2O_3	PbO_2	Bi_2O_5	Po	At

■ Basic oxide ■ Amphoteric oxide ■ Acidic oxide

Looking across a period on the periodic table, metals give way to non-metals. Similarly, there is a trend from basic oxides (that are ionically bonded) through a few amphoteric oxides to acidic oxides (that are covalently bonded). The noble gases, of course, do not form oxides.

8.4.1 Acidic Oxides

Acidic oxides are oxides of non-metals. They turn blue litmus solution to red. Most acidic oxides are soluble in water. They react with water to produce an acid. All acidic oxides react with alkali to give salt and water.

Example of acidic oxides are sulphur dioxide, sulphur trioxide, carbon dioxide, phosphorus pentoxide.

Properties of Acidic Oxides

1. Acidic oxides do not react with acids.
2. They react with bases or alkalis to form salt and water.
3. They dissolve in water to form acidic solutions.
4. They are usually gases at room temperature.

8.4.2 Basic Oxides

Basic oxides are oxides of metals. They turn red litmus solution to blue. Most metal oxides are insoluble in water but some of these dissolve in water to form alkalis. They react with water to produce a base. All basic oxides react with acid to form salt and water.

Example of basic oxides are sodium oxide, magnesium oxide, potassium oxide, copper oxide and calcium oxide.

Properties of Basic Oxides

1. Basic oxides do not react with bases.
2. They react with acids to form salt and water.
3. The basic oxides are usually insoluble in water. Those that dissolve in water form alkaline solutions.

8.4.3 Neutral Oxide

Some non-metals react with oxygen to form neutral oxides. Neutral oxides do not show neither acidic nor basic characteristics. For example, carbon monoxide, nitric oxide, and nitrous oxide. These compounds are also called neutral compounds.

Properties of Neutral Oxides

1. Neutral oxides do not react with either acids or bases.
2. They have no effect on litmus solution.

8.4.4 Amphoteric Oxides

Some metals react with oxygen to produce amphoteric oxides. Amphoteric oxides exhibit both acidic and basic characteristics.

These oxides react with acids as well as bases to form salt and water. Example of amphoteric oxides are aluminium oxide and zinc oxide.

Properties of Amphoteric Oxides

1. Amphoteric oxides react with both acids and bases.
2. They change red litmus to blue and blue litmus to red.
3. They undergo neutralisation reaction to give salt and water.

> **Note:** Amphoteric oxides show basic as well as acidic behaviour. Amphoteric oxides react with both acids and bases to form salt and water.

Classification of oxides as basic, acidic, amphoteric and neutral is given in Table 8.1.

Table 8.1: Classification of Oxides

Acidic Oxides	Basic Oxides[a] (Alkaline oxides)	Amphoteric Oxides[b]	Neutral Oxides
• Lead dioxide	• Sodium oxide	• Zinc oxide	• Water
• Chromium trioxide	• Magnesium oxide	• Aluminium oxide	• Carbon monoxide
• Carbon dioxide	• Copper oxide		• Nitrogen oxide
• Sulphur dioxide	• Calcium oxide		
• Sulphur trioxide	• Potassium oxide		
• Silicon dioxide			
• Phosphorus Pentoxide			

[a] Basic oxides are oxides of metals, especially alkali and alkaline earth metals.

[b] Other elements which form amphoteric oxides are gallium, indium, scandium, titanium, zirconium, silver, gold, germanium, tin, antimony, bismuth, and tellurium.

EXERCISE 8.6

1. Name the types of oxides. Give one example of each.
2. (*i*) ________ are oxides of non-metals.

(*ii*)________ are oxides of metals.

(*iii*)Some metals react with oxygen to produce ________.

(*iv*)Some non-metals react with oxygen to produce ________.

3. Make a list of oxides of your choice and categorize them into their respective types .

8.5 USES AND PRODUCTION OF SLAKED LIME (*ISHWAGARA*)

Calcium oxide (lime or quicklime) reacts vigorously with water to form calcium hydroxide (slaked lime):

$$\underset{\substack{\text{Calcium}\\\text{oxide}\\\text{(lime or}\\\text{quicklime)}}}{CaO\,(s)} + \underset{\text{Water}}{H_2O\,(l)} \xrightarrow{\text{Combination}} \underset{\substack{\text{Calcium}\\\text{hydroxide}}}{Ca(OH)_2\,(s)}$$

This is a combination reaction in which two compounds, calcium oxide and water combine to form single compound calcium hydroxide. A large amount of heat is released when calcium oxide reacts with water to form calcium hydroxide. Solid calcium hydroxide is also known as slaked lime (*ishwagara*). Slaked lime is a white powder.

A solution of slaked lime produced by the reaction of quicklime and water is used for white washing walls. Calcium hydroxide reacts slowly with the carbon dioxide in air to form a thin layer of calcium carbonate on the walls. Calcium carbonate is formed after two to three days of white washing and gives a shiny finish to the walls. It is interesting to note that the chemical formula for marble is also $CaCO_3$.

$$\underset{\substack{\text{Calcium}\\\text{hydroxide}}}{Ca(OH)_2\,(aq)} + CO_2\,(g) \longrightarrow \underset{\text{Calcium carbonate}}{CaCO_3\,(s)} + H_2O\,(l)$$

EXERCISE 8.7

1. Write two uses of *ishwagara*.
2. Chemical formula of marble is:

(*a*)$Ca(OH)_2$ (*b*) CaO (*c*) $CaCO_3$ (*d*) $CaCl_2$

3. A large amount of heat is absorbed when lime reacts with water. (True or False)
4. Slaked lime is a ________ powder.
5. Solid calcium hydroxide is also known as:

(*a*)Marble (*b*) Talcum (*c*) Granite (*d*) None of these.

8.6 SUMMARY

- Oxides are compounds of oxygen with another element.
- Metals combine with oxygen to give basic oxides or amphoteric oxides.
- Non-metals combine with oxygen to form acidic oxides or neutral oxides.
- Amphoteric oxides exhibit both acidic and basic characteristics.
- Neutral oxides show neither acidic nor basic characteristics.
- Basic oxide turns red litmus solution into blue.
- Acidic oxide turns blue litmus solution into red.
- Oxides can be prepared by thermal decomposition of carbonates, hydroxides and nitrates.
- Metal oxides react with acid to give salt and water.
- A few metal oxides react with water to form alkalis.
- Non-metal oxides react with water to produce acid.
- Slaked lime (*ishwagara*) is a white powder produced by reaction of quicklime and water.
- Slaked lime is used for white washing and formation of calcium carbonate.

8.7 GLOSSARY

- **Alkaline:** having a pH greater than 7
- **Dazzling:** extremely bright
- **Covalent:** the sharing of electrons between atoms.
- **Deflagrate:** sharp combustion, burning with flame.
- **Deflagrating spoon:** a long vertical handled spoon with cover used in deflagration experiments.
- **Endothermic:** a reaction or process accompanied by or requiring the absorption of heat.
- **Exothermic:** a reaction or process accompanied by the release of heat.
- **Ignition:** the action of setting something on fire or starting to burn.
- **Prolonged:** continuing for a long time.
- **Reluctant:** resistant
- **Sprinkled:** spray
- **Thermal decomposition:** a chemical decomposition caused by heat.

8.8 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Metals react with oxygen to produce

(*a*) basic oxides (*b*) amphoteric oxides

(*c*) both (*a*) and (*b*) (*d*) neither (*a*) nor (*b*)

2. Non-metals react with oxygen to give

(*a*)acidic oxides (*b*) neutral oxides (*c*) both (*a*) and (*b*) (*d*) none of these

3. Metal which reacts with oxygen only on strong heating is

(*a*) zinc (*b*) sodium (*c*) potassium (*d*) iron

4. Metal hydroxide decomposes to form

(*a*)metal oxide and water (*b*)metal oxide and carbon dioxide

(*c*)both (*a*) and (*b*) (*d*) none of these

5. The colour of nitrogen dioxide gas is

(*a*) red (*b*) brown (*c*) blue (*d*) white

6. Some metal nitrates decompose to give

(*a*)metal oxide (*b*) nitrogen oxide (*c*) oxygen (*d*) all of these

7. The reaction between quicklime and water is

(*a*)an exothermic reaction (*b*) an endothermic reaction

(*c*)both (*a*) and (*b*) (*d*) none of these

8. Aluminium oxide and zinc oxide are examples of

(*a*)amphoteric oxides (*b*) neutral oxides

(*c*)basic oxides (*d*) none of these

9. Nitric oxide and nitrous oxide are examples of

(*a*)amphoteric oxides (*b*) basic oxides

(*c*)neutral oxides (*d*) none of these

10. Production of slaked lime from calcium oxide and water is a

(*a*)combination reaction (*b*)combustion reaction

(*c*)single replacement reaction (*d*)double replacement reaction

II. Open Ended Questions

1. Compare acidic oxide and basic oxide.
2. Distinguish between amphoteric and neutral oxides.
3. List two examples of each
   (*i*) Acidic oxide
   (*ii*) Basic oxide
   (*iii*) Amphoteric oxide
   (*iv*) Neutral oxide
4. Explain the preparation of oxide by thermal decomposition of carbonate.
5. State the properties of acidic oxides and basic oxides.
6. What happens when
   (*i*) calcium oxide reacts with water

(*ii*) magnesium combines with oxygen
(*iii*) sulphur burns in oxygen
(*iv*) lithium carbonate decomposes
(*v*) carbon dioxide dissolves in water

7. Write balanced chemical equation for each reaction of Q6.
8. Make a table showing a list of acidic and basic oxides.
9. Can you name some compounds which decompose to produce metal oxide and carbon dioxide.
10. Carry out an experiment to show the product obtained when magnesium reacts with oxygen. Write a balanced equation.

PROJECT

Collect some corroded and some fresh metals. Observe and make a presentation distinguishing corroded and fresh metal. Present your observations in the class.

Unit 9

Electrolytes and Non-electrolytes

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- define an electrolyte and a non-electrolyte.
- give examples of weak and strong electrolytes and non-electrolytes.
- state applications of electrolytes in daily life.

KNOWLEDGE GAIN

Trophies and medals are electroplated with gold and silver.

9.1 ELECTROLYTE AND NON-ELECTROLYTE

ACTIVITY 9.1: Distinction between Electrolyte and Non-electrolyte

Materials Required : Salt (NaCl) solution, sugar solution, pure water, 3 nine-volt cells, 6 carbon electrodes, 3 bulbs, 3 beakers and wires.

- Take 3 beakers. Label them A, B and C.
- Half-fill the beaker A with pure water, half-fill the beaker B with sugar solution, half-fill the beaker C with salt solution.
- Set all apparatuses as shown in figures.

Pure water, $H_2O(l)$ does not conduct electricity

Sucrose solution, $C_{12}H_{22}O_{11}(aq)$ non-electrolyte does not conduct electricity

Sodium chloride solution, $NaCl(aq)$ electrolyte conducts electricity

Do you know, why bulb glows only in the sodium chloride solution?

In Activity 9.1, you will observe that only sodium chloride solution conducts electricity. This is because the solution contains ions—Na^+ and Cl^-. Sodium chloride (NaCl) ionises into Na^+ and Cl^- in water. It is an *electrolyte*. Pure water and sugar solution do not contain any ion. So they do not conduct electricity. Sugar is a non-electrolyte.

Non-electrolytes are covalent compounds that do not dissociate into ions when they are dissolved in water. They dissolve in water as molecules. Sugar and urea are good examples of non-electrolytes.

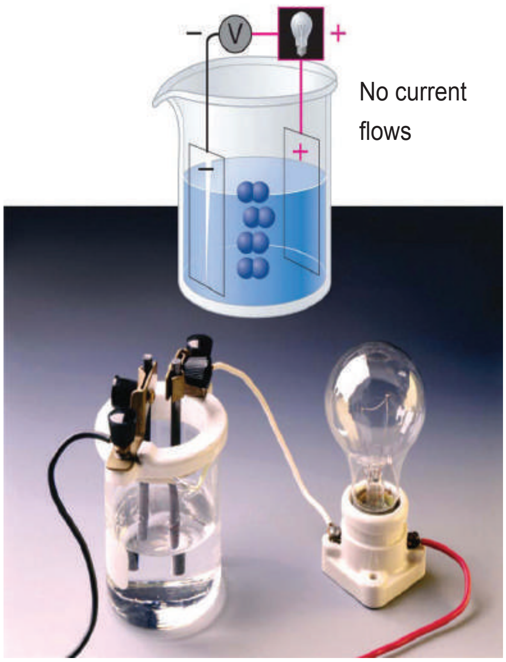
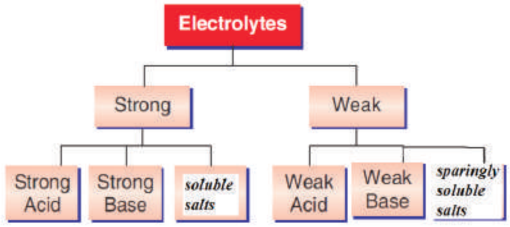


Figure 9.1: Non-electrolyte.

Electrolytes are compounds whose aqueous solution contains ions. In other words, a substance that dissociates into ions* when they are dissolved in water. For example, sodium chloride, magnesium chloride, copper sulphate and ammonium chloride. All electrolytes are divided into two groups—strong electrolytes and weak electrolytes.



EXERCISE 9.1

1. Pure water and sugar solution do not contain ________.
2. Non-electrolytes are ____________ compounds.
3. Give two examples of non-electrolytes.
4. What do you mean by electrolytes?
5. Sodium chloride is an example of electrolyte. (True or False)

9.2 DEFINITION OF ELECTROLYSIS

Electrolysis is the process by which ionic substances are decomposed (broken down) into simpler substances when an electric current is passed through them. Ions are free to move when an ionic substance is dissolved in water. In Unit 5, we have already studied the electrolysis of water.

9.3 STRONG ELECTROLYTE

An electrolyte which is completely ionised in water and thus produces a large amount of ions is called **strong electrolyte**. For example, hydrochloric acid, nitric acid and sulphuric acid, sodium hydroxide and potassium hydroxide, sodium chloride and

* An ion is an electrically charged atom or group of atoms. The positively charged ions are called cations and the negatively charged ions are called anions.

potassium nitrate.

These electrolytes have high electrical conductivity (Figure 9.2) because of high concentration of ions in their solution.

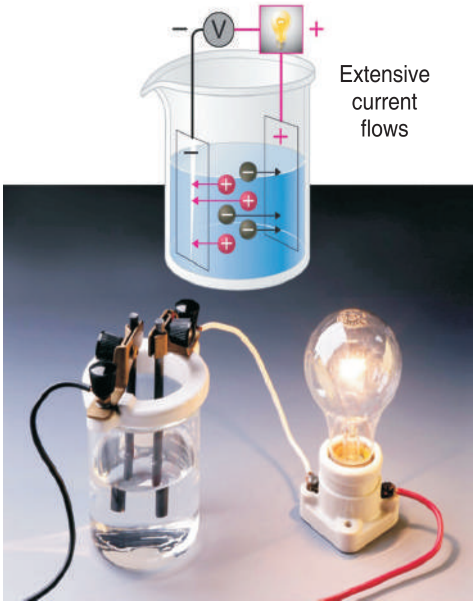


Figure 9.2: Strong electrolyte.

Hydrochloric acid, sulphuric acid and nitric acid are also called mineral acids. The acids prepared from the minerals of the earth are called mineral acid. These are very strong acids and react very rapidly with metals, metal carbonates and metal hydrogen carbonates. In laboratory, these acids are mixed with water to dilute them.

While using mineral acids, follow all safety measures and handle them very carefully because they cause severe burns on skin and eat up clothes, wood, metals and stone due to their corrosive nature. These are never stored in metal containers. Mineral acids are stored in containers made up of glass.

EXERCISE 9.2

1. What do you mean by strong electrolytes?
2. __________ and __________ are mineral acids.
3. Give two examples of strong electrolytes.
4. Write two safety measures to use sulphuric acid.
5. In laboratory, dilutes acids are stored in metallic container. (True or False)

9.4 WEAK ELECTROLYTE

An electrolyte which is partially ionised in water and thus produces a small amount of ions is called **weak electrolyte**. For example, acetic acid, carbonic acid, sulphurous acid, organic acid, ammonium hydroxide, calcium hydroxide, magnesium hydroxide, barium sulphate and silver nitrate. These electrolytes react quite slowly with metals, metal carbonates, and metal hydrogen carbonates.

Weak electrolytes have low electrical conductivity (Figure 9.3) because of low concentration of ions in their solution.

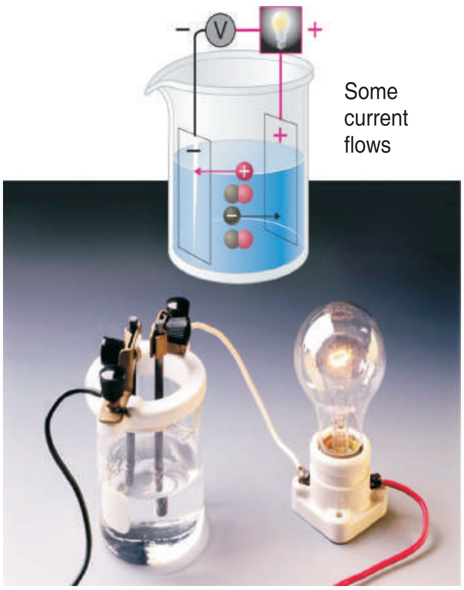


Figure 9.3: Weak electrolyte.

Organic Acids

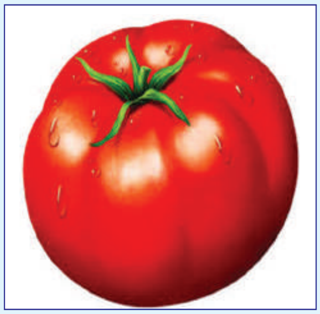
The acids present in fruits are called **organic acids**. Organic acids occur naturally. Examples of organic acids are citric acid, lactic acid, oxalic acid and tartaric acid. Organic acids are weak acids. It is not harmful to eat fruits containing organic acids.

Sources of organic acids

- Citric acid is present in citrus fruits such as lemons and oranges.

Lemon

- Lactic acid is present in curd or sour milk.
- Oxalic acid is present in tomatoes.



Tomato

- Tartaric acid is present in unripe grapes and tamarind.

Tamarind

EXERCISE 9.3

1. What do you mean by weak electrolytes?
2. Write three examples of weak electrolytes.
3. The acids present in fruits are called __________.
4. Name the acid present in

(*a*) milk (*b*) lemon.

5. Why weak electrolytes have low electrical conductivity?

9.5 CONDUCTIVITY OF ELECTRICITY BY ELECTROLYTES

ACTIVITY 9.2: Identifying the Conductivity of Electricity by Electrolytes

Chemicals Required

Hydrochloric acid, aqueous ammonia, Sodium chloride solution, Sugar solution, Calcium hydroxide solution, Ethanol, Sodium hydroxide solution, tap water, Acetic acid (ethanoic acid), distilled water.

Materials Required

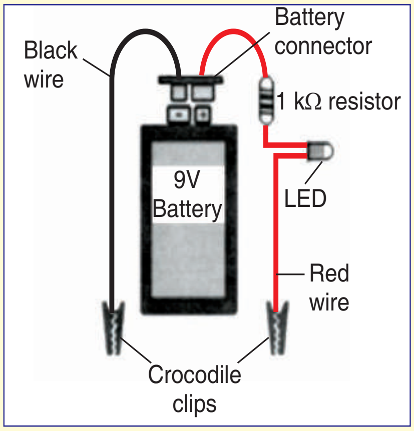
Homemade conductivity tester, and beaker (50 ml).

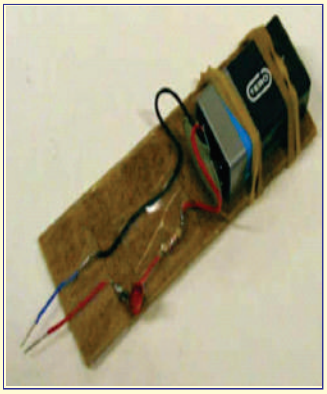
Procedure

Prepare your homemade conductivity tester as given below.

Materials

- 1 light emitting diode (LED)
- 1 kΩ resistor
- 1 nine-volt battery
- 1 battery connector
- 1 piece of fibre board or plywood
- 1 rubber band
- Half metre black wire and half metre red wire
- Tape





Homemade Conductivity Tester

Table: Conductivity Measurements

LED Brightness	Conductivity
Off	Non-electrolyte
Dim	Weak electrolyte
Medium	Weak electrolyte
Bright	Strong electrolyte
Very bright	Strong electrolyte

- Take about 25 ml hydrochloric acid in a 50 ml beaker.
- Dip the open wire ends in the beaker to test if they conduct electricity.
- Observe the brightness of the LED in tester.
- Compare your observation with the given table.
- Light will glow if the electrolyte is strong or weak. It will not glow if it is a non-electrolyte
- Repeat the activity with each sample.
- Make an appropriate report of your observation.

Caution

Be careful not to touch the exposed wire of the electrodes that are used for testing!

9.6 APPLICATION OF ELECTROLYTES

ACTIVITY 9.3

Research and make presentations about the applications of electrolytes in daily life.

The following are application of electrolytes:

- Sodium chloride (NaCl) is more commonly known as "salt" or "table salt," and is the major ingredient in the edible salt that is sprinkled on food to make it taste better. Sodium chloride is present in the ocean and is the main thing that makes ocean water taste so salty.
- Nitric acid (HNO_3) is a corrosive acid. It is used in the manufacture of fertiliser. It is also used as one of the ingredients in certain types of liquid rocket fuel.
- Hydrochlroic acid (HCl) is a strong acid widely used in the chemical industry.
- Calcium chloride ($CaCl_2$) is also a salt and is different from the table salt. Calcium chloride is a compound made of calcium and chloride, and is one of the types of salt used to control ice on sidewalks and roadways. The compound is often produced from limestone.
- Potassium nitrate (KNO_3) is used in a wide variety of substances. It can be used as a food additive, but is also an ingredient in various types of rocket fuels and in fireworks. Potassium nitrate was an ingredient in gun powder for many years.
- Sodium hydroxide (NaOH) is an important ingredient in many detergents, soaps, and drain cleaners. It is highly dangerous because of its ability to decompose lipids and proteins in skin, causing burns when not handled properly.
- Sulphuric acid (H_2SO_4) is used in car batteries to get the electricity to flow. Like sodium hydroxide, it can cause very severe chemical burns when it comes into contact with skin, so it is very important to handle with extreme caution.
- Ammonium chloride (NH_4Cl) is a main component of dry cell. Dry cell is commonly known as Leclanché cell. The other electrolyte used in dry cell is manganese dioxide (MnO_2).
- Sodium acetate (CH_3COONa) is an electrolyte that includes sodium as a main ingredient. It is often used to seal concrete so that it is protected against bad weather. It is also used as an ingredient

in certain types of foods, such as salt and vinegar potato chips, because of its salty and tangy flavour when mixed with other seasonings.

- Magnesium hydroxide ($Mg(OH)_2$) is also known as milk of magnesia because of its milk-like appearance. It is a main component of many types of laxatives and antacids, and in underarm deodorants and antiperspirants. It can also be applied to the scalp for dandruff control.
- $CuSO_4$ is used as an electrolyte in copper refining
- $AgNO_3$ is used for electroplating other metals

EXERCISE 9.4

1. Fill in the Blanks

(*a*) Nitric acid is a __________ acid.

(*b*) Sodium hydroxide is used in the making __________ and __________.

(*c*) __________ is used in car batteries for electricity.

(*d*) __________ is a main component of dry cell.

(*e*) Magnesium hydroxide is also known as __________.

9.7 SUMMARY

- Electrolytes are compounds whose aqueous solution contains ions.
- Electrolytes are grouped into—strong electrolytes and weak electrolytes.
- An electrolyte which ionises completely in water is called strong electrolyte. For example, H_2SO_4, HCl and HNO_3.
- Strong electrolytes have high electrical conductivity.
- An electrolyte which ionises partially in water is called weak electrolyte. For example, CH_3COOH, $Ca(OH)_2$ and $Mg(OH)_2$.
- Weak electrolytes have low electrical conductivity.
- Electrolysis is the process by which ionic substances are decomposed (broken down) into simpler substances when an electric current is passed through them
- The acids present in fruits are called organic acids.
- The acids prepared from the minerals of the earth are called mineral acids.
- Sulphuric acid is used in car batteries.
- Ammonium chloride and manganese dioxide are used in Leclanché cell(dry batteries).

9.8 GLOSSARY

- **Antacid:** a medicine used to reduce excess acid in the stomach.
- **Antiperspirants:** a substance that is applied to the skin, especially under the arms, to reduce sweating.
- **Concrete:** a building material made from a mixture of broken stone or gravel, sand, cement, and water.
- **Corrosive:** a substance that gradually distroys living tissues on other matter.

- **Dandruff:** small pieces of dead skin in a person's hair.
- **Deodorants:** a substance which removes unpleasant smells, especially bodily odours.
- **Detergents:** a water-soluble cleansing agent which combines with impurities and dirt to make them more soluble.
- **Diode:** a semiconductor device with two terminals, typically allowing the flow of current in one direction only.
- **Ingredient:** any of the foods or substances that are combined to make a particular dish.
- **Laxative:** a medicine used to treat and prevent constipation.
- **Lipids:** an important component of living cells, a fatty acids that stores energy in living cells.
- **Severe:** intense, extreme, rigorous.
- **Sprinkle:** a small quantity or amount of something scattered over an object or surface.
- **Tangy:** having a strong, piquant flavour or smell.
- **Urea:** a water-soluble compound containing nitrogen, carbon, oxygen and hydrogen used as fertiliser.

9.9 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Sugar solution does not contain

(*a*) atom (*b*) molecule (*c*) ion (*d*) none of these

2. Sodium chloride and copper sulphate are examples of

(*a*) electrolytes (*b*) compounds (*c*) both (*a*) and (*b*) (*d*) elements

3. Sucrose and urea are good examples of

(*a*) strong electrolyte (*b*) weak electrolyte (*c*) non-electrolyte (*d*) all of these

4. During electrolysis of water, hydrogen gas is produced __________ as compared with oxygen.

(*a*) two times (*b*) three times (*c*) four times (*d*) five times

5. Generally, strong electrolytes react very rapidly with

(*a*) metals (*b*) metal carbonates

(*c*) metal hydrogen carbonates (*d*) all of these

6. Citric acid and tartaric acid are examples of

(*a*) mineral acid (*b*) organic acid (*c*) only (*a*) (*d*) only (*b*)

7. Choose the weak electrolyte(s).

(*a*) Carbonic acid (*b*) Ammonium hydroxide

(*c*) Calcium hydroxide (*d*) All of these

8. Choose the mineral acid.

(*a*) Sulphuric acid (*b*) Hydrochloric acid

(*c*) Nitric acid (*d*) All of these

9. Choose the correct statement(s).

(*a*) Hydrochloric acid is widely used in chemical industry

(*b*) Magnesium hydroxide is also known as milk of magnesia

(*c*) Nitric acid is used in the manufacture of fertiliser

(*d*) All of the above

10. Electrolyte used in car batteries to get the electricity is ______ .

(*a*) sulphuric acid (*b*) nitric acid (*c*) sodium acetate (*d*) calcium chloride

II. Open Ended Questions

1. Give two examples of

(*a*) Electrolytes (*b*) Non-electrolytes

2. Distinguish between strong and weak electrolytes. Give examples.
3. Explain non-electrolytes.
4. What do you mean by organic acid and mineral acid?
5. How is strong acid stored in laboratory?
6. Carry out an experiment to differentiate between electrolytes and non-electrolytes.
7. Name the electrolyte used in/as

(*a*) car batteries (*b*) fertiliser

(*c*) table salt (*d*) detergents

8. What are the applications of electrolytes in daily life?

III. Practical-based Questions

Look at the figures and answer the following questions.

Beaker A Beaker B Beaker C

1. Which of the beakers contains non-electrolyte?

   (*a*) Beaker A (*b*) Beaker B (*c*) Beaker C (*d*) All of these

2. Which beaker contains solution of weak electrolyte?

   (*a*) Beaker A (*b*) Beaker B (*c*) Beaker C (*d*) All of these

3. Which beaker contains solution of strong electrolyte?

   (*a*) Beaker A (*b*) Beaker B (*c*) Beaker C (*d*) All of these

4. Which beaker contains solution of organic acid?

   (*a*) Beaker A (*b*) Beaker B (*c*) Beaker C (*d*) None of these

5. The beaker B may contain

   (*a*) Pure water (*b*) Glucose solution

   (*c*) Urea solution (*d*) All of these

Unit 10

Properties of Organic Compounds and Uses of Alkanes

LEARNING OBJECTIVES

After reading this unit, you will be able to:

- identify organic compounds and their origin.
- describe the physical and chemical properties of alkanes (methane).
- state the uses of methane and some other alkanes.
- explain structural isomerism.

KNOWLEDGE GAIN

The candle wax contains a mixture of alkanes having high molecular masses.

ACTIVITY 10.1: Distinguishing Daily Used Things as Organic or Inorganic

Make a list of things which you
- eat or drink daily
- wear daily
- use daily

Can you distinguish the objects you have listed as organic or inorganic?

Do you know whether milk is organic or not?

Make an album showing organic compounds used in our daily life.

The compounds of carbon are known as organic compounds. The oxides of carbon, carbonates, hydrogen carbonate and carbides are not organic. All living things, plants and animals are composed of organic compounds. The term 'Organic' is derived from organism meaning a living body. Initially, organic compounds were considered to be derived from animal and vegetable sources.

It was observed that inorganic compounds such as sodium chloride can be prepared directly from their elements in the laboratory but organic compounds such as sugar and urea could not be prepared from their elements.

This led to the common belief that the organic compounds could only be synthesised in nature by living organisms. In 1828, Friedrich Wöhler, a German chemist, prepared the first organic compound (urea) in the laboratory. This artificial synthesis of urea shook the common belief and it was discarded. Later, methane and acetic acid were prepared in the laboratory.

A large number of things which we use in daily life are organic compounds. For example,

- Our food materials such as grains, pulses, milk, sugar, etc.
- Our clothes materials such as cotton, silk, wool, nylon, etc.
- Fuels such as petrol, diesel, coal, wood, LPG, etc.

Other materials such as paper, cotton, plastics, leather are also carbon compounds. Carbon compounds such as carbohydrate, proteins, amino acids, vitamins, fats, oil and enzymes play important roles in life process. Without carbon compounds life is not possible.

10.1 ORGANIC CHEMISTRY

Organic chemistry is a discipline of chemistry that is concerned with the study of carbon compounds except oxides, carbonates, hydrogencarbonates, cyanides and carbides. Organic chemistry encompasses the synthesis, identification, modelling, and chemical reactions of such compounds.

Organic chemistry focuses on the structure, properties, and applications of various carbon-containing molecules that make up important biological molecules such as proteins, enzymes, carbohydrates, lipids, nucleic acids, and vitamins. A number of organic compounds such as cotton, wool, and natural petroleum are found in nature. There are many others that are purely man-made. Industries have been able to synthesise many organic compounds such as plastics, photographic film, synthetic fabrics, and pharmaceutical drugs using applications of organic chemistry.

10.2 DIFFERENCE BETWEEN ORGANIC AND INORGANIC CHEMISTRY

Chemistry is often divided into two subgroups known as organic and inorganic. Organic chemistry focuses on all compounds that contain the element carbon, usually in the form of carbon and hydrogen bonds. Inorganic chemistry deals with the compounds that do not have carbon, though there are some exceptions. Many of these inorganic compounds are classified as salts.

ACTIVITY 10.2: Distinction between Organic and Inorganic Compounds

Chemicals Required

Organic compounds: Urea, Sugar, Ethanol
Inorganic compounds: NaCl, MgOH, $Ca(OH)_2$

Materials Required

Homemade conductivity tester (as you have made in Activity 9.2)

Procedure

- Take 150 ml aqueous solution of urea in a beaker.
- Dip the open wire ends in the beaker.
- Observe the LED.
- Note your observation.
- Repeat the activity with each sample.
- Make an appropriate report of your observation.

Caution

- Do not insert the open wire end of tester into the AC plug.

In Activity 10.2, you will observe that aqueous solution of organic compounds do not conduct electricity, whereas inorganic compounds conduct electricity in aqueous solution.

Table 10.1 summarises the difference between organic and inorganic compounds.

Table 10.1: Distinction between Organic and Inorganic Compounds

	Organic Compounds	Inorganic Compounds
1.	Organic compounds usually contain carbon, hydrogen, and oxygen. They always contain carbon-hydrogen bonds with some exceptions such as urea $H_2N\text{-}CO\text{-}NH_2$.	Inorganic compounds contain all the known elements. They may contain carbon but do not contain carbon-hydrogen bonds.
2.	Organic compounds are mainly found in most of the living things.	Inorganic compounds are found in non-living things.
3.	Generally, they are insoluble in water but soluble in organic solvents.	They are generally soluble in water and non-soluble in organic solvents.
4.	They are highly inflammable.	They are non-inflammable.

5.	In most of the aqueous solutions, organic compounds are poor conductors of heat and electricity.	In aqueous solutions, inorganic compounds are known to be good conductors of heat and electricity.
6.	They have low melting and boiling points compared to inorganic compounds.	They have high melting and boiling points compared to organic compounds.

10.3 OCCURRENCE OF ORGANIC COMPOUNDS

Simplest organic compounds occur naturally in crude oil and natural gas. Crude oil and natural gas are usually found together, formed by the same gradual decay of marine animals and plants. Crude oil is a mixture of different hydrocarbons; most of which are alkanes.

The alkanes are heavily exploited as fuels. Alkanes are obtained from crude oil by fractional distillation. Fractional distillation separates crude oil into different fractions, that is, the various hydrocarbon components.

The important sources of methane are:

- It occurs in natural gas (90% CH_4 + small amount of C_2H_6, C_3H_8 and C_4H_{10}) above the layer of petroleum under earth.
- It is one of the chief constituents (32–34%) of coal gas.
- It is the major constituent of gobar gas, sewage gas and biogas.
- It occurs in marshy lands where it is formed due to the microbial decomposition of vegetable and animal organic matter such as plant wastes, animal dung, poultry sweeps, etc.

10.4 HOMOLOGOUS SERIES

Hydrocarbons with related structures and properties are usually separated into "families" known as **homologous series**. Each member of a series differs from the next member by a single carbon atom and two hydrogen atoms. With each successive member, the properties of the compounds, such as melting and boiling points, change in a regular way.

- All members of a homologous series can be represented by same general formula.
- All the compounds of a homologous series show similar chemical properties.
- The members of a homologous series show a gradual change in their physical properties with increase in molecular mass.
- Any two adjacent members of a homologous differ by a —CH2 group.

10.5 GENERAL FORMULA OF ALKANES

In alkanes, each carbon atom is bonded directly to four other atoms. Since in these compounds, all the carbon atoms utilise their full combining capacity by forming single covalent bonds with other atoms, these compounds are called **saturated compounds**.

General formula of homologous series of alkanes is C_nH_{2n+2}, where n is a positive integer. Thus, **alkanes** may be defined as *the saturated hydrocarbons having general formula* C_nH_{2n+2}.

10.6 NOMENCLATURE OF ALKANES

International Union of Pure and Applied Chemistry (IUPAC) Rules for Naming Straight Chain Alkanes

The name of a straight chain alkane, may be divided into two parts: ***Prefix*** and ***suffix***. Prefix designates the number of carbon atoms in the chain. For chains containing one to four carbon atoms, special prefixes are used and for chains of five or more carbon atoms, Greek number prefixes are used. IUPAC prefixes for a few carbon chains are given in Table 10.2.

Table 10.2: Prefixes for Carbon Chain Lengths

Chain length	Prefix	Chain length	Prefix
C_1	Meth-	C_6	Hex-
C_2	Eth-	C_7	Hept-
C_3	Prop-	C_8	Oct-
C_4	But-	C_9	Non-
C_5	Pent-	C_{10}	Dec-

In order to derive the IUPAC name, a suffix is added to the prefix. The primary suffix used for alkanes is "**ane**".

IUPAC names of straight chain alkanes containing carbons 1–10 are given in Table 10.3.

Table 10.3: Names of Alkanes

Condensed structural formula	IUPAC name	Molecular formula
CH_4	Methane	CH_4
$CH_3\,CH_3$	Ethane	C_2H_6
$CH_3\,CH_2\,CH_3$	Propane	C_3H_8
$CH_3CH_2CH_2CH_3$	Butane	C_4H_{10}
$CH_3CH_2CH_2CH_2CH_3$	Pentane	C_5H_{12}
$CH_3CH_2\,CH_2\,CH_2\,CH_2CH_3$	Hexane	C_6H_{14}
$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$	Heptane	C_7H_{16}
$CH_3CH_2CH_2\,CH_2CH_2CH_2CH_2CH_3$	Octane	C_8H_{18}
$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$	Nonane	C_9H_{20}
$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$	Decane	$C_{10}H_{22}$

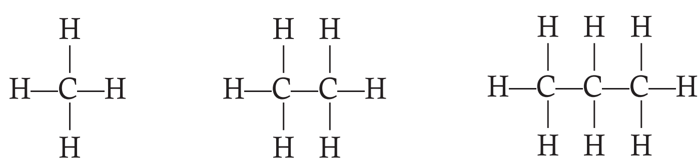
10.7 STRUCTURAL ISOMERISM

Isomerism Isomerism is the phenomenon of chemical compounds having the same molecular formula but with different arrangements of atoms.

Structural isomerism is a form of isomerism in which compounds have the same molecular formula but with different structural formulae.

Alkanes can exhibit structural isomerism.

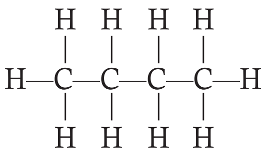
Each of the first three members of alkanes; methane, ethane and propane has only one structural formula as shown below. Structural formula shows the arrangement of atoms in a molecule of compound.



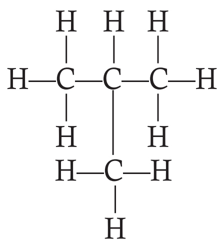
Methane Ethane Propane

Thus, *no structural isomerism is possible up to propane.*

The fourth member, that is, C_4H_{10}, can have two structural formulae as shown below:



(n-Butane) Butane



2-Methylpropane
(isobutane)

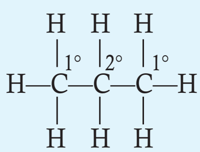
An alkane with unbranched chain of carbon atoms is called normal alkane or *n*-alkane.

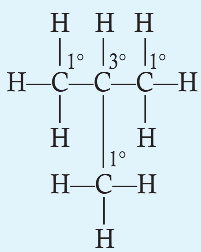
Classes of Carbon Atoms

Carbon atoms in alkanes can be classified on the basis of the number of other carbon atoms to which they are attached.

A carbon atom which is attached directly to only one is classified as a primary (1°) carbon atom.

A carbon atom attached directly to two other carbon atoms is called a secondary (2°) carbon atom. Similarly, a tertiary (3°) carbon atom is attached directly to three other carbon atoms.





Example 1

$_1CH_3—_2CH_2—_3CH_2—_4CH_2—_5CH_3$

Prefix used for 5 carbon atoms is Pent

Suffix is ane

IUPAC Name: *n*-Pentane

Example 2

$_1CH_3—_2CH_2—_3CH_2—_4CH_2—_5CH_2—_6CH_2—_7CH_3$

Prefix used for 7 carbon atoms is Hept

Suffix is ane

IUPAC Name: *n*-Heptane.

Example 3

$_1CH_3—_2CH_2—_3CH_2—_4CH_2—_5CH_2—_6CH_3$

Prefix used for 6 carbon atoms is Hex

Suffix is ane

IUPAC Name: *n*-Hexane.

10.8 PHYSICAL PROPERTIES OF ALKANES

ACTIVITY 10.3: Illustrating Physical Properties of Alkanes

In groups, discuss the physical properties of alkanes.

1. **Physical state:** Alkanes contain only C—C and C—H bonds. They have very small difference of electronegativity between carbon and hydrogen atoms. The first four members (C_1 to C_4) are gases; the next thirteen members, (C_5 to C_{17}) are liquids while the higher members are waxy solids.
2. **Solubility:** alkanes are soluble in organic solvents such as ether, ethanol, carbon tetrachloride, etc. and insoluble in polar solvents such as water.
3. **Boiling Points:** Alkanes generally have low boiling points. *The boiling points of*

n-alkanes increase regularly with the increase in the number of carbon atoms.

4. **Melting Points:** The melting points of alkanes generally increase with the increase in molecular mass. For example, let us look at the melting points of propane to *n*-octane in Table 10.4.

   The physical properties of the first ten straight chain alkanes are given in Table 10.4.

5. **Density:** The densities of alkanes increase with increasing molecular masses but become constant at about 0.8 g cm^{-3}. This means that liquid alkanes are lighter than water.

Table 10.4: Physical Properties of the First 10 n-Alkanes

Alkane	Formula/Physical State (at 20°C)	Boling Point (K)	Melting Point (K)	Density (g cm^{-3}) at 20°C
Methane	$CH_4(g)$	109	90	0.55
Ethane	$C_2H_6(g)$	184	90	0.51
Propane	$C_3H_8(g)$	231	84	0.50
Butane	$C_4H_{10}(g)$	273	135	0.58
Pentane	$C_5H_{12}(l)$	309	143	0.63
Hexane	$C_6H_{14}(l)$	342	178	0.66
Heptane	$C_7H_{16}(l)$	371	182	0.68
Octane	$C_8H_{18}(l)$	399	216	0.70
Nonane	$C_9H_{20}(l)$	424	222	0.72
Decane	$C_{10}H_{22}(l)$	447	243	0.73

EXERCISE 10.1

1. Alkanes contain only _______ and ______ bonds.
2. The first fifteen members of alkanes are gases. (True or False)
3. The density of alkanes is

(*a*) more than water

(*b*) less than water

(*c*) equal to water

(*d*) cannot be calculated.

4. Which of the following is not the correct formula for alkanes?

(*a*) C_2H_2 (*b*) C_2H_4

(*c*) C_3H_6 (*d*) all of these

5. Alkanes generally have low boiling points. (True or False)

10.9 CHEMICAL PROPERTIES OF ALKANES

ACTIVITY 10.4: Illustrating Chemical Properties of Alkanes

In groups, discuss the chemical properties of alkanes.

Alkanes are saturated hydrocarbons. They contain only strong C—C and C—H bonds. Hence, alkanes are quite less reactive and are known as **paraffins** (*parum*–little, *affinis*–reactivity or affinity). Alkanes are generally inert towards acids, bases, oxidising and reducing agents. Alkanes being saturated compounds can not undergo addition but substitution and combustion reactions. Thus, the typical reactions of alkanes are substitution reactions.

Important reactions of alkanes are as follows:

1. **Combustion:** Alkanes on burning in air or oxygen get completely oxidised to carbon dioxide and water with the evolution of large amount of heat.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + 890\ kJ$$

The general chemical equation for combustion of alkanes may be written as:

$$C_nH_{2n+2} + \left(\frac{3n+1}{2}\right)O_2 \longrightarrow nCO_2 + (n+1)H_2O$$

In the presence of limited supply of air, alkanes burn to give carbon monoxide, which is a highly poisonous gas.

$$2CH_4 + 3O_2 \longrightarrow 2CO + 4H_2O$$

DO YOU KNOW?

Carbon black is formed during incomplete combustion of alkanes with insufficient amount of air or oxygen. It is used in the manufacture of ink, printer ink, black pigments and as filters.

$$CH_4(g) + O_2(g) \xrightarrow[\text{Combustion}]{\text{Incomplete}} C(s) + 2H_2O(l).$$

2. **Halogenation:** *Halogenation involves the replacement of one or more H atoms of alkane by corresponding number of halogen atoms.*

Chlorination can be carried out by treatment of alkane with chlorine, either in the presence of diffused sunlight or ultraviolet light or at high temperature (573–773 K). For example, when methane is chlorinated, we get mixture of four different substitution products.

$$CH_4 + Cl_2 \xrightarrow[\text{or Heat}]{\text{Diffused sunlight}} \underset{\text{Chloromethane}}{CH_3Cl} + HCl$$

$$CH_3Cl + Cl_2 \longrightarrow \underset{\text{Dichloromethane}}{CH_2Cl_2} + HCl$$

$$CH_2Cl_2 + Cl_2 \longrightarrow \underset{\text{Trichloromethane}}{CHCl_3} + HCl$$

$$CHCl_3 + Cl_2 \longrightarrow \underset{\text{Tetrachloro-methane}}{CCl_4} + HCl$$

3. **Thermal cracking:** Cracking means the breaking down of long chain saturated hydrocarbon to form a mixture of alkane and alkenes of shorter length. When alkanes are heated to very high temperatures in the absence of air, they split into smaller molecules. Cracking of alkanes is a thermal decomposition that involves the breaking of carbon-carbon bonds to form smaller molecules of hydrocarbons.

$$C_{10}H_{22} \xrightarrow[550°C]{\text{Heat}} C_8H_{16} + C_2H_4$$

Cracking of ethane gives ethene and hydrogen.

EXERCISE 10.2

1. Alkanes are also known as _______.
2. Alkanes on burning in air/oxygen get completely oxidised to ______ and ______.
3. Complete the following reactions:

(*a*) $CH_4 + Cl_2 \xrightarrow{?} CH_3Cl + ?$

(*b*) $C_{10}H_{22} \xrightarrow[600\,°C]{} ? + ?$

4. Define thermal cracking of alkanes.
5. _______ is used in the manufacture of printer ink and black pigments.

10.10 LABORATORY PREPARATION OF METHANE

Methane can be prepared in the laboratory by heating a mixture of anhydrous sodium ethanoate and soda-lime. Soda-lime is a mixture of sodium hydroxide (NaOH) and calcium oxide (CaO) in the ratio 3 : 1.

$$\underset{\text{Sodium ethanoate}}{CH_3COONa} + \underset{\text{Sodium hydroxide}}{NaOH} \longrightarrow \underset{\text{Methane}}{CH_4} + \underset{\text{Sodium carbonate}}{Na_2CO_3}$$

- Take a mixture of anhydrous sodium ethanoate and soda-lime in the ratio 2 : 1 in a hard glass test-tube fitted with a delivery tube as shown in Figure 10.1.

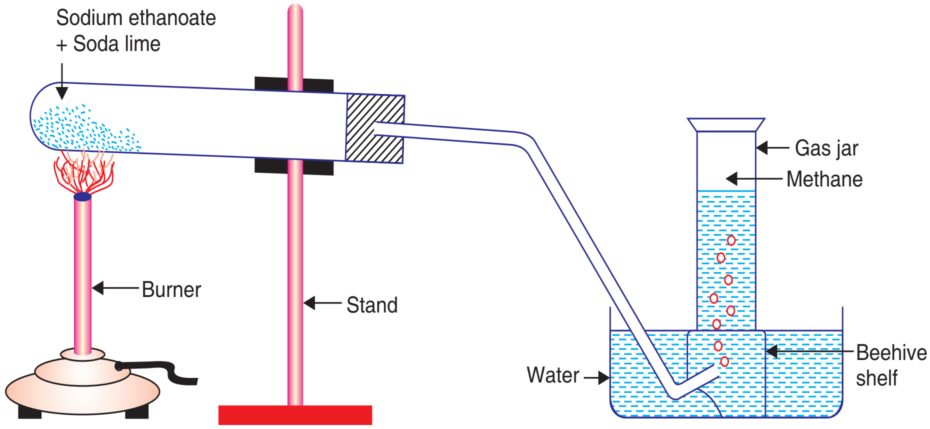


Figure 10.1: Laboratory preparation of methane.

- Heat the mixture slowly.
  Observe what happens and record.
  Methane gas is produced which is collected in a gas jar by the downward displacement of water. As methane gas is collected over water, it shows that *methane is insoluble in water.*

In Rwanda, methane gas is extracted in Lake Kivu. Methane is used as source of heat.

EXERCISE 10.3

1. Methane can be prepared in the laboratory by heating a mixture of _______ and ______.
2. Methane is soluble in water. (True or False)
3. In laboratory, methane gas is collected in a gas jar by _______ of water.
4. What is soda lime?
5. Methane is used as a source of heat. (True or False)

ACTIVITY 10.5: Illustrating Extraction of Methane

Carry out research and make a presentation on how methane gas is naturally formed. (A field visit may be necessary to a biogas plant).

10.11 USES OF ALKANES

Important uses of alkanes are as follows:

1. Petroleum is largely a mixture of different alkanes. On refining, petroleum gives Liquefied Petroleum Gas (LPG), gasoline, kerosene, diesel, furnace oil and wax which are used as fuels.

Natural gas is mainly methane. Compressed Natural Gas (CNG) is used as fuel to run automobiles.

2. Some higher alkanes are used as lubricating oils and as vaseline. Vaseline is used as a basic material in many cosmetics.
3. Higher alkanes on sulphonation yield corresponding sulphonic acids, which are used for preparing detergents.
4. Alkanes are used as starting materials for the preparation of many other useful organic compounds. For example, methane on chlorination gives chloromethane, dichloromethane, trichloromethane (chloroform) and tetrachloromethane which are used as solvents.
5. Carbon formed during decomposition of methane is in the form of finely divided particles and is known as **carbon black**. Carbon black is used for making printer's ink and paints. It is also used in rubber industry.
6. Natural gas is a rich ***source of hydrogen*** gas which is needed in the manufacture of fertilisers.

EXERCISE 10.4

1. _______ is largely a mixture of different alkanes.
2. Name two products obtained on refining of petroleum.
3. Write full form of (*a*) LPG, (*b*) CNG.
4. What do you mean by carbon black?
5. Alkanes are used as lubricating oils and vaseline. (True or False)

10.12 SUMMARY

- Organic chemistry is the branch of chemistry that deals with the carbon compounds.
- In organic chemistry, we study the structure, properties, composition, reaction and preparation of carbon containing compounds. The most organic (carbon) compounds comprise at least one carbon-hydrogen bond.
- Alkanes occur naturally in crude oil and natural gas. Crude oil and natural gas are formed by the gradual decay of marine animals and plants.
- Crude oil is a mixture of different hydrocarbons; most of which are alkanes.
- The compounds in a homologous series:
  - ➢ are represented by a general formula.
  - ➢ show gradation in physical properties.
  - ➢ have similar chemical properties because they have the same type of bonding.
  - ➢ have successive members differing by a CH_2 group.
- General formula of homologous series of alkanes is C_nH_{2n+2}, where n is a positive integer.
- In order to derive the IUPAC name, a suffix is added to the prefix. The primary suffix used for alkanes is "ane".

- Isomerism is the phenomenon of two chemical compounds having the same molecular formula but with different arrangements of atoms.
- Structural isomerism is a form of isomerism in which compounds have the same molecular formula but with different structural formulae.
- No structural isomerism is possible from methane to propane.
- Alkanes are colourless. The first four members of alkanes (C_1 to C_4) are gases; the next thirteen members (C_5 to C_{17}) are liquids while the higher members are waxy solids.
- The boiling points and melting points of alkanes directly correspond to the size of the molecule. The melting and boiling points of the shorter chain alkanes are low, but the melting and boiling points of alkanes increase as the number of carbon atoms in the carbon chain increases.
- Alkanes are insoluble in water but soluble in ether. The densities of alkanes increase with increasing molecular masses but become constant at about 0.8 g cm^{-3}. Liquid alkanes are less dense than water (alkanes float on top of water).
- Alkanes react rapidly with oxygen releasing energy, which makes alkanes useful as fuels. Alkanes react with halogens such as chlorine gas and bromine water in the presence of ultraviolet light.
- Cracking of alkanes is a thermal decomposition that involves the breaking of carbon-carbon bonds to form smaller molecules of hydrocarbons.
- Common name of trichloromethane is chloroform.
- Methane can be prepared in laboratory by heating a mixture of anhydrous sodium ethanoate and soda-lime.

$$CH_3COONa + NaOH \longrightarrow CH_4 + Na_2CO_3$$

- Soda-lime is a mixture of sodium hydroxide (NaOH) and calcium oxide (CaO) in the ratio 3:1.
- Alkanes are used as fuels.
- Alkanes are used for preparing detergents, lubricating oils and vaseline.

10.13 GLOSSARY

- **Cosmetics:** substances used to enhance the appearance or odour of the human body.
- **Hydrocarbon:** a compound of hydrogen and carbon.
- **Inorganic compound:** any compound that lacks a carbon atom plus CO_2, CO, carbonates, hydrogencarbonates, cyanides and carbides.
- **Lipids:** a group of naturally occurring molecules that include fats, waxes, fat-soluble vitamins such as vitamins A, D, E, and K.
- **Organic compounds:** compounds containing carbon except carbon oxides, carbonates, hydrogencarbonates, carbides and cyanides.

- **Pharmaceutical drugs:** a drug used to diagnose, cure, treat, or prevent disease.

10.14 UNIT ASSESSMENT

I. Multiple Choice Questions

1. Organic chemistry deals with the compounds containing

(*a*) carbon (*b*) phosphorus (*c*) silicon (*d*) chlorine

2. Alkanes occur naturally in

(*a*) crude oil (*b*) natural gas (*c*) both (*a*) and (*b*) (*d*) none of these

3. The compounds in a homologous series have

(*a*) Same general formula (*b*) Similar chemical properties

(*c*) Differ by a CH_2 group. (*d*) All of these

4. The general formula of alkane is

(*a*) C_nH_{2n} (*b*) $C_n H_{2n-2}$ (*c*) C_nH_{2n+2} (*d*) none of these

5. IUPAC name of alkane containing four carbon atoms is

(*a*) Butane (*b*) Propane (*c*) Pentane (*d*) Hexane

6. IUPAC name of $CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$ is

(*a*) Octane (*b*) Heptane (*c*) Heptene (*d*) Hexane

7. The fourth member of alkane has .... structural isomers.

(*a*) one (*b*) two (*c*) three (*d*) four

8. Alkanes are soluble in

(*a*) ether (*b*) CCl_4 (*c*) both (*a*) and (*b*) (*d*) water

9. Alkane reacts with excess oxygen to give

(*a*) $CO_2 + H_2O$ (*b*) $CO + H_2O$

(*c*) $CO_2 + H_2O$ + Heat (*d*) $CO + H_2O$ + Heat

10. Methane is prepared in laboratory by heating a mixture of

(*a*) $CH_3COONa + NaOH$ (*b*) $CH_3COONa + Ca(OH)_2$

(*c*) $CH_3COOH + NaOH$ (*d*) $CH_3COOH + Ca(OH)_2$

II. Open Ended Questions

1. Distinguish between organic and inorganic chemistry.

2. Compare and contrast the physical properties of organic and inorganic compounds.

3. Define homologous series.

4. Write IUPAC names of first ten alkanes.

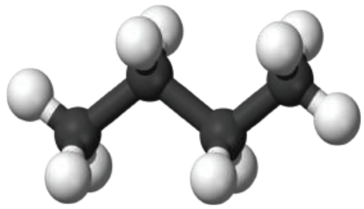
5. Discuss the physical and chemical properties of alkane.

6. How is methane gas prepared in the laboratory?
7. State the uses of alkanes in daily life.

III. Practical-based Questions

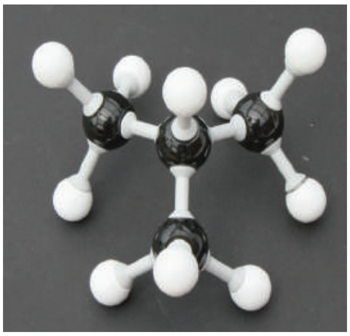
General Instruction: In the following figures, black balls represent carbon and white balls represent hydrogen.

1. What is the name of the organic compound shown in the figure?



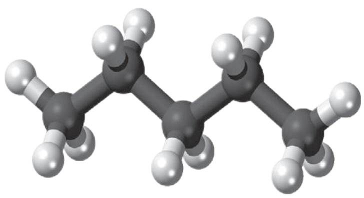
(*a*) Methane (*b*) Ethane (*c*) Propane (*d*) Butane

2. Name the compound shown in the following figure.

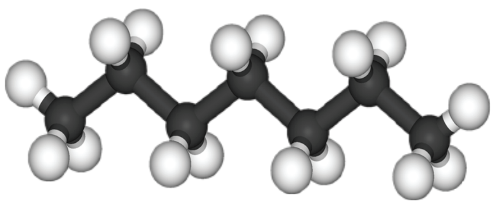


(*a*) Propane (*b*) iso-butane (*c*) Butane (*d*) None of these

3. Which compound has higher melting point?



Compound A



Compound B

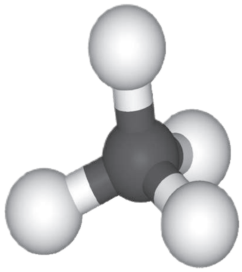
(*a*) Compound A

(*b*) Compound B

(*c*) Both compounds have same melting point

(*d*) Compounds do not exist

4. Choose the correct statement for the following compound.



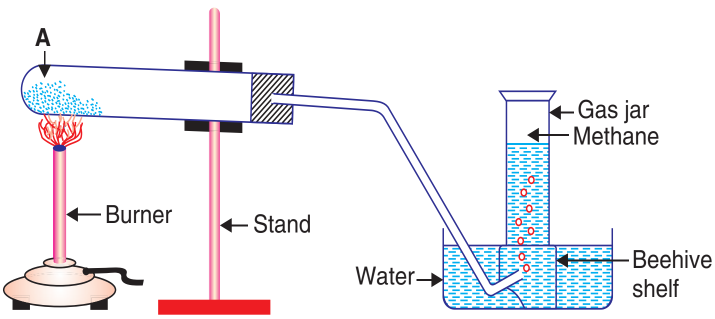
(*a*) This is the simplest alkane

(*b*) It can be prepared in the laboratory

(*c*) This can be used as a source of heat

(*d*) All of the above

5. If the figure shows the laboratory preparation of methane, which of the given statements is correct?



(*a*) A = Sodium Acetate + Soda Lime

(*b*) A = Sodium Chloride + Soda Lime

(*c*) A = Sodium Carbonate + Slaked Lime

(*d*) A = Sodium Hydrogencarbonate + Lime

PROJECT

Make a 3D model of heptane using plastic balls and sticks.

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