Table of Contents

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E - book

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Assessment

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DIGI links

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Learning Objectives

At the end of the lesson the students will be able to:

- acquire the ability to learn about the atoms and molecules.
- comprehend atomic mass and molecular mass.
- have information about gram atomic mass and gram molecular mass.
- perceive the intended meaning of Avogadro's hypothesis of gases.
- interpret the application of Avogadro's hypothesis.
- determine the atomicity of a molecule.
- interpret the relation between vapour density and relative molecular mass.
- have the facts about the relationship between the volume of a gas and the number of molecules present in it.
- grasp the idea of mole concept and solve many problems using it.
- calculate the percentage of composition of a compund.

INTRODUCTION

You have learnt, in your lower classes that matter is around us everywhere. Matter is made of atoms. Curiously the idea of atom was first proposed by the Greek philosophers in the fifth century BC (BCE). But, their theory was more philosophical than scientific.

The first scientific theory of the atom was proposed by John Dalton. Few of the postulates of Dalton's theory about an atom were found incorrect by the later on studies made by J.J. Thomson, Rutherford, Neils Bohr and Schrodinger. In the light of the result of the researches most of the limitations of the Dalton's theory were removed and a new theory known as the modern atomic theory was put forward. **'The main postulates of modern atomic theory'** are as follows:

- **An atom is no longer indivisible** (after the discovery of the electron, proton, and neutron).
- Atoms of the same element may have different atomic mass. (discovery of **isotopes** $_{17}Cl^{35}$, $_{17}Cl^{37}$).
- Atoms of different elements may have same atomic masses (discovery of **Isobars** $_{18}Ar^{40}$, $_{20}Ca^{40}$).
- Atoms of one element can be transmuted into atoms of other elements. In other words, atom is no longer indestructible (discovery of **artificial transmutation**).

- Atoms may not always combine in a simple whole number ratio (E.g. Glucose $C_6H_{12}O_6$ C:H:O = 6:12:6 or 1:2:1 and Sucrose $C_{12}H_{22}O_{11}$ C:H:O = 12:22:11).
- Atom is the **smallest particle that takes part in a chemical reaction.**
- The mass of an atom can be converted into energy ($E = mc^2$).

The modern atomic theory is the basis for all the studies of chemical and physical processes that involve atoms. You have studied the most fundamental ideas about an atom in your lower classes. Let us discuss some more concepts about atoms in this lesson.

7.1 ATOM AND ATOMIC MASS

As you know, anything that has mass and occupies space is called matter. Atoms are the building blocks of matter. Since matter has mass, it must be due to its atoms. According to the modern atomic theory, an atom contains subatomic particles such as protons, neutrons and electrons. **Protons and neutrons have considerable mass, but electrons don't have such a considerable mass**. Thus, the mass of an atom is mainly contributed by its protons and neutrons and hence **the sum of the number of protons and neutrons of an atom is called its mass number**.

Individual atoms are very small and it is difficult to measure their masses. You can measure the mass of macroscopic materials in gram or kilogram. The mass of an atom is measured in atomic mass unit (amu).

Atomic mass unit is one-twelfth of the mass of a carbon-12 atom; an isotope of carbon, which contains 6 protons and 6 neutrons.

(**Note:** The symbol ‘amu’ is no longer used in the modern system and instead, it uses the symbol ‘U’ to denote unified atomic mass. The mass of a proton or neutron is approximately 1 amu).

7.1.1 Relative Atomic Mass (RAM)

As an atom is very small, its absolute mass cannot be determined directly. The early pioneers of chemistry used to measure the atomic mass of an atom relative to an atom of another element. They measured the masses of equal number of atoms of two or more elements at a time, to determine their relative masses. They established one element as a standard, gave it an arbitrary value of atomic mass and using this value they measured the relative mass of other elements. The mass obtained by this way is called relative atomic mass. In the beginning, the mass of hydrogen atom was chosen as a standard and masses of other atoms were compared with it, because of the existence of isotopic character of hydrogen (${}_1H^1$, ${}_1H^2$, ${}_1H^3$). Later hydrogen atom was replaced by oxygen atom as the standard. Now, the stable isotope of carbon (C-12) with atomic mass 12 is used as the standard for measuring the relative atomic mass of an element.

Relative atomic mass of an element is the ratio between the average mass of its isotopes to $\frac{1}{12^{th}}$ part of the mass of a carbon-12 atom. It is denoted as A_r. It is otherwise called “Standard Atomic Weight”.

Relative Atomic Mass

$$(A_r) = \frac{\text{Average mass of the isotopes of the element}}{\frac{1}{12^{th}} \text{ of the mass of one Carbon-12 atom}}$$

Modern methods of determination of atomic mass by Mass Spectrometry uses C-12 as standard. For most of the elements, the relative atomic mass is very closer to a whole number and it is rounded off to a whole number, to make calculations easier. Table 7.1 lists some of the elements of periodic table and their A_r values.

Table 7.1 Relative atomic mass of elements (C-12 Scale)

Element	Symbol	A_r
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24
Sulphur	S	32

DO YOU KNOW?

Relative Atomic Mass is only a ratio, so it has no unit. If the atomic mass of an element is expressed in grams, it is called as **Gram Atomic Mass**

Gram Atomic Mass of hydrogen = 1 g
Gram Atomic Mass of carbon = 12 g
Gram Atomic Mass of nitrogen = 14 g
Gram Atomic Mass of oxygen = 16 g

7.1.2 Average Atomic Mass (AAM)

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How can one measure the atomic mass of an element? It is somewhat more complicated because most of the naturally occuring elements exist as a mixture of isotopes, each of which has its own mass. Thus, it is essential to consider this isotopic mixture while calculating the atomic mass of an element.

The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.

But, the abundance of isotopes of each element may differ. So, the abundancy of all these isotopes are taken into consideration while calculating the atomic mass. Then, what do we mean by a weighted average? Let us consider an element which exists as a mixture of 50% of an isotope having a mass of 9 amu, and 50% of another isotope having a mass of 10 amu. Then, its average atomic mass is calculated by the following equation:

Average atomic mass

= (Mass of 1st isotope × % abundance of 1st isotope) + (Mass of 2nd isotope × % abundance of 2nd isotope)

Thus, for the given element the average

$$\text{atomic mass} = \left(9 \times \frac{50}{100}\right) + \left(10 \times \frac{50}{100}\right)$$

$$= 4.5 + 5 = 9.5 \text{ amu}$$

(Note: In the calculations involving percentages, you need to convert percentage abundance into fractional abundance. For example, 50 percent is converted into 50/100 or 0.50 as shown in the a foresaid calculation.)

The atomic masses of elements, given in the periodic table, are average atomic masses. Sometimes, the term atomic weight is used to mean average atomic mass. It is observed, from the periodic table that atomic masses of most of the elements are not whole numbers. For instance, the atomic mass of carbon given in the periodic table is 12.01 amu, not 12.00 amu. The reason is that while calculating the atomic mass of carbon, both of its natural isotopes such as carbon-12. and carbon-13 are considered. The natural abundance of C-12 and C-13 are 98.90 % and 1.10 % respectively. The average of the atomic mass of carbon is calculated as follows:

Average atomic mass of carbon

$$= \left(12 \times \frac{98.9}{100}\right) + \left(13 \times \frac{1.1}{100}\right)$$

$$= (12 \times 0.989) + (13 \times 0.011)$$

$$= 11.868 + 0.143 = 12.011 \text{ amu}$$

So it is important to understand that if it is mentioned that the atomic mass of carbon is 12 amu, it refers to the average atomic mass of the carbon isotopes, not the mass of the individual atoms of carbon.

Table 7.2 Atomic mass of some elements

Atomic Number	Name	Symbol	Atomic Mass (amu)
1	Hydrogen	H	1.008
2	Helium	He	4.003
3	Lithium	Li	6.941
4	Beryllium	Be	9.012
5	Boron	B	10.811

Calculation of average atomic mass – Solved Examples

Example 1: Oxygen is the most abundant element in both the Earth's crust and the human body. It exists as a mixture of three stable isotopes in nature as shown in Table 7.3:

Table 7.3 Isotopes of oxygen

Isotope	Mass (amu)	% abundance
$_8O^{16}$	15.9949	99.757
$_8O^{17}$	16.9991	0.038
$_8O^{18}$	17.9992	0.205

The atomic mass of

$$\text{oxygen} = (15.9949 \times 0.99757) + (16.9991 \times 0.00038) + (17.9992 \times 0.00205)$$
$$= 15.999 \text{ amu.}$$

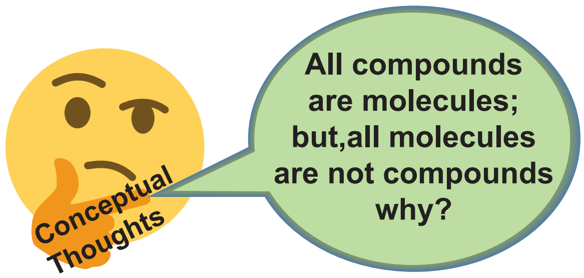
Example 2: Boron naturally occurs as a mixture of boron-10 (5 protons + 5 neutrons) and boron-11 (5 protons + 6 neutrons) isotopes. The percentage abundance of B-10 is 20 and that of B-11 is 80. Then, the atomic mass of boron is calculated as follows:

Atomic mass of

$$\text{boron} = (10 \times \frac{20}{100}) + (11 \times \frac{80}{100})$$
$$= (10 \times 0.20) + (11 \times 0.80)$$
$$= 2 + 8.8$$
$$= 10.8 \text{ amu}$$

7.2 MOLECULE AND MOLECULAR MASS

Except noble gases, atoms of most of the elements are found in the combined form with itself or atoms of other elements. It is called as a molecule. **A molecule is a combination of two or more atoms held together by strong chemical forces of attraction, i.e. chemical bonds.**



7.2.1 Classification of molecules

A molecule may contain atoms of the same element or may contain atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions. Thus, a molecule may be an **element or a compound**. If the molecule is made of similar kind of atoms, then it is called **homoatomic molecule**.

The molecule that consist of atoms of different elements is called **heteroatomic molecule.** A compound is a heteroatomic molecule. The number of atoms present in the molecule is called its '**atomicity**'.

Table 7.4 Classification of molecules

Atomicity	No. of atoms present	Name
1	1	Monoatomic
2	2	Diatomic
3	3	Triatomic
>3	>3	Polyatomic

Let us consider oxygen. Oxygen gas exists in two allotropic forms: Oxygen (O_2) and Ozone (O_3). In oxygen molecule, there are two oxygen atoms. So its atomicity is two. Since both the atoms are similar, oxygen (O_2) is a homodiatomic molecule. Other elements

Activity 7.1

Complete the following table by filling the appropriate values /terms

Element	No. of Protons	No. of Neutrons	Mass Number	Stable Isotopes (abundance)	Atomic Mass (amu)
	7			N-14 (99.6 %)	
		8		N-15 (0.4 %)	
Silicon	**14**		**28**	**Si-28 (92.2 %)**	
	14			**Si-29 (4.7 %)**	
		16		**Si-30 (3.1 %)**	
	17			Cl-35 (75 %)	
	17			Cl-37 (25 %)	

that exist as diatomic molecules are hydrogen (H_2), nitrogen (N_2) and halogens: fluorine (F_2), chlorine (Cl_2), bromine (Br_2) and iodine (I_2).

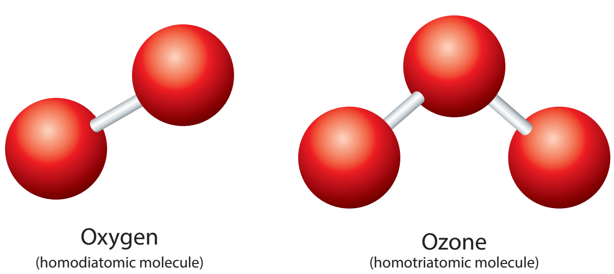


Figure 7.1 Homoatomic molecules

Ozone (O_3) contains three oxygen atoms and hence it is called homotriatomic molecule. If a molecule contains more than three atoms, then it is called **polyatomic molecule**.

Conisder hydrogen chloride. It consists of two atoms, but of different elements, i.e. hydrogen and chlorine. So, its atomicity is two. It is a heterodiatomic molecule. Similarly, the water molecule contains two hydrogen atoms and one oxygen atom. So its atomicity is three. It is a heterotriatomic molecule.

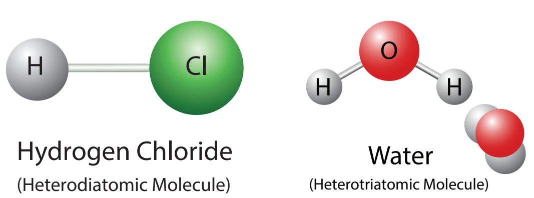


Figure 7.2 Heteroatomic molecules

Activity 7.2

Classify the following molecules based on their atomicity and fill in the table:

Fluorine (F_2), Carbon dioxide (CO_2), Phosphorous (P_4), Sulphur (S_8), Ammonia (NH_3), Hydrogen iodide (HI), Sulphuric Acid (H_2SO_4), Methane (CH_4), Glucose ($C_6H_{12}O_6$), Carbon monoxide (CO)

Molecule	Di - atomic	Tri - atomic	Poly - atomic
Homo			
Hetero			

7.2.2 Relative Molecular Mass (RMM)

As the molecules are made of atoms, they also have their own mass. The mass of the molecule of an element or compound is measured in the C-12 scale and hence called relative molecular mass.

The Relative Molecular Mass of a molecule is the ratio between the mass of one molecule of the substance to $\frac{1}{12^{th}}$ mass of an atom of Carbon -12.

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DO YOU KNOW?

Relative Molecular Mass is only a ratio. So, it has no unit. If the molecular mass of a compound is expressed in grams, it is called ***Gram Molecular Mass.***

Gram Molecular Mass of water = 18 g
Gram Molecular Mass of carbon dioxide = 44 g
Gram Molecular Mass of ammonia = 17 g
Gram Molecular Mass of HCl = 36.5 g

The relative molecular mass is obtained by adding together the relative atomic masses of all the atoms present in a molecule.

Calculation of relative molecular mass – Solved examples:

Example 1: Relative molecular mass of sulphuric acid (H_2SO_4) is calculated as follows: Sulphuric acid conatins 2 atoms of hydrogen, 1 atom of sulphur and 4 atoms of oxygen.

Therefore, Relative molecular mass of sulphuric acid = (2 × mass of hydrogen) + (1 × mass of sulphur) + (4 × mass of oxygen)

$= (2 \times 1) + (1 \times 32) + (4 \times 16)$

$= 98$

i.e., one molecule of H_2SO_4 is 98 times as heavy as $\frac{1}{12^{th}}$ of the mass of a carbon –12.

Example 2: Relative molecular mass of water (H_2O) is calculated as follows: A water molecule is made of 2 atoms of hydrogen and one atom of oxygen.

So, the relative molecular mass of water

= (2 × mass of hydrogen) + (1 × mass of oxygen)

$= (2 \times 1) + (1 \times 16)$

$= 18$

i.e., one molecule of H_2O is 18 times as heavy as $\frac{1}{12^{th}}$ of the mass of a carbon –12.

7.3 DIFFERENCE BETWEEN ATOMS AND MOLECULES

Even though atoms are the basic components of molecules, they differ in many aspects when compared to the molecules. Table 7.5 consolidates the major difference between atoms and molecules.

Table 7.5 Difference between atoms and molecules

Atom	Molecule
An atom is the smallest particle of an element	A molecule is the smallest particle of an element or compound.
Atom does not exist in free state except in noble gas	Molecule exists in a free a state
Except some of noble gas, other atoms are highly reactive	Molecules are less reactive
Atom does not have a chemical bond	Atoms in a molecule are held by chemical bonds

7.4 MOLE CONCEPT

So far we discussed about matters in terms of individual atoms and molecules. Atomic mass units provide a relative scale for the masses of the elements. Since the atoms have such small masses, no usable scale can be devised to weigh them in the calibrated units of atomic mass units. In any real situation, we deal with macroscopic samples containing enormous number of atoms. Therefore, it is convenient to have a special unit to describe a very large number of atoms. The idea of a 'unit' to denote a particular number of objects is not new. For example, **the pair (2 items) and the dozen (12 items)**, are all familiar units. Chemists measure atoms and molecules in 'moles'. So, you can now understand that 'mole' denotes a number of particles.

In the SI system, the ***mole (mol) is the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 g (or 0.012 kg) of the carbon-12 isotope.*** The actual number of atoms in 12 g of carbon-12 is determined experimentally. This is called **Avogadro's Number (N_A),** named after an Italian scientist **Amedeo Avogadro** who proposed its significance. Its value is **6.023×10^{23}**. So one mole of a substance contains **6.023×10^{23}** entities. Thus, 5 moles of oxygen molecules contain $5 \times 6.023 \times 10^{23}$ molecules.

Mole Concept: The study of the collection of particles by using mole as the counting unit, in order to express the mass and volume of such unit particles in a bulk of matter is known as **mole concept**.

The number of moles of a substance can be calculated by various means depending on the data available, as follows:

- Number of moles of molecules.
- Number of moles of atoms.
- Number of moles of a gas (Standard molar volume at STP = 22.4 litre).
- Number of moles of ions.

Note:

STP-Standard Temperature and Pressure(273.15 K,1.00 atm)

Mole of atoms:

One mole of an element contains 6.023×10^{23} atoms and it is equal to its gram atomic mass.

i.e., one mole of oxygen atom contains 6.023×10^{23} atoms of oxygen and its gram atomic mass is 16 g.

Mole of molecules:

One mole of matter contains 6.023×10^{23} molecules and it is equal to its gram molecular mass.

i.e., one mole of oxygen molecule contains 6.023×10^{23} molecules of oxygen and its gram molecular mass is 32 g.

Molar volume:

One mole of any gas occupies 22.4 litre or 22400 ml at S.T.P. This volume is called as molar volume.

Calculation of number of moles by Different modes

Number of moles

= Mass / Atomic Mass

= Mass / Molecular mass

= Number of Atoms / 6.023×10^{23}

= Number of Molecules / 6.023×10^{23}

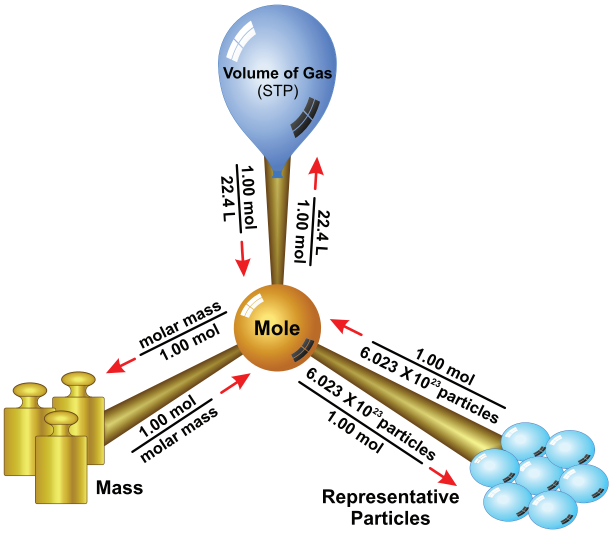

Figure 7.3 Mole concept

7.5 PERCENT COMPOSITION

So for, we were dealing with the number of entities present in a given substance. But many times, the information regarding the percentage of a particular element present in a compound is required.

The percentage composition of a compound represents the mass of each element present in 100 g of the compound.

Let us understand the percentage composition of oxygen and hydrogen by taking the example of H_2O. It can be calculated using the formula

Mass % of an element

$$= \frac{\textit{mass of that element in the compound}}{\textit{molecular mass of the compound}} \times 100$$

Now,

molecular mass of $H_2O = 2(1) + 16$

$= 18$ g

$$\text{Mass \% of hydrogen} = \frac{2}{18} \times 100$$

$= 11.11$ %

$$\text{Mass \% of oxygen} = \frac{16}{18} \times 100$$

$= 88.89$ %

This percentage composition is useful to determine the empirical formula and molecular formula.

Example 1: Find the mass percentage composition of methane (CH_4).

molecular mass of $CH_4 = 12 + 4$

$= 16$ g

$$\text{Mass \% of carbon} = \frac{12}{16} \times 100$$

$= 75$ %

$$\text{Mass \% of hydrogen} = \frac{4}{16} \times 100$$

$= 25$ %

7.6 AVOGADRO HYPOTHESIS

In 1811 Avogadro framed a hypothesis based on the relationship between the number of molecules present in equal volumes of gases in different conditions.

The Avogadro's law states that ***"equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules"***

It follows that the volume of any given gas must be proportional to the number of molecules in it. If 'V' is the volume and 'n' is the number of molecules of a gas, then Avogadro law is represented, mathematically, as follows:

$V \propto n$

$V = \text{constant} \times n$

Thus, one litre (1 dm^3) of hydrogen contains the same number of molecules as in one litre of oxygen, i.e. the volume of the gas is directly proportional to the number of molecules of the gas.

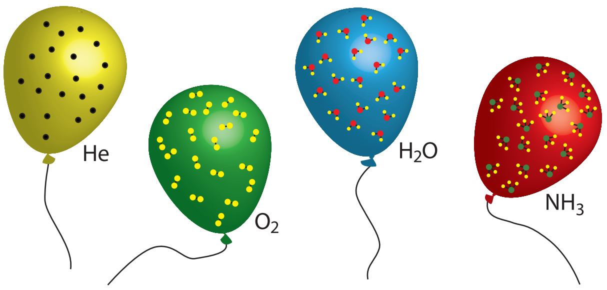

Figure 7.4 Avogadro Hypothesis

Explanation

Let us consider the reaction between hydrogen and chlorine to form hydrogen chloride gas

$H_{2(g)} + Cl_{2(g)} \rightarrow 2\ HCl_{(g)}$

1 vol + 1 vol → 2 volumes

According to Avogadro's law 1 volume of any gas is occupied by "n" number of molecules.

n molecules + n molecules → 2n molecules

if n = 1 then

1molecule + 1 molecule → 2 molecules.

½ molecule + ½ molecule → 1 molecule

1 molecule of hydrogen chloride gas is made up of ½ molecule of hydrogen and ½ molecule of chlorine. Hence, the molecules can be subdivided. This law is in agreement with Dalton's atomic theory.

Activity 7.3

Under same conditions of temperature and pressure if you collect 3 litre of O_2, 5 litre of Cl_2 and 6 litre of H_2,

i. Which has the highest number of molecules?

ii. Which has the lowest number of molecules?

7.7 APPLICATIONS OF AVOGADRO'S LAW

i. It explains Gay-Lussac's law.
ii. It helps in the determination of atomicity of gases.
iii. Molecular formula of gases can be derived using Avogadro's law
iv. It determines the relation between molecular mass and vapour density.
v. It helps to determine gram molar volume of all gases (i.e, 22.4 litre at S.T.P)

7.8 RELATIONSHIP BETWEEN VAPOUR DENSITY AND RELATIVE MOLECULAR MASS

i. Relative molecular mass: (Hydrogen scale)

The Relative Molecular Mass of a gas or vapour is the ratio between the mass of one molecule of the gas or vapour to mass of one atom of Hydrogen.

ii. Vapour Density:

Vapour density is the ratio of the mass of a certain volume of a gas or vapour, to the mass of an equal volume of hydrogen, measured under the same conditions of temperature and pressure.

Vapour Density (V.D.)

$$= \frac{\text{Mass of a given volume of gas or vapour at S.T.P.}}{\text{Mass of the same volume of hydrogen}}$$

According to Avogadro's law, equal volumes of all gases contain equal number of molecules.

Thus, let the number of molecules in one volume = n, then

V.D. at S.T.P

$$= \frac{\text{Mass of 'n' molecules of a gas or vapour at S.T.P.}}{\text{Mass of 'n' molecules of hydrogen}}$$

Cancelling 'n' which is common, you get

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 1 molecules of hydrogen}}$$

However, since hydrogen is diatomic

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 2 atoms of hydrogen}}$$

When you compare the formula of vapour density with relative molecular mass, they can be represented as

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{2 \times \text{Mass of 1 atom of hydrogen}} \quad \text{(Eqn 7.1)}$$

Relative molecular mass (hydrogen scale) =

$$\frac{\text{Mass of 1 molecule of a gas or vapour at STP}}{\text{Mass of 1 atom of hydrogen}} \quad \text{(Eqn 7.2)}$$

You can therefore substitute the above equation to an Eqn 7.1 and arrive at the following formula

$$\text{V.D.} = \frac{\text{Relative molecular mass}}{2}$$

Now on cross multiplication, you have

2 × vapour density = Relative molecular mass of a gas

(Or)

Relative molecular mass = 2 × Vapour density

7.9 SOLVED PROBLEMS

I. Calculation of molecular mass

Calculate the gram molecular mass of the following.

1) H_2O 2) CO_2 3) $Ca_3(PO_4)_2$

Solution:

1) H_2O

Atomic masses of H = 1, O = 16

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Gram molecular mass of H_2O

$= (1 \times 2) + (16 \times 1)$

$= 2 + 16$

Gram molecular mass of H_2O = 18 g

2) CO_2

Atomic masses of C = 12, O = 16

Gram molecular mass of CO_2

$= (12 \times 1) + (16 \times 2)$

$= 12 + 32$

Gram molecular mass of CO_2 = 44 g

3) $Ca_3(PO_4)_2$

Atomic masses of Ca = 40, P = 30, O = 16.

Gram molecular mass of $Ca_3(PO_4)_2$

$= (40 \times 3) + [30 + (16 \times 4)] \times 2$

$= 120 + (94 \times 2)$

$= 120 + 188$

Gram molecular mass of $Ca_3(PO_4)_2$ = 308 g

II. Calculation based on number of moles from mass and volume

1) Calculate the number of moles in 46 g of sodium?

$$\text{Number of moles} = \frac{\text{Mass of the element}}{\text{Atomic mass of the element}}$$

= 46 / 23

= 2 moles of sodium

2) 5.6 litre of oxygen at S.T.P

$$\text{Number moles} = \frac{\text{Given volume of } O_2 \text{ at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$\text{Number of moles of oxygen} = \frac{5.6}{22.4}$$

= 0.25 mole of oxygen

3) Calculate the number of moles of a sample that contains 12.046×10^{23} atoms of iron ?

$$\text{Number of moles} = \frac{\text{Number of atoms of iron}}{\text{Avogadro's number}}$$

$= 12.046 \times 10^{23} / 6.023 \times 10^{23}$

= 2 moles of iron

III. Calculation of mass from mole

Calculate the mass of the following

1) 0.3 mole of aluminium (Atomic mass of Al = 27)

$$\text{Number of moles} = \frac{\text{Mass of Al}}{\text{Atomic mass of Al}}$$

Mass = No. of moles $\times$ atomic mass

So, mass of Al = 0.3×27

= 8.1 g

2) 2.24 litre of SO_2 gas at S.T.P

Molecular mass of $SO_2 = 32 + (16 \times 2)$

$= 32 + 32 = 64$

$$\text{Number of moles of } SO_2 = \frac{\text{Given volume of } SO_2 \text{ at S.T.P}}{\text{Molar volume } SO_2 \text{ at S.T.P}}$$

$$\text{Number of moles of } SO_2 = \frac{2.24}{22.4}$$

= 0.1 mole

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

Mass = No. of moles $\times$ molecular mass

Mass = 0.1×64

Mass of SO_2 = 6.4 g

3) 1.51×10^{23} molecules of water

Molecular mass of H_2O = 18

$$\text{Number of moles} = \frac{\text{Number of molecules of water}}{\text{Avogadro's number}}$$

$= 1.51 \times 10^{23} / 6.023 \times 10^{23}$

= 1 / 4

= 0.25 mole

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

0.25 = mass / 18

Mass = 0.25×18

Mass = 4.5 g

4) 5×10^{23} molecules of glucose ?

Molecular mass of glucose = 180

$$\text{Mass of glucose} = \frac{\text{Molecular mass} \times \text{number of particles}}{\text{Avogadro's number}}$$

$= (180 \times 5 \times 10^{23}) / 6.023 \times 10^{23}$

$= 149.43$ g

IV. Calculation based on number of atoms/ molecules.

1) Calculate the number of molecules in 11.2 litre of CO_2 at S.T.P

$$\text{Number of moles of } CO_2 = \frac{\text{Volume at S.T.P}}{\text{Molar volume}}$$

$= 11.2 / 22.4$

$= 0.5$ mole

Number of molecules of CO_2 = number of moles of CO_2 × Avogadro's number

$= 0.5 \times 6.023 \times 10^{23}$

$= 3.011 \times 10^{23}$ molecules of CO_2

2) Calculate the number of atoms present in 1 gram of gold (Atomic mass of Au = 198)

$$\text{Number of atoms of Au} = \frac{\text{Mass of Au} \times \text{Avogadro's number}}{\text{Atomic mass of Au}}$$

$$\text{Number of atoms of Au} = \frac{1}{198} \times 6.023 \times 10^{23}$$

Number of atoms of Au = 3.042×10^{21} g

3) Calculate the number of molecules in 54 gm of H_2O?

$$\text{Number of molecules} = \frac{(\text{Avogadro number} \times \text{Given mass})}{\text{Gram molecular mass}}$$

Number of molecules

of water = $6.023 \times 10^{23} \times 54 \ / \ 18$

$= 18.069 \times 10^{23}$ molecules

4) Calculate the number of atoms of oxygen and carbon in 5 moles of CO_2.

- 1 mole of CO_2 contains 2 moles of oxygen
- 5 moles of CO_2 contain 10 moles of oxygen

Number of atoms of oxygen = Number of moles of oxygen × Avogadro's number

$= 10 \times 6.023 \times 10^{23}$

$= 6.023 \times 10^{24}$ atoms of Oxygen

- 1 mole of CO_2 contains 1 mole of carbon
- 5 moles of CO_2 contains 5 moles of carbon

No. of atoms of carbon = No.of moles of carbon × Avogadro's number

$= 5 \times 6.023 \times 10^{23}$

$= 3.011 \times 10^{24}$ atoms of Carbon

V. Calculation based on molar volume

Calculate the volume occupied by:

1) 2.5 mole of CO_2 at S.T.P

$$\text{Number of moles of } CO_2 = \frac{\text{Given volume at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$2.5 \text{ mole of } CO_2 = \frac{\text{Volume of } CO_2 \text{ at S.T.P}}{22.4}$$

Volume of CO_2 at S.T.P = 22.4×2.5

= 56 litres.

2) 12.046×10^{23} of ammonia gas molecules

$$\text{Number of moles} = \frac{\text{Number of molecules}}{\text{Avogadro's number}}$$

$= 12.046 \times 10^{23} / 6.023 \times 10^{23}$

= 2 moles

Volume occupied by NH_3

= number of moles × molar volume

$= 2 \times 22.4$

= 44.8 litres at S.T.P

3) 14 g nitrogen gas

Number of moles = 14 / 28

= 0.5 mole

Volume occupied by N_2 at S.T.P

= no. of moles × molar volume

$= 0.5 \times 22.4$

= 11.2 litres.

VI. Calculation based on % composition

Calculate % of S in H_2SO_4

molecular mass of H_2SO_4

$= (1 \times 2) + (32 \times 1) + (16 \times 4)$

$= 2 + 32 + 64$

$= 98$ g

$$\% \text{ of S in } H_2SO_4 = \frac{\text{Mass of sulphur}}{\text{Molecular mass of } H_2SO_4} \times 100$$

$$\% \text{ of S in } H_2SO_4 = \frac{32}{98} \times 100$$

$= 32.65$ %

Points to Remember

- Two or more forms of an element having the same atomic number, but different mass number are called Isotopes ($_{17}Cl^{35}$, $_{17}Cl^{37}$).
- Atoms of different elements having the same mass number, but different atomic numbers are called Isobars ($_{18}Ar^{40}$, $_{20}Ca^{40}$).
- Atoms of different elements having the same number of neutrons, but different atomic number and different mass number are called Isotones ($_6C^{13}$, $_7N^{14}$).
- Relative atomic mass of an element is the ratio between the mass of one atom of the element to 1/12th of the mass of the atom of carbon -12.
- Average atomic mass of an element is calculated by adding the masses of its isotopes, each multiplied by their natural abundance on the Earth.
- Relative molecular mass of a molecule is the ratio between the mass of one molecule of the substance to 1/12th of the mass of the atom of carbon – 12.
- The Avogadro's law states that "equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules".
- The vapour density is defined as "the ratio between the masses of equal volumes of a gas (or a vapour) and hydrogen under the same condition".
- Atomicity of a monoatomic element = Molecular mass / Atomic Mass.
- Molecular mass = 2 × Vapour density.

DLRFYB

TEXTBOOK EVALUATION

I. Choose the best answer.

1. Which of the following has the smallest mass?
   a. 6.023×10^{23} atoms of He
   b. 1 atom of He
   c. 2 g of He
   d. 1 mole atoms of He

2. Which of the following is a triatomic molecule?
   a. Glucose b. Helium
   c. Carbon dioxide d. Hydrogen

3. The volume occupied by 4.4 g of CO_2 at S.T.P
   a. 22.4 litre b. 2.24 litre
   c. 0.24 litre d. 0.1 litre

4. Mass of 1 mole of Nitrogen atom is
   a. 28 amu b. 14 amu
   c. 28 g d. 14 g

5. Which of the following represents 1 amu?
   a. Mass of a C – 12 atom
   b. Mass of a hydrogen atom
   c. $1/12^{th}$ of the mass of a C – 12 atom
   d. Mass of O – 16 atom

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6. Which of the following statement is incorrect?
   a. 12 gram of C – 12 contains Avogadro's number of atoms.
   b. One mole of oxygen gas contains Avogadro's number of molecules.
   c. One mole of hydrogen gas contains Avogadro's number of atoms.
   d. One mole of electrons stands for 6.023×10^{23} electrons.

7. The volume occupied by 1 mole of a diatomic gas at S.T.P is
   a. 11.2 litre b. 5.6 litre
   c. 22.4 litre d. 44.8 litre

8. In the nucleus of ${}_{20}Ca^{40}$, there are
   a. 20 protons and 40 neutrons
   b. 20 protons and 20 neutrons
   c. 20 protons and 40 electrons
   d. 40 protons and 20 electrons

9. The gram molecular mass of oxygen molecule is
   a. 16 g b. 18 g
   c. 32 g d. 17 g

10. 1 mole of any substance contains _____ molecules.
    a. 6.023×10^{23} b. 6.023×10^{-23}
    c. 3.0115×10^{23} d. 12.046×10^{23}

II. Fill in the blanks

1. Atoms of different elements having ________ mass number, but _________ atomic numbers are called isobars.
2. Atoms of different elements having same number of ____________ are called isotones.
3. Atoms of one element can be transmuted into atoms of other element by _________
4. The sum of the numbers of protons and neutrons of an atom is called its __________
5. Relative atomic mass is otherwise known as ___________
6. The average atomic mass of hydrogen is ____________ amu.
7. If a molecule is made of similar kind of atoms, then it is called _____________ atomic molecule.
8. The number of atoms present in a molecule is called its _____________
9. One mole of any gas occupies _________ ml at S.T.P
10 Atomicity of phosphorous is ____________

III. Match the following

1.	8 g of O_2	-	4 moles
2.	4 g of H_2	-	0.25 moles
3.	52 g of He	-	2 moles
4.	112 g of N_2	-	0.5 moles
5.	35.5 g of Cl_2	-	13 moles

IV. True or False: (If false give the correct statement)

1. Two elements sometimes can form more than one compound.
2. Noble gases are Diatomic
3. The gram atomic mass of an element has no unit
4. 1 mole of Gold and Silver contain same number of atoms
5. Molar mass of CO_2 is 42g.

V. Assertion and Reason:

Answer the following questions using the data given below:

i) A and R are correct, R explains the A.

ii) A is correct, R is wrong.

iii) A is wrong, R is correct.

iv) A and R are correct, R doesn't explains A.

1. Assertion: The Relative Atomic mass of aluminium is 27

Reason: An atom of aluminium is 27 times heavier than 1/12th of the mass of the C – 12 atom.

2. **Assertion:** The Relative Molecular Mass of Chlorine is 35.5 a.m.u.

   Reason: The natural abundance of Chlorine isotopes are not equal.

VI. Short answer questions

1. Define: Relative atomic mass.
2. Write the different types of isotopes of oxygen and its percentage abundance.
3. Define: Atomicity
4. Give any two examples for heterodiatomic molecules.
5. What is Molar volume of a gas?
6. Find the percentage of nitrogen in ammonia.

VII. Long answer questions

1. Calculate the number of water molecule present in one drop of water which weighs 0.18 g.

2. $N_2 + 3\ H_2 \rightarrow 2\ NH_3$

   (The atomic mass of nitrogen is 14, and that of hydrogen is 1)

   1 mole of nitrogen (________g) +
   3 moles of hydrogen (__________ g) →

   2 moles of ammonia (__________ g)

3. Calculate the number of moles in

   i) 27g of Al   ii) 1.51×10^{23} molecules of NH_4Cl

4. Give the salient features of "Modern atomic theory".
5. Derive the relationship between Relative molecular mass and Vapour density.

VIII. HOT question

1. Calcium carbonate is decomposed on heating in the following reaction

$$CaCO_3 \rightarrow CaO + CO_2$$

   i. How many moles of Calcium carbonate are involved in this reaction?

   ii. Calculate the gram molecular mass of calcium carbonate involved in this reaction

   iii. How many moles of CO_2 are there in this equation?

IX. Solve the following problems

1. How many grams are there in the following?

   i. 2 moles of hydrogen molecule, H_2

   ii. 3 moles of chlorine molecule, Cl_2

   iii. 5 moles of sulphur molecule, S_8

   iv. 4 moles of phosphorous molecule, P_4

2. Calculate the % of each element in calcium carbonate. (Atomic mass: C-12, O-16, Ca -40)

3. Calculate the % of oxygen in $Al_2(SO_4)_3$. (Atomic mass: Al-27, O-16, S -32)

4. Calculate the % relative abundance of B -10 and B -11, if its average atomic mass is 10.804 amu.

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2) Raymond Chang. (2010). Chemistry. New York, NY: The Tata McGraw Hill Companies. Inc.

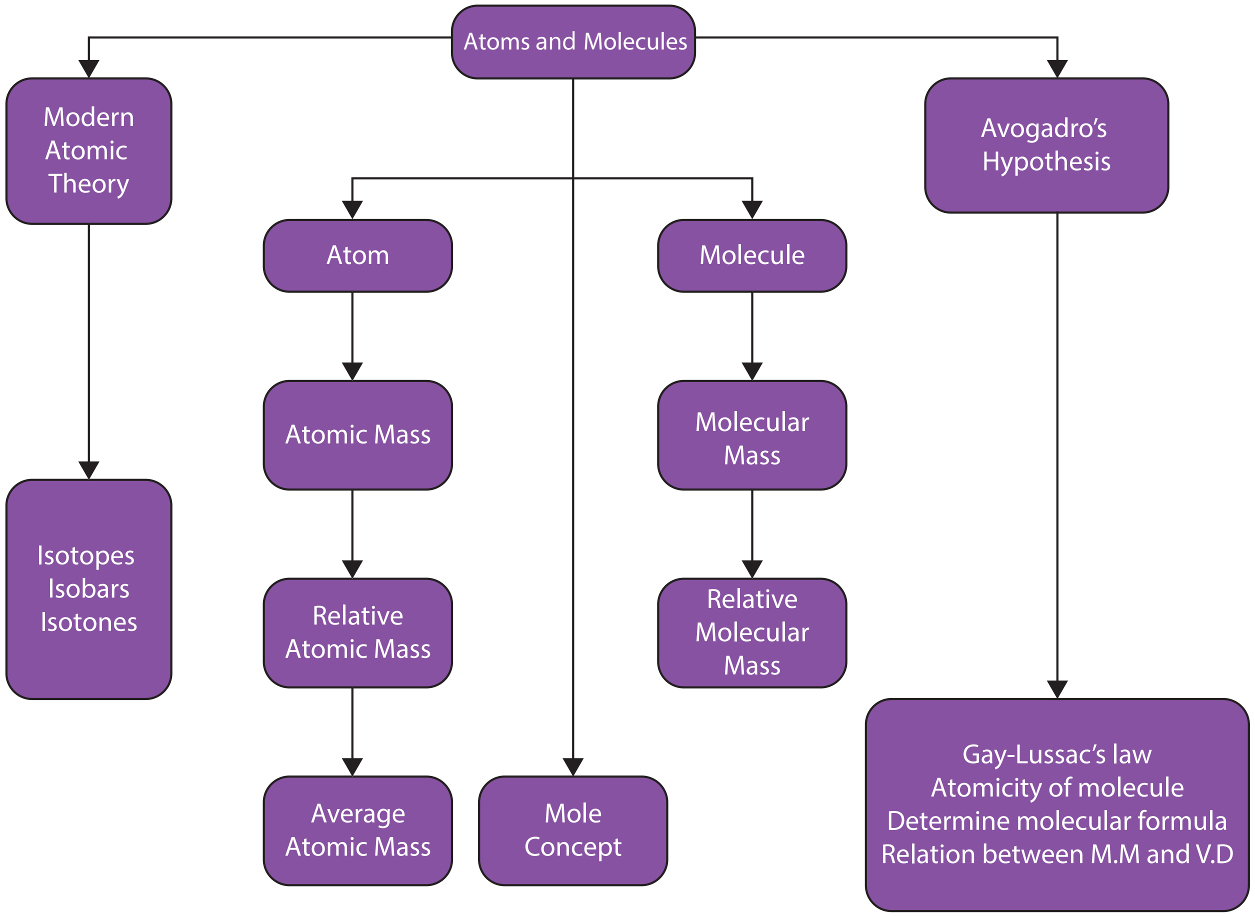
3) Julia Burdge. (2011). Chemistry. New York, NY: The Tata McGraw Hill Companies.Inc.

INTERNET RESOURCES

https://www2.estrellamountain.edu/faculty/farabee/biobk/BioBookCHEM1.html

https://www.toppr.com/guides/chemistry/atoms-and-molecules/

CONCEPT MAP



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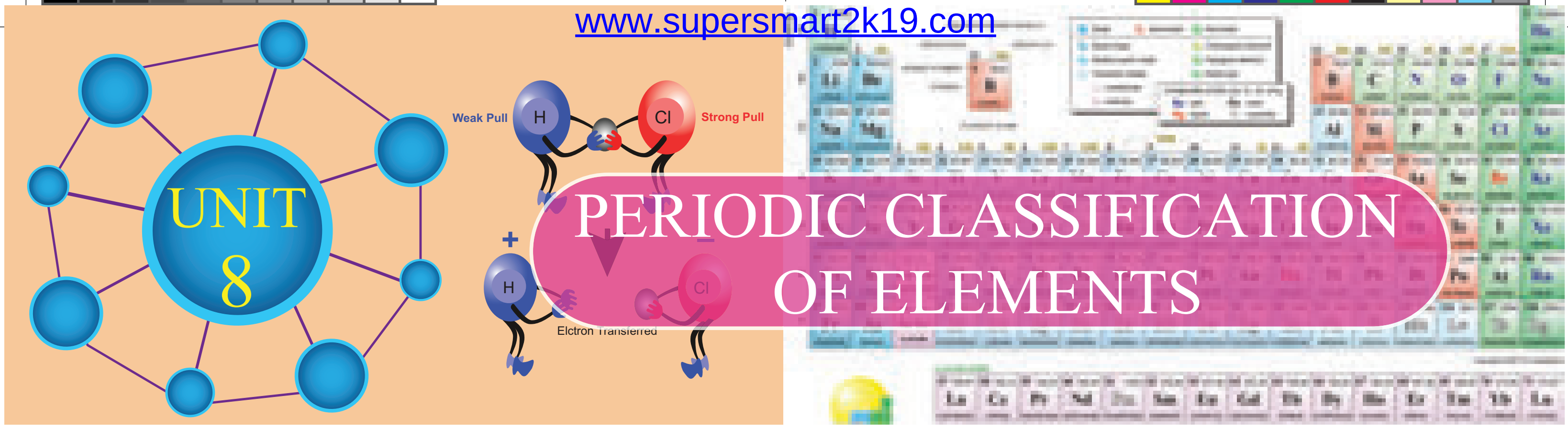


UNIT 8

PERIODIC CLASSIFICATION OF ELEMENTS

E2AAZM

Learning Objectives

After a thorough perusal of this unit, the students will be able to:

- recognize the basis of the modern periodic law and its development.
- list the features of groups and periods of the modern periodic table.
- explain the trend in periodic properties along the periods and groups.
- distinguish between ores and minerals .
- list out the types of separation of impurities from the ores.
- recall the various places of occurrences of minerals in the state of Tamil Nadu.
- put forth the properties of metals.
- identify the stages involved in metallurgical processes.
- think scientifically on alloys and their types.
- develop an idea on amalgam.
- understand the reason for corrosion and the methods of its prevention.

INTRODUCTION

The eighteenth and nineteenth centuries witnessed a rapid development in chemistry in all spheres of scientific activities. By 1860, scientists had already discovered 60 elements and determined their atomic masses. They noticed that some elements had similar properties and hence arranged them into groups. During this period, several new elements were discovered. These elements were found to have different properties. It was realized that instead of studying the properties of all these elements individually, it would be more convenient to divide them into groups and periods in such a way that each group contained a certain number of elements **(like an array of fruits and vegetables showing orderliness)** with similar properties and periods showing a regular gradation. So, scientists made several attempts to arrange elements in a logical way. You have studied about all these early attempts of arrangement of elements in standard IX. In continuation of the knowledge gained in the topic **periodic classification of elements** in standard IX with earlier concepts and their subsequent deliberations, you get set to go ahead with the higher order of thinking to enhance your knowledge on the properties of elements.

8.1 MODERN PERIODIC LAW

Mendeleev's periodic table had some discrepancies, which were difficult to overcome. For example, the atomic mass of argon (39.95 amu) is greater than that of potassium (39.10 amu), but argon comes before potassium in the periodic table. If elements were arranged solely according to increasing atomic mass, argon would appear in the position occupied by potassium in our modern periodic table (see in Figure 8.1). No chemist would place argon, a gas with no tendency to react, in the same group as lithium and sodium,which are two highly reactive metals. This kind of discrepancies suggested that some fundamental property other than atomic mass must be the basis of periodicity. The fundamental property turned out to be the number of protons in an atom's nucleus, something that could not have been known by Mendeleev and his contemporaries.

Henry Moseley, a British scientist in 1912, discovered a new property of elements called atomic number, which provided a better basis for the periodic arrangement of the elements. It is a well-known fact that atomic number of an element is equal to the number of protons or the number of electrons present in the neutral atom of an element. The periodic law was, therefore, modified to frame a **modern periodic law,** which states that

"The physical and chemical properties of the elements are the periodic functions of their atomic numbers".

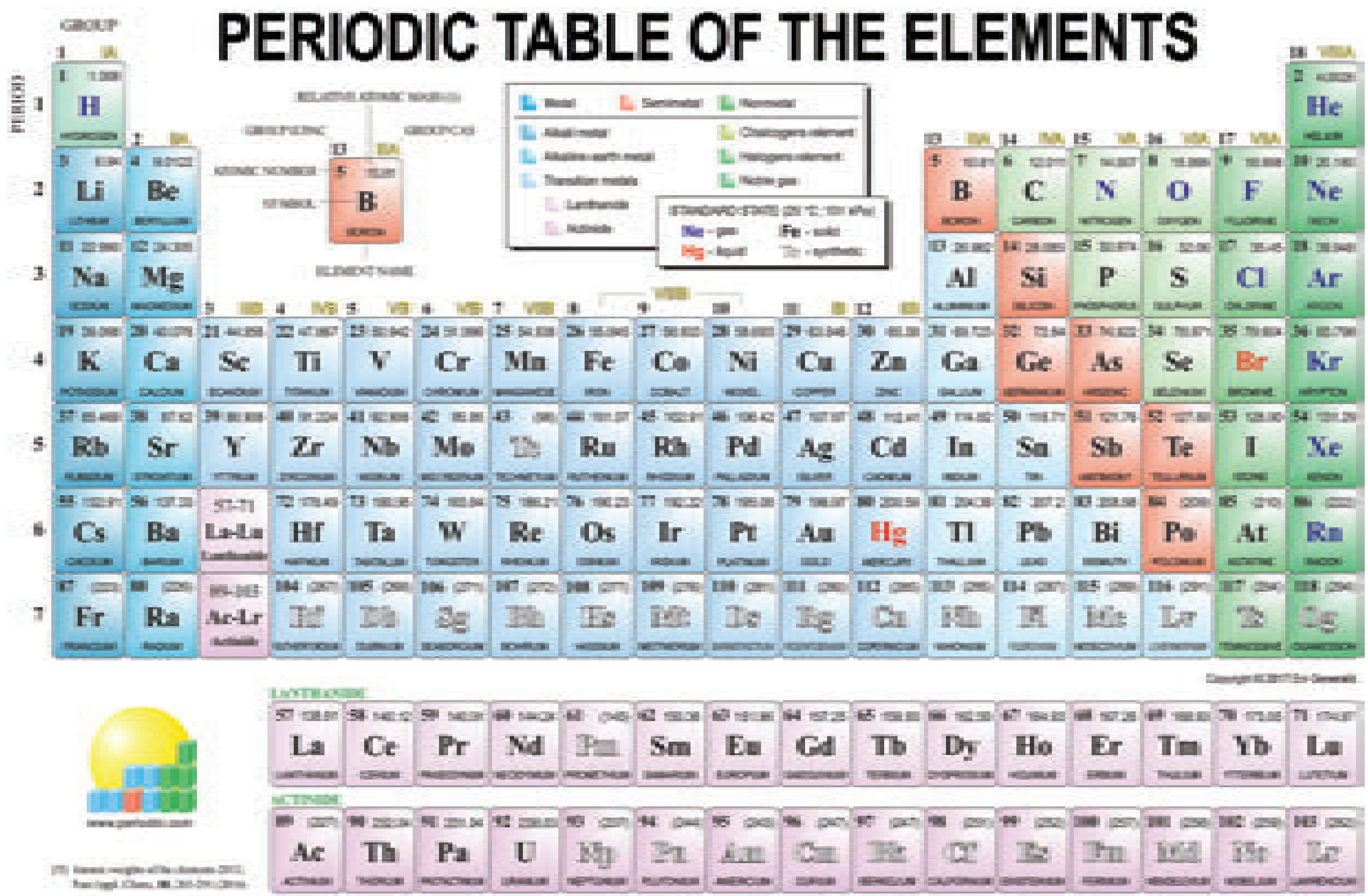
8.2 MODERN PERIODIC TABLE

With reference to the modern periodic law, the elements were arranged in the increasing order of their atomic numbers to form the modern periodic table. **The modern periodic table is a tabular arrangement of elements in periods and groups, highlighting the regular repetition of properties of the elements.** Figure 8.1 shows the modern periodic table of 118 elements discovered so far.

As you have studied the features of the modern periodic table in standard IX, here let us confine to the study of the features of periods and groups.

8.2.1 Features of Periods

- The **horizontal rows are called periods**. There are **seven** periods in the periodic table.
- **First period** (Atomic number 1 and 2): This is the shortest period. It contains only two elements (Hydrogen and Helium).
- **Second period** (Atomic number 3 to 10): This is a short period. It contains eight elements (Lithium to Neon).
- **Third period** (Atomic number 11 to 18): This is also a short period. It contains eight elements (Sodium to Argon).
- **Fourth period** (Atomic number 19 to 36): This is a long period. It contains eighteen elements (Potassium to Krypton). This includes 8 representative elements and 10 transition elements.
- **Fifth period** (Atomic number 37 to 54): This is also a long period. It contains 18 elements (Rubidium to Xenon). This includes 8 representative elements and 10 transition elements.
- **Sixth period** (Atomic number 55 to 86): This is the longest period. It contains 32 elements (Caesium to Radon). This includes 8 representative elements, 10 transition elements and 14 inner transition elements (Lanthanides).
- **Seventh period** (Atomic number 87 to 118): Like the sixth period, this period also accommodates 32 elements. Recently 4 elements have been included by IUPAC.



*(CAS) Chemical Abstract Service | *(IUPAC) International union of Pure and Applied Chemistry

Figure 8.1 Modern Periodic Table

8.2.2 Features of Groups

- The vertical columns in the periodic table starting from top to bottom are called **groups.** There are **18 groups** in the periodic table.
- Based on the common characteristics of elements in each group, they can be grouped as various families.

Group Number	Family
1	Alkali Metals
2	Alkaline earth metals
3 to 12	Transition metals
13	Boron Family
14	Carbon Family
15	Nitrogen Family
16	Oxygen Family (or) Chalcogen family
17	Halogens
18	Noble gases

- The Lanthanides and Actinides, which form part of Group 3 are called **inner transition elements**.
- Except 'group 18', all the elements present in each group have the same number of electrons in their valence shell and thus have the same valency. For example, all the elements of group 1 have one electron in their valence shells ($1s^1$). So, the valency of all the alkali metals is '1'.
- As the elements present in a group have identical valence shell electronic configurations, they possess similar chemical properties.
- The physical properties of the elements in a group such as melting point, boiling point and density vary gradually.
- The atoms of the 'group 18' elements have stable electronic configuration in their valence shells and hence they are unreactive.

8.3 PERIODIC TRENDS IN PROPERTIES

EGJ9ZB

The electronic configurations of elements help us to explain the periodic recurrence of physical and chemical properties. Anything which repeats itself after a regular interval is called **periodic** and this behaviour is called **periodicity.** Some of the atomic properties of the elements are periodic.

Properties such as atomic radius, ionic radius, ionisation energy, electronegativity, electron affinity, show a regular periodicity and hence they are called **periodic properties**. The main significance of the modern periodic table is that it gives a clear understanding of the general properties and trends within a group or a period to predict with considerable accuracy, the properties of any element, even though that element may be unfamiliar to us. Let us discuss the periodic trend of some of the properties.

8.3.1 Atomic Radius

Atomic radius of an atom is defined as the distance between the centre of its nucleus and the outermost shell containing the valence electron. Direct measurement of the radius of an isolated atom is not possible. Except for noble gases, usually the atomic radius is referred to as **covalent radius** or **metallic radius** depending on the nature of the bonding between the concerned atoms. Atomic radius in metal atoms is known as **metallic radius.** It is defined as **half the distance between the nuclei of adjacent metal atoms** (Figure 8.2

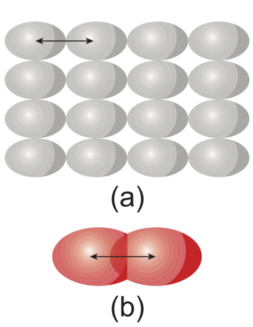


Figure 8.2
(a) Metallic Radius
(b) Covalent Radius

(a)). In non-metallic elements, their atomic radius is known as **Covalent radius.** It is defined as **half the distance between the adjacent nuclei of two covalently bonded atoms of the same element in a molecule** (Figure 8.2 (b)). For example, let us consider H_2 molecule. The distance between the two hydrogen nuclei of the molecule is 0.74 Å. So its covalent radius is 0.74/2 = 0.37 Å.

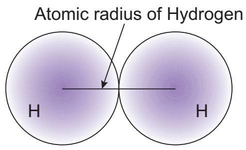


Figure 8.3 Atomic radius

When you look at the variation of the atomic radii in the periodic table, there are two distinct trends. Along the period, from left to right, the atomic radius of the elements decreases whereas along the groups, from the top to bottom, the atomic radius increases. The increase, down a group, is due to the increase in the valence shell number down the group. As the shell number increases, the distance between the valence shell and the nucleus increases. In contrast, when you observe along the period, the shell number remains the same but the number of protons (i.e. atomic number) increases. More and more positive charges impose a strong attraction over the electrons and thus the electron cloud shrinks towards the nucleus, which results in the decrease in the atomic size. Figure 8.4 shows how the atomic radius decreases from lithium to boron.

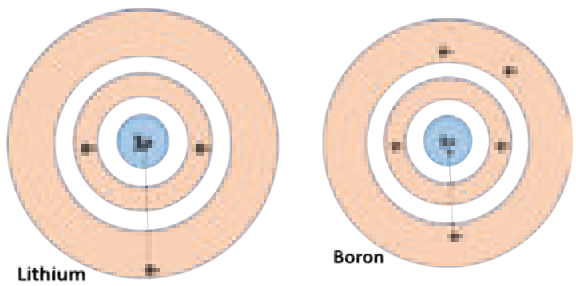


Figure 8.4 Variation of atomic radius

8.3.2 Ionic Radii

It is defined as the distance from the centre of the nucleus of the ion upto the point where it exerts its influence on the electron cloud of the ion. You know that ions are formed when an atom lose or gain electrons. When a neutral atom loses an electron, it becomes a positively charged ion called **cation,** whereas the gain of an electron by a neutral atom forms a negatively charged ion called **anion.** The size of the ions is important to determine their behaviours in solutions and the structure of ionic solids. The size of a cation is always smaller than its corresponding neutral atom. But, the anion is larger than its neutral atom.

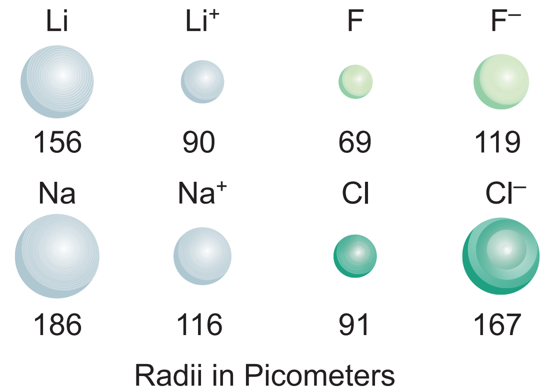


Figure 8.5 Relative ionic radii of cation and anion

Note: As the positive charge increases the size of the cation decreases

As the negative charge increases the size of the anion increases

For instance, lithium and sodium lose the single electron from their outermost energy level to form cations. The ions so formed are smaller because the remaining electrons are at a inner cells and attracted more strongly by the nucleus. Fluorine and chlorine become negative ions by gaining an electron. When electrons are added, the charge on the nucleus is not great enough to hold the increased number of electrons as closely as it holds the electrons in the neutral atom. So, **as seen in atomic radius, ionic radii also decrease along the period from left to right and increase down the group.**

8.3.3 Ionisation Energy

Ionisation energy is the minimum energy required to remove an electron from an isolated gaseous atom in its ground state to form a cation. It is otherwise called **ionisation enthalpy.** It is measured in kJ/mol. Higher the ionisation energy, it is more difficult to remove the electron.

As the atomic size decreases from left to right in a period, more energy is required to remove the electrons. **So, the ionisation energy increases along the period.** But, down the group, the atomic size increases and hence the valence electrons are loosely bound. They require relatively less energy for the removal. Thus, **ionisation energy decreases down the group in the periodic table.**

8.3.4 Electron Affinity

Electron affinity is the amount of energy released when an isolated gaseous atom gains an electron to form its anion. It is also measured in kJ/mol and represented by the following equation:

$$A_{(g)} + e^- \rightarrow A^-_{(g)} + \text{Energy}$$
$$Cl_{(g)} + e^- \rightarrow Cl^-_{(g)} + \text{energy}$$

Like ionisation energy, electron affinity also increases from left to right in a period and decreases from top to bottom in a group.

More to Know

Noble gases show no tendency to accept electrons because the outer s and p orbitals of noble gases are completely filled. No more electrons can be added to them and hence their electron affinities are zero.

8.3.5 Electronegativity

Electronegativity of an element is the measure of the tendency of its atom to attract the shared pair of electrons towards itself in a covalent bond. Let us consider HCl molecule. Both the hydrogen and chlorine atoms share one electron each to form the covalent bond between them. chlorine atom has a higher electronegativity and hence it pulls the shared electrons towards itself more strongly than hydrogen. Thus, when the bond breaks, the bonding electrons are left with chlorine forming H^+ and Cl^- ions. It is represented, diagrammatically, as shown below:

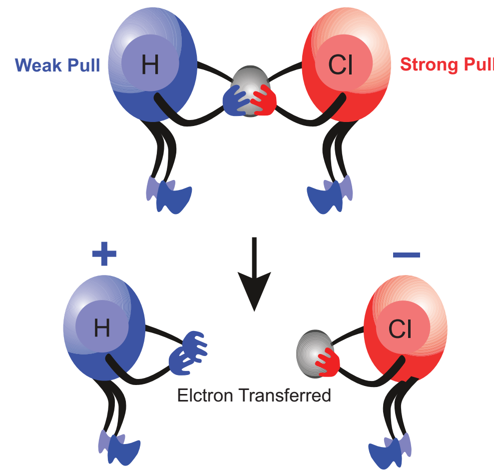


Figure 8.6 Relative electronegativity of H and Cl

Electronegativity is based on various experimental data such as bond energy, ionization potential, electron affinity, etc.

Pauling scale is the widely used scale to determine the electronegativity, which in turn predicts the nature of bonding (ionic or covalent) between the atoms in a molecule.

Electronegativity of some of the elements are given below

F = 4.0, Cl = 3.0, Br = 2.8, I = 2.5, H = 2.1, Na = 1

If the difference in electronegativity between two elements is 1.7, the bond has 50% ionic character and 50% covalent character.

If the difference is less than 1.7, the bond is considered to be more covalent.

If the difference is greater than 1.7, the bond is considered to be more ionic.

Along the period, from left to right in the periodic table, the electronegativity increases because of the increase in the nuclear charge which in turn attracts the electrons more strongly. On moving down a group, the electronegativity of the elements decreases because of the increased number of valence shells.

Periodic Property	In Periods	In Groups
Atomic radius	Decreases	Increases
Ionic radius	Decreases	Increases
Ionisation energy	Increases	Decreases
Electron affinity	Increases	Decreases
Electronegativity	Increases	Decreases

Test yourself

Predict the nature of the bond in the following molecules.

(i) NaCl (ii) NaBr (iii) NaI

(iv) NaF (v) NaH

8.4 METALLURGY

Human life is associated with various metals. We use metals in our day to day activities. It is the utmost need to have some metals like sodium, potassium, calcium, iron, etc. in the human body. Deficiency of these metals affects the metabolic activities thereby causing diseases. So, metals play a vital role in our life. In this section, let us discuss how metals are obtained from various sources by the process of metallurgy.

EW351M

Metallurgy is a science of extracting metals from their ores and modifying the metals into alloys for various uses, based on their physical and chemical properties and their structural arrangement of atoms. A metallurgical process involve three main steps as follows:

(i) **Concentration or Separation of the ore:** It is the process of removal of impuries from the ore.

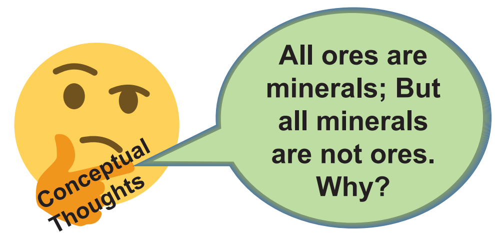
(ii) **Production of the metal:** It is the convertion of the ore into metal.

(iii) **Refining of the metal:** It is the process of purification of the metal.

8.4.1 Terminology in metallurgy

Minerals: A mineral may be a single compound or a complex mixture of various compounds of metals found in the Earth.

Ore: The mineral from which a metal can be readily and economically extracted on a large scale is said to be an ore.



For example: Clay (Al_2O_3. 2 SiO_2. 2 H_2O) and bauxite (Al_2O_3.2 H_2O) are the two minerals of aluminium, but aluminium can be profitably extracted only from bauxite. Hence, bauxite is an ore of aluminium and clay is its mineral.

Mining: The process of extracting the ores from the Earth's crust is called mining.

Gangue or Matrix: The rocky impurity associated with an ore is called gangue or matrix.

Flux: It is the substance added to the ore to reduce the fusion temperature and to remove the impurities. E.g. Calcium oxide (basic), Silica (acidic). If the gangue is acidic, then basic flux is added and vice versa.

Slag: It is the fusible product formed when a flux reacts with a gangue during the extraction of metals.

$$\text{Flux} + \text{Gangue} \rightarrow \text{Slag}$$

Smelting: Smelting is the process of reducing the roasted metallic oxide from the metal in its molten condition. In this process, impurities are removed as slag by the addition of flux.

8.4.2 Types of separation or concentration of an ore

There are four major types of separation of ores based on the nature of the ore. The different kinds of ores of metals are given in Table 8.1

Concentration of the crushed ore is done mainly by the following methods: -

(i) **Hydraulic (Gravity Separation) method**

Principle: The difference in the densities or specific gravities of the ore and the gangue is the main principle behind this method. Oxide ores are purified by this method. e.g., Haematite Fe_2O_3 the ore of iron.

Note: When the ore is heavier than the impurity, this method can be used.

Method: The ore is poured over a sloping, vibrating corrugated table with grooves and a jet of water is allowed to flow over it. The denser ore particles settle down in the grooves and lighter gangue particles are washed down by water.

(ii) **Magnetic separation method**

Principle: The magnetic properties of the ores form the basis of separation. When either the ore or the gangue is magnetic, this method is employed. e.g., Tinstone SnO_2, the ore of tin.

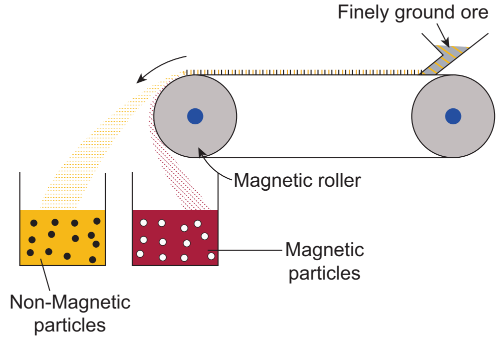


Figure 8.8 Magnetic separation

Method: The crushed ore is placed over a conveyer belt which rotates around two metal wheels, one of which is magnetic. The magnetic particles are attracted to the magnetic wheel and fall separately apart from the non-magnetic particles.

(iii) **Froth floatation**

Principle: This process depends on the preferential wettability of the ore with oil (pine oil) and the gangue particles by water. Lighter ores, such as sulphide ores, are concentrated by this method. e.g., Zinc blende (ZnS).

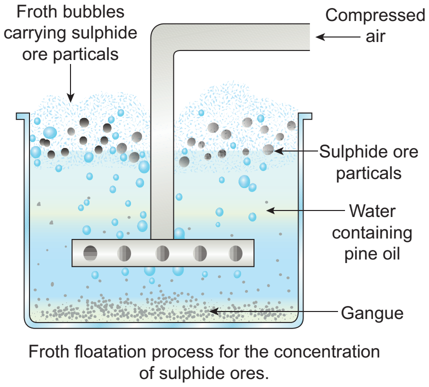


Figure 8.9 Froth floatation

Note: When the impurity is heavier than the ore, this method can be used.

Method: The crushed ore is taken in a large tank containing pine oil and water and agitated with a current of compressed air.The ore is wetted by the oil and gets separated from the gangue in the form of froth. Since the ore is lighter, it comes on the surface with the froth and the impurities are left behind. e.g., Zinc blende (ZnS).

(iv) **Chemical method or Leaching**

This method is employed when the ore is in a very pure form.

Table 8.1 Types of ores

Oxide Ores	Carbonate Ores	Halide Ores	Sulphide Ores
Bauxite ($Al_2O_3 \cdot 2H_2O$)	Marble ($CaCO_3$)	Cryolite(Na_3AlF_6)	Galena (PbS)
Cuprite(Cu_2O)	Magnesite (($MgCO_3$)	Fluorspar(CaF_2)	Iron pyrite (FeS_2)
Haematite (Fe_2O_3)	Siderite($FeCO_{3)}$	Rock salt (NaCl)	Zinc blende (ZnS)

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More to Know

Extraction of metal from metal oxide can be categorized into three types.

More reactive metals	Medium reactive metals	Less reactive metals
Na,K,Ca,Mg,Al	Zn,Fe,Pb,Cu	Ag,Hg
Electrolytic reduction of metal oxide into metal	Chemical reduction of metal oxide into metal using coke	Thermal decomposition of metal oxide into metal

The ore is treated with a suitable reagent such that the ore is soluble in it but the impurities are not. The impurities are removed by filtration. The solution of the ore, ie., the filtrate is treated with a suitable reagent which precipitates the ore. E.g. Bauxite $Al_2O_3.2H_2O$, (the ore of aluminium).

8.5 OCCURRENCE OF ORES IN TAMIL NADU

Lime stone: Coimbatore, Cuddalore, Dindugul

Gypsum: Tiruchi and Coimbatore Distiricts

Titanium minerals: Kanyakumari, Tirunelveli and Tuticorin.

Chromite: Coimbatore and Salem district.

Magnetite:.Dharmapuri, Erode, Salem, Thiruvannamalai.

Tungsten: Madurai and Dindugal.

(Reference: mineral resources of Tamil Nadu-ENVIS Centre, Tamil Nadu)

8.6 PROPERTIES OF METALS

8.6.1 Physical properties

1. **Physical state:** All metals are solids at room temperature except mercury and gallium.
2. **Lustre:** Metals possess a high lustre (called metallic lustre).
3. **Hardness:** Most of the metals are hard and strong (exceptions: sodium and potassium can be cut with a knife)
4. **Melting point and Boiling point:** Usually, metals possess high melting and boiling points and vaporize only at high temperatures (exceptions: gallium, mercury, sodium and potassium).
5. **Density:** Metals have a high density (exceptions: sodium and potassium are less dense than water).
6. **Ductility:** Metals are usually ductile. In other words, they can be drawn into thin wires without breaking.
7. **Malleability:** Metals are usually malleable, i.e, they can be beaten into thin sheets without cracking (except zinc and mercury).
8. **Conduction of heat and electricity:** Metals are good conductors of heat and electricity; silver and copper excel in this property (exception: tungsten)
9. **Solubility:** Usually, metals do not dissolve in liquid solvents.

8.6.2 Chemical Properties

- **Valence electrons:** Atoms of metals usually have 1,2 or 3 electrons in their outermost shell.
- **Formation of ions:** Metals form Positive ions by the loss of electrons and hence they are electro positive.
- **Discharge of ions:** Metals are discharged at the cathode during the electrolysis of their compounds.
- **Atomicity:** Molecules of metals in their vapour state are usually monoatomic.
- **Nature of oxides:** Oxides of metals are usually basic.

8.7 METALLURGY OF ALUMINIUM

Aluminium is the metal found most abundantly in the Earth's crust. Since it is a reactive metal, it occurs in the combined state. The important ores of aluminium are as follows

Ores of Aluminium	Formula
Bauxite	$Al_2O_3 \cdot 2H_2O$
Cryolite	Na_3AlF_6
Corundum	Al_2O_3

Bauxite is the chief ore of aluminium. The extraction of aluminium from bauxite involves two steps:

(i) Conversion of bauxite into alumina – Baeyer's Process

The conversion of Bauxite into Alumina involves the following steps:

Bauxite ore is finely ground and heated under pressure with a solution of concentrated caustic soda solution at 150° C to obtain sodium meta aluminate.

$$Al_2O_3 + 2\ NaOH \rightarrow 2\ NaAlO_2 + H_2O$$

On diluting sodium meta aluminate with water, a precipitate of aluminium hydroxide is formed.

$$NaAlO_2 + 2\ H_2O \rightarrow Al(OH)_3 + NaOH$$

The precipitate is filtered, washed, dried and ignited at 1000°C to get alumina.

$$2Al(OH)_3 \xrightarrow{1000°c} Al_2O_3 + 3H_2O$$

(ii) Electrolytic reduction of alumina – Hall's Process

Aluminium is produced by the electrolytic reduction of fused alumina (Al_2O_3) in the electrolytic cell.

Cathode: Iron tank lined with graphite

Anode: A bunch of graphite rods suspended in molten electrolyte.

Electrolyte: Pure alumina+ molten cryolite + fluorspar (fluorspar lowers the fusion temperature of electrolyte)

Temperature: 900 - 950 °C

Voltage used: 5-6 V

Overall reaction: $2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2\uparrow$

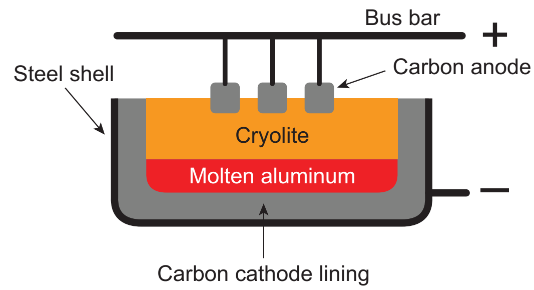


Figure 8.10 Hall's Process

Aluminium is deposited at the cathode and oxygen gas is liberated at the anode. Oxygen combines with graphite to form CO_2.

Physical Properties of Aluminium

- It is a silvery white metal
- It has low density (2.7) and it is light
- It is malleable and ductile
- It is a good conductor of heat and electricity.
- Its melting point is 660 °C.
- It can be polished to produce a shiny attractive appearance.

Chemical Properties of Aluminium

i. Reaction with air: It is not affected by dry air. On heating at 800 °C, aluminium burns very brightly forming it's oxide and nitride.

$$4\ Al + 3\ O_2 \rightarrow 2\ Al_2O_3 \text{(Aluminium oxide)}$$

$$2\ Al + N_2 \rightarrow 2\ AlN \text{ (Aluminium nitride)}$$

ii. Reaction with water: Water does not react with aluminium due to the layer of oxide on it. When steam is passed over red hot aluminium, hydrogen is produced.

$$2\ Al + \underset{\text{(steam)}}{3\ H_2O} \rightarrow \underset{\text{(aluminium oxide)}}{Al_2O_3} + 3\ H_2\uparrow$$

iii. Reaction with alkalis: It reacts with strong caustic alkalis forming aluminates.

$$2\ Al + 2\ NaOH + 2\ H_2O \rightarrow 2\ NaAlO_2 + 3\ H_2\uparrow$$
(Sodium meta aluminate)

iv. Reaction with acids: With dilute and con. HCl it liberates H_2 gas.

$$2\ Al + 6\ HCl \rightarrow 2\ AlCl_3 + 3\ H_2\uparrow$$
(Aluminium chloride)

Aluminium liberates hydrogen on reaction with dilute sulphuric acid and liberates sulphur dioxide on reaction with hot concentrated sulphuric acid

$$2\ Al + 3\ H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3\ H_2\uparrow$$
(Dilute)

$$2\ Al + 6\ H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6\ H_2O + 3\ SO_2\uparrow$$
(Con.)

More to Know

Dilute or concentrated nitric acid does not attack aluminium, but it renders aluminium passive due to the formation of an oxide film on its surface.

v. As reducing agent: Aluminium is a powerful reducing agent. When a mixture of aluminium powder and iron oxide is ignited, the latter is reduced to metal. This process is known as **aluminothermic process**.

$$Fe_2O_3 + 2\ Al \rightarrow 2\ Fe + Al_2O_3 + \text{Heat.}$$

Uses

Aluminium is used in

- household utensils
- electrical cable industry
- making aeroplanes and other industrial mechine parts

8.8 METALLURGY OF COPPER

Occurrence:

It was named as cuprum by the Romans because they got it from the Island of Cyprus. Copper is found in the native state as well as combined state.

Ores of copper	Formula
Copper pyrites	$CuFeS_2$
Cuprite or ruby copper	Cu_2O
Copper glance	Cu_2S

The chief ore of copper is copper pyrite. It yields nearly 76% of the world production of copper. Extraction of copper from copper pyrites involves the following steps

i. Concentration of ore: The ore is crushed and the concentrated by froth floatation process.

ii. Roasting: The concentrated ore is roasted in excess of air. During the process of roasting, the moisture and volatile impurities are removed. Sulphur, phosphorus, arsenic and antimony are removed as oxides. Copper pyrite is partly converted into sulphides of copper and iron.

$$2\ CuFeS_2 + O_2 \longrightarrow Cu_2S + 2\ FeS + SO_2\uparrow$$

iii. Smelting: The roasted ore is mixed with powdered coke and sand and is heated in a blast furnace to obtain matte (Cu_2S + FeS) and slag. The slag is removed as waste.

iv. Bessemerisation: The molten matte is transferred to Bessemer converter in order to obtain blister copper. Ferrous sulphide from matte is oxidized to ferrous oxide, which is removed as slag using silica.

$$2\ FeS + 3\ O_2 \longrightarrow 2\ FeO + 2\ SO_2\uparrow$$

$$FeO + SiO_2 \longrightarrow FeSiO_3\ \text{(slag)}$$
(Iron silicate)

$$2\ Cu_2S + 3O_2 \longrightarrow 2\ Cu_2O + 2\ SO_2\uparrow$$

$$2\ Cu_2O + Cu_2S \longrightarrow 6\ Cu + SO_2\uparrow$$
(Blister copper)

v. Refining: Blister copper contains 98% of pure copper and 2% of impurities and is purified by **electrolytic refining**. This method is used to get metal of a high degree of purity. For electrolytic refining of copper, we use:

Cathode: A thin plate of pure copper metal.

Anode: A block of impure copper metal.

Electrolyte: Copper sulphate solution acidified with sulphuric acid.

When electric current is passed through the electrolytic solution, pure copper gets deposited at the cathode and the impurities settle at the bottom of the anode in the form of sludge called anode mud.

Physical Properties of Copper

Copper is a reddish brown metal, with high lustre, high density and high melting point (1356°C).

Chemical Properties of Copper

i. Action of Air and Moisture: Copper gets covered with a green layer of basic copper carbonate in the presence of CO_2 and moisture.

$$2\,Cu + O_2 + CO_2 + H_2O \rightarrow CuCO_3.Cu(OH)_2$$

ii. Action of Heat: On heating at different temperatures in the presence of oxygen, copper forms two types of oxides CuO and Cu_2O.

$$2\,Cu + O_2 \xrightarrow{\text{below 1370K}} 2\,CuO$$

(copper II oxide- black)

$$4\,Cu + O_2 \xrightarrow{\text{above 1370K}} 2\,Cu_2O$$

(copper I oxide – red)

iii. Action of Acids:

a) With dilute HCl and dilute H_2SO_4:

Dilute acids such as HCl and H_2SO_4 have no action on these metals in the absence of air. Copper dissolves in these acids in the presence of air.

$$2\,Cu + 4\,HCl + O_2\,(air) \longrightarrow 2\,CuCl_2 + 2\,H_2O$$

b) With dil. HNO_3:

Copper reacts with dil. HNO_3 with the liberation of Nitric Oxide gas.

$$3\,Cu + 8\,HNO_3 \rightarrow 3\,Cu(NO_3)_2 + 2\,NO \uparrow + 4H_2O$$

iv) Action of Chlorine:

Chlorine reacts with copper, resulting in the formation of copper(II) chloride.

$$Cu + Cl_2 \longrightarrow CuCl_2$$

v) Action of Alkalis:

Copper is not attacked by alkalis.

Uses of Copper:

i. It is extensively used in manufacturing electric cables and other electric appliances.

ii. It is used for making utensils, containers, calorimeters and coins,

iii. It is used in electroplating.

iv. It is alloyed with gold and silver for making coins and jewels

8.9 METALLURGY OF IRON

Occurrence:

Iron is the second most abundant metal available next to aluminium. It occurs in nature as oxides, sulphides and carbonates. The ores of iron are as follows:

Ores of iron	Formula
Haematite	Fe_2O_3
Magnetite	Fe_3O_4
Iron pyrite	FeS_2

Iron is chiefly extracted from haematite ore (Fe_2O_3)

i. Concentration by Gravity Separation: The powdered ore is washed with a steam of water. As a result, the lighter sand particles and other impurities are washed away and the heavier ore particles settle down.

ii. Roasting and Calcination: The concentrated ore is strongly heated in a limited supply of air in a reverberatory furnace. As a result, moisture is driven out and sulphur, arsenic and phosphorus impurities are oxidized off.

iii. Smelting (in a Blast Furnace): The charge consisting of roasted ore, coke and limestone in the ratio 8:4:1 is smelted in a blast furnace by introducing it through the hopper arrangement at the top. There are three important regions in the furnace.

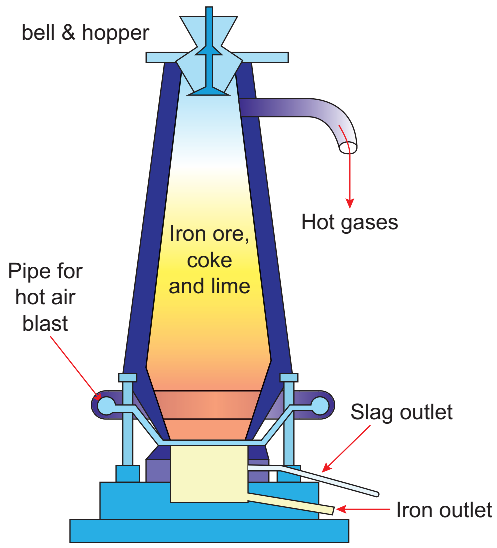


Figure 8.11 Blast Furnace

(a) The Lower Region (Combustion Zone)- The temperature is at 1500°C. In this region, coke burns with oxygen to form CO_2 when the charge comes in contact with a hot blast of air.

$$C + O_2 \xrightarrow[\Delta]{1500°C} CO_2 + Heat$$

It is an exothermic reaction since heat is liberated.

(b) The Middle Region (Fusion Zone) – The temperature prevails at 1000°C. In this region, CO_2 is reduced to CO.

$$CO_2 + C \xrightarrow[\Delta]{1000°C} 2\ CO - Heat$$

Limestone decomposes to calcium oxide and CO_2.

$$CaCO_3 \xrightarrow[\Delta]{1000°C} CaO + CO_2 - Heat$$

These two reactions are endothermic due to absorption of heat. Calcium oxide combines with silica to form calcium silicate slag.

$$CaO + SiO_2 \longrightarrow CaSiO_3$$

(c) The Upper Region (Reduction Zone)- The temperature prevails at 400°C . In this region carbon monoxide reduces ferric oxide to form a fairly pure spongy iron.

$$Fe_2O_3 + 3CO \xrightarrow{400°C} 2Fe + 3CO_2 \uparrow$$

The molten iron is collected at the bottom of the furnace after removing the slag.

The iron thus formed is called pig iron. It is remelted and cast into different moulds. This iron is called cast iron.

Physical properties:

i. It is a lustrous metal, greyish white in colour.

ii. It has high tensility, malleability and ductility.

iii. It can be magnetized.

Chemical properties:

i. **Reaction with air or oxygen:** Only on heating in air, iron forms magnetic oxide.

$$3\ Fe + 2\ O_2 \longrightarrow Fe_3O_4\ (black)$$

ii. **Reaction with moist air:** When iron is exposed to moist air, it forms a layer of brown hydrated ferric oxide on its surface. This compound is known as rust and the phenomenon of formation of rust is known as **rusting**.

$$4\ Fe + 3\ O_2 + x\ H_2O \longrightarrow 2\ Fe_2O_3 \cdot xH_2O(rust)$$

iii. **Reaction with steam:** When steam is passed over red hot iron, magnetic oxide is formed.

$$3\ Fe + 4\ H_2O\ (steam) \longrightarrow Fe_3O_4 + 4\ H_2 \uparrow$$

iv. **Reaction with chlorine:** Iron combines with chlorine to form ferric chloride.

$$2Fe + 3Cl_2 \longrightarrow 2FeCl_3\ (ferric\ chloride)$$

v. **Reaction with acids:** With dilute HCl and dilute H_2SO_4 it liberates H_2 gas.

$$Fe + 2HCl \longrightarrow FeCl_2 + H_2 \uparrow$$

$$Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2 \uparrow$$

With dilute HNO_3 in cold condition it gives ferrous nitrate and ammonium nitrate.

$$4\ Fe + 10\ HNO_3 \rightarrow 4\ Fe(NO_3)_2 + NH_4NO_3 + 3\ H_2O$$

With con. H_2SO_4 it forms ferric sulphate and liberates SO_2.

$$2\,Fe + 6\,H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 3\,SO_2 + 6\,H_2O$$

When iron is dipped in con. HNO_3 it becomes chemically passive or inert due to the formation of a layer of iron oxide (Fe_3O_4) on its surface.

Types and uses of iron

Pig iron (Iron with 2.0% - 4.5% of carbon)**:** It is used in making pipes, stoves, radiators, railings, manhole covers and drain pipes.

Steel (Iron with 0.25% - 2.0% of carbon)**:** It is used in the construction of buildings, machinery, transmission cables and T.V towers and in making alloys.

Wrought iron (Iron with < 0.25% of carbon)**:** It is used in making springs, anchors and electromagnets.

8.10 ALLOYS

An alloy is a homogeneous mixture of two or more metals or of one or more metals with certain non-metallic elements.

The properties of alloys are often different from those of its components. Pure gold is brittle to be used. The addition of small percentage of copper enhances its strength and utility.

8.10.1 Amalgam

An amalgam is an alloy of mercury with another metal. These alloys are formed through metallic bonding with the electrostatic force of attraction between the electrons and the positively charged metal ions. Silver tin amalgam is used for dental filling.

Reasons for alloying:

i. To modify appearance and colour
ii. To modify chemical activity.
iii. To lower the melting point.
iv. To increase hardness and tensile strength.
v. To increase resistance to electricity.

8.10.2 Method of making alloys

(a) By fusing the metals together. E.g. Brass is made by melting zinc and copper.

(b) By compressing finely divided metals. E.g. Wood metal: an alloy of lead, tin, bismuth and cadmium powder is a fusible alloy.

Alloys as solid solutions:

Alloys can be considered as solid solutions in which the metal with high concentration is solvent and other metals are solute.

For example, brass is a solid solution of zinc (solute) in copper (solvent).

8.10.3 Types of Alloys

Based on the presence of Iron, alloys can be classified into:

- **Ferrous alloys:** Contain Iron as a major component. A few examples of ferrous alloys are Stainless Steel, Nickel Steel etc.
- **Non-ferrous alloys:** These alloys do not contain Iron as a major component. For example, Aluminium alloy, Copper alloy etc.

Copper Alloys (Non- ferrous)

Alloys	Uses
Brass (Cu, Zn)	Electrical fittings, medal, decorative items, hardware
Bronze (Cu, Sn)	Statues, coins, bells, gongs

Aluminium Alloys (Non- ferrous)

Alloys	Uses
Duralumin (Al, Mg, Mn, Cu)	Aircrafts, tools, pressure cookers
Magnalium (Al, Mg)	Aircraft, scientific instruments

Iron Alloys(Ferrous)

Alloys	Uses
Stainless steel (Fe,C, Ni,Cr)	Utensils, cutlery, automobile parts
Nickel steel (Fe,C,Ni)	Cables , aircraftparts, propeller

8.11 CORROSION

It is the gradual destruction of metals by chemical or electrochemical reaction with the environment. It is a natural process which converts a metal into its oxide, hydroxide or sulphide so that it loses its metallic characteristics.

Rusting

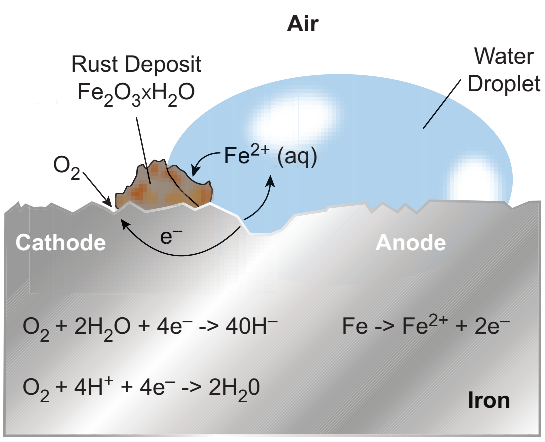


Figure 8.12 Rusting

Rust is chemically known as hydrated ferric oxide (it is formulated as $Fe_2O_3{\cdot}xH_2O$). Rusting results in the formation of scaling reddish brown hydrated ferric oxide on the surface of iron and iron containing materials.

8.11.1 Types of Corrosion

i. **Dry Corrosion or Chemical Corrosion:** The corrosive action in the absence of moisture is called dry corrosion. It is the process of a chemical attack on a metal by a corrosive liquids or gases such as O_2, N_2, SO_2, H_2S etc. It occurs at high temperature. Of all the gases mentioned above O_2 is the most reactive gas to impart the chemical attack.

ii. **Wet Corrosion or Electrochemical Corrosion:** The corrosive action in the presence of moisture is called wet corrosion. It occurs as a result of electrochemical reaction of metal with water or aqueous solution of salt or acids or bases.

8.11.2 Methods of preventing corrosion

i. **Alloying:** The metals can be alloyed to prevent from the process of corrosion. E.g: Stainless Steel

ii. **Surface Coating:** It involves application of a protective coating over the metal. It is of the following types:

a) **Galvanization:** It is the process of coating zinc on iron sheets by using electric current.

b) **Electroplating:** It is a method of coating one metal over another metal by passing electric current.

c) **Anodizing:** It is an electrochemical process that converts the metal surface into a decorative, durable and corrosion resistant. Aluminium is widely used for anodizing process.

d) **Cathodic Protection:** It is the method of controlling corrosion of a metal surface protected is coated with the metal which is easily corrodible. The easily corrodible metal is called Sacrificial metal to act as anode ensuring cathodic protection.

8.12 PAMBAN BRIDGE

It is a railway bridge which connects the town of Rameshwaram on Pamban Island to mainland India. Opened on 1914, it was India's first sea bridge in India until the opening of the BandraWorli Sea Link in 2010. We can control the corrosion and renovation of historical pamban bridge by a periodical protective coating which will be the strong example for applied chemistry to uphold our history.

Figure 8.12 Pamban Bridge

Points to Remember

- ❖ Modern periodic law states that, the physical and chemical properties of the elements are the periodic functions of their atomic numbers.
- ❖ The table in which elements are arranged in rows and columns in regular gradation is called periodic table.
- ❖ Smelting is the process of reducing the roasted metallic oxide into metal in molten condition.
- ❖ Dilute or con. HNO_3 does not attack aluminium metal ,as it renders aluminium passive due to oxide film formation on its surface.
- ❖ The charge used in the metallurgy of iron consists of roasted ore, coke and limestone in the ratio, 8:4:1.
- ❖ Copper vessel on exposure to air and moisture forms a green layer on its surface due to basic copper carbonate.
- ❖ An alloy is a homogeneous mixture of two or more metals.
- ❖ An amalgam is an alloy of mercury with another metal. E.g. Ag-Sn amalgam is used for dental filling.
- ❖ The chemical name of rust is hydrated ferric oxide and its formula is $Fe_2O_3 \cdot xH_2O$.

TEXTBOOK EVALUATION

FBKZ2Y

I. Choose the best answer.

1. The number of periods and groups in the periodic table are______.
   a) 6,16 b) 7,17
   c) 8,18 d) 7,18

2. The basis of modern periodic law is______.
   a) atomic number
   b) atomic mass
   c) isotopic mass
   d) number of neutrons

3. _____ group contains the member of halogen family.
   a) 17^{th} b) 15^{th}
   c) 18^{th} d) 16^{th}

4. _____ is a relative periodic property
   a) atomic radii b) ionic radii
   c) electron affinity d) electronegativity

5. Chemical formula of rust is ________.
   a) $FeO.xH_2O$ b) $FeO_4.xH_2O$
   c) $Fe_2O_3.xH_2O$ d) FeO

6. In the alumino thermic process the role of Al is _____.
   a) oxidizing agent
   b) reducing agent
   c) hydrogenating agent
   d) sulphurising agent

7. The process of coating the surface of metal with a thin layer of zinc is called______.
   a) painting b) thinning
   c) galvanization d) electroplating

8. Which of the following have inert gases 2 electrons in the outermost shell.
   a) He b) Ne
   c) Ar d) Kr

9. Neon shows zero electron affinity due to ______.
   a) stable arrangement of neutrons
   b) stable configuration of electrons
   c) reduced size
   d) increased density

10. ______ is an important metal to form amalgam.
   a) Ag   b) Hg
   c) Mg   d) Al

II. Fill in the blanks

1. If the electronegativity difference between two bonded atoms in a molecule is greater than 1.7, the nature of bonding is ______
2. ________ is the longest period in the periodical table.
3. ________ forms the basis of modern periodic table.
4. If the distance between two Cl atoms in Cl_2 molecule is 1.98Å, then the radius of Cl atom is ________.
5. Among the given species A^-,A^+, and A, the smallest one in size is ________.
6. The scientist who propounded the modern periodic law is ________.
7. Across the period, ionic radii ________ (increases,decreases).
8. ________ and ________ are called inner transition elements.
9. The chief ore of Aluminium is ________.
10. The chemical name of rust is ________.

III. Match the following

1. Galvanisation	-	Noble gas elements
2. Calcination	-	Coating with Zn
3. Redox reaction	-	Silver-tin amalgam
4. Dental filling	-	Alumino thermic process
5. Group 18 elements	-	Heating in the absence of air

IV. True or False: (If false give the correct statement)

1. Moseley's periodic table is based on atomic mass.
2. Ionic radius increases across the period from left to right.
3. All ores are minerals; but all minerals cannot be called as ores;
4. Al wires are used as electric cables due to their silvery white colour.
5. An alloy is a heterogenous mixture of metals.

V. Assertion and Reason

Answer the following questions using the data given below:

i) A and R are correct, R explains the A.

ii) A is correct, R is wrong.

iii) A is wrong, R is correct.

iv) A and R are correct, R doesn't explains A.

1. **Assertion :** The nature of bond in HF molecule is ionic

   Reason : The electronegativity difference between H and F is 1.9
2. **Assertion :** Magnesium is used to protect steel from rusting

   Reason : Magnesium is more reactive than iron
3. **Assertion :** An uncleaned copper vessel is covered with greenish layer.

   Reason : copper is not attacked by alkali

VI. Short answer questions

1. A is a reddish brown metal, which combines with O_2 at < 1370 K gives B, a black coloured compound. At a temperature > 1370 K, A gives C which is red in colour. Find A,B and C with reaction.
2. A is a silvery white metal. A combines with O_2 to form B at 800°C, the alloy of A is used in making the aircraft. Find A and B
3. What is rust? Give the equation for formation of rust.
4. State two conditions necessary for rusting of iron.

VII. Long answer questions

1. a) State the reason for addition of caustic alkali to bauxite ore during purification of bauxite.

   b) Along with cryolite and alumina, another substance is added to the electrolyte mixture. Name the substance and give one reason for the addition.

2. The electronic configuration of metal A is 2,8,18,1.

   The metal A when exposed to air and moisture forms B a green layered compound. A with con. H_2SO_4 forms C and D along with water. D is a gaseous compound. Find A,B,C and D.

3. Explain smelting process.

VIII. HOT questions

1. Metal A belongs to period 3 and group 13. A in red hot condition reacts with steam to form B. A with strong alkali forms C. Find A,B and C with reactions

2. Name the acid that renders aluminium passive. Why?

3. a) Identify the bond between H and F in HF molecule.

   b) What property forms the basis of identification?

   c) How does the property vary in periods and in groups?

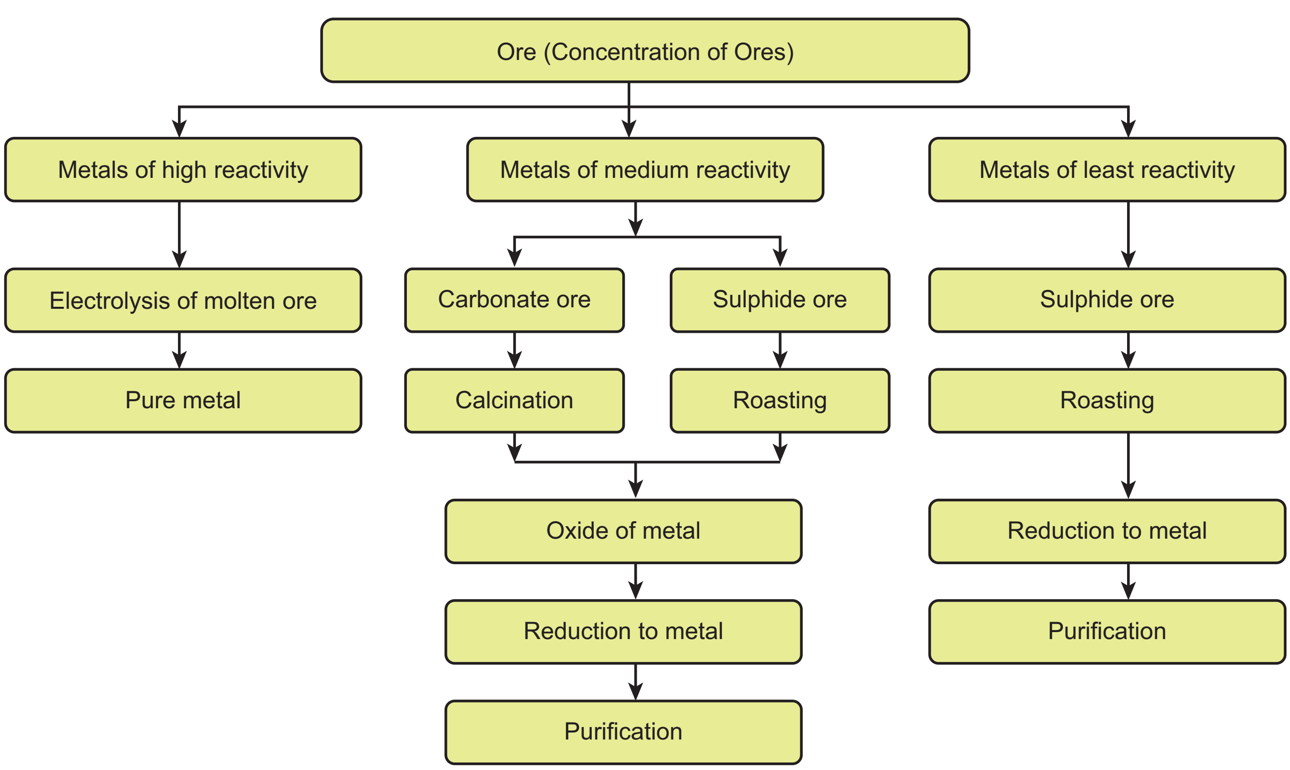
REFERENCE BOOKS

1. Inorganic chemistry by PL Soni
2. Physical chemistry by Puri and Sharma
3. Inorganic chemistry by Atkins
4. Oxford Inorganic chemistry

INTERNET RESOURCES

- https://www.webelements.com
- www.rsc.orgperiodic-table
- https://www.tcyonline.com

CONCEPT MAP



UNIT 9

SOLUTIONS

Learning Objectives

FQUY2M

After studying this lesson, students will able to

- define solution.
- recognize the types of solutions.
- analyse the factors influencing solubility.
- explain the various modes of expression of concentration of solution.
- calculate the solubility of solutes in solvents.
- correlate the hydrated salts and anhydrous salts.
- distinguish between deliquescent and hygroscopic substances.

INTRODUCTION

You have learnt about mixtures in your lower classes. Most of the substances that we encounter in our daily life are mixtures of two or more substances. The substances present in a mixture may exist in one or more physical state. For example, when we burn wood, the smoke released is a mixture of solid carbon and gases like CO_2, CO, etc.

In some cases of mixtures, their components can be separated easily whereas in some other cases they cannot be. Consider the two mixtures, one which contains salt and water, and the another which contains sand and water. Water is the one of the components in both the mixtures. In the first case salt disolves in water. In the second case the sand does not disolve in water. Sand in water can be separated by filtration but salt cannot be separated as it dissolves in water to form a homogeneous mixture. This kind of homogenous mixtures are termed as "**solutions**".

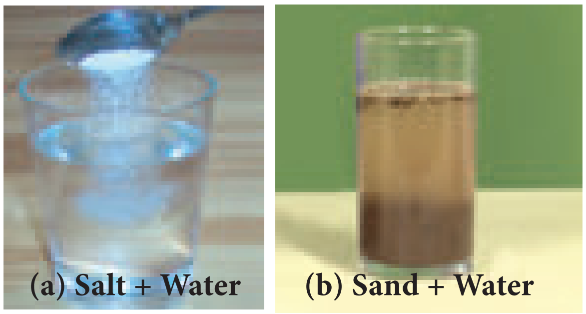


Figure 9.1 a) Homogeneous mixture
b)heterogeneous mixture

9.1 SOLUTIONS IN DAY-TO-DAY LIFE

One of the naturally existing solutions is sea water. We cannot imagine life on earth without sea water. It is a mixture of many dissolved salts. The another one is air. It is a mixture of gases like nitrogen, oxygen, carbon dioxide and other gases.

All the life forms on the earth are associated with solutions. Plants take solutions of nutrients for their growth from the soil. Most of the liquids found in human body including blood, lymph and urine are solutions. Day to day human activities like washing, cooking, cleaning and few other activities involve the formation of solutions with water. Similarly, the drinks what we take, like fruit juice, aerated drinks, tea, coffee etc. are also solutions. Therefore, the ability of water to form solutions is responsible for sustenance of life.

9.2 COMPONENTS OF SOLUTIONS

We know that, a **solution is a homogeneous mixture of two or more substances**. In a solution, the component which is present in lesser amount (by weight), is called **solute** and the component, which is present in a larger amount (by weight) is called **solvent**. The solute gets distributed uniformly throughout the solvent and thus forming the mixture homogeneous. So, the solvent acts as a dissolving medium in a solution. The process of uniform distribution of solute into solvent is called **dissolution**. Figure 9.2 shows the schematic representation of solution.

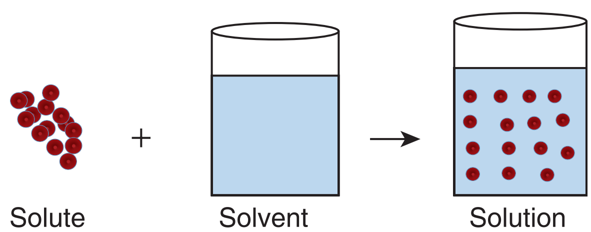
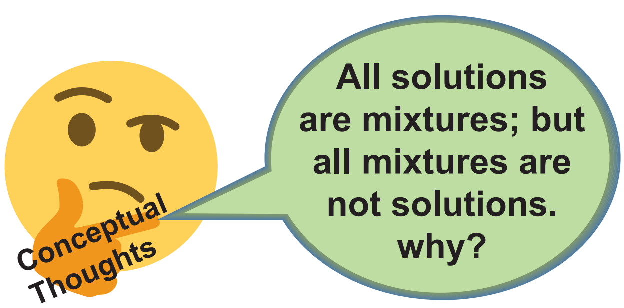


Figure 9.2 Formation of solution



A solution must at least be consisting of two components (a solute and a solvent). Such solutions which are made of one solute and one solvent (two components) are called **binary solutions**. e.g. On adding copper sulphate crystals to water, it dissolves in water forming a solution of copper sulphate as shown in Figure 9.3. It contains two components i.e. one solute- copper sulphate and one solvent-water. So it is a binary solution. Similarly, a solution may contain more than two components. For example if salt and sugar are added to water, both dissolve in water forming a solution. Here two solutes are dissolved in one solvent. Such kind of solutions which contain three components are called **ternary solutions**.

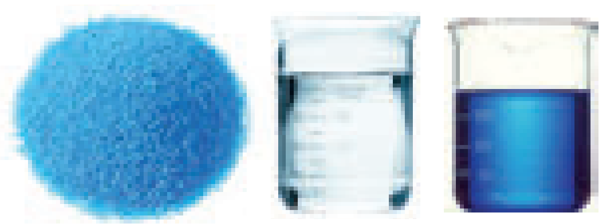


Figure 9.3 Formation of Copper sulphate solution

9.3 Types of Solutions

9.3.1 Based on the physical state of the solute and the solvent

We know that substances normally exist in three physical states (phases) i.e., solid, liquid and gas. In binary solutions, both the solvent and solute may exist in any of these physical states. But the solvent constitutes the major part of the solution. Its physical state is the primary factor which determine the characteristics of the solution. Therefore, there are different types of binary solutions as listed in Table 9.1.

Table 9.1 Types of binary solutions

Solute	Solvent	Example
Solid solution		
Solid	Solid	Copper dissolved in gold (Alloys)
Liquid	Solid	Mercury with sodium (amalgam)
Liquid solution		
Solid	Liquid	Sodium chloride dissolved in water
Liquid	Liquid	Ethyl alcohol dissolved in water
Gas	Liquid	carbon-di-oxide dissolved in water (Soda water)
Gaseous solution		
Liquid	Gas	Water vapour in air (cloud)
Gas	Gas	Mixture of Helium-Oxygen gases,

9.3.2 Based on the type of solvent

Most of the substances are soluble in water. That is why, water is called as 'Universal solvent". However some substances do not dissolve in water. Therefore, other solvents such as ethers, benzene, alcohols etc., are used to prepare a solution. On the basis of type of solvent, solutions are classified into two types. They are aqueous solutions and non-aqueous solutions.

a) Aqueous solution:

The solution in which water acts as a solvent is called aqueous solution. E.g. Common salt in water, Sugar in water, Copper sulphate in water etc.

b) Non – Aqueous solution:

The solution in which any liquid, other than water, acts as a solvent is called non-aqueous solution. Solvent other than water is referred to as non-aqueous solvent. Generally, alcohols, benzene, ethers, carbon disulphide, acetone, etc., are used as non-aqueous solvents. Examples for non-aqueous solutions: Sulphur dissolved in carbon disulphide, Iodine dissolved in carbon tetrachloride.

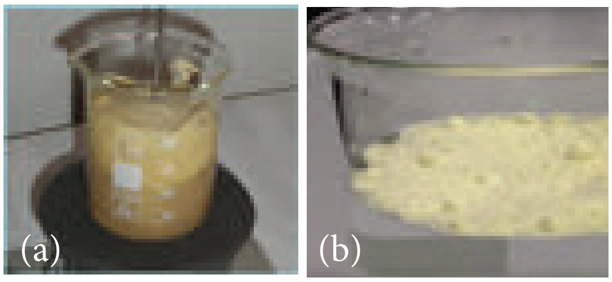


Figure 9.4 a) Sulphur is soluble in CS_2
b) Sulphur is insoluble in water

9.3.3 Based on the amount of solute

The amount of the solute that can be dissolved in the given amount of solvent is limited under any given conditions. Based on the amount of solute, in the given amount of solvent, solutions are classified into the following types:

(i) Saturated solution

(ii) Unsaturated solution

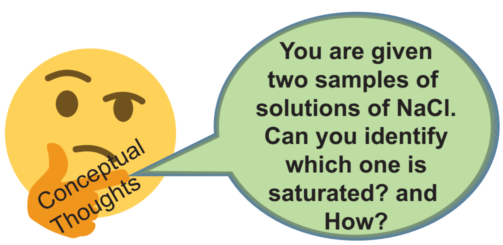
(iii) Super saturated solution

(i) Saturated solution: A solution in which no more solute can be dissolved in a definite amount of the solvent at a given temperature is called saturated solution. e.g. 36 g of sodium chloride in 100 g of water at 25°C forms saturated solution.

Further addition of sodium chloride, leave it undissolved.

(ii) Unsaturated solution: Unsaturated solution is one that contains less solute than that of the saturated solution at a given temperature. e.g. 10 g or 20 g or 30 g of Sodium chloride in 100 g of water at 25°C forms an unsaturated solution.

(iii) Super saturated solution: Supersaturated solution is one that contains more solute than the saturated solution at a given temperature. e.g. 40 g of sodium chloride in 100 g of water at 25°C forms super saturated solution. This state can be achieved by altering any other conditions liken temperature, pressure. Super saturated solutions are unstable, and the solute is reappearing as crystals when the solution is disturbed.



9.3.4 Concentrated and dilute solutions

It is another kind of classification of unsaturated solutions. It expresses the relative concentration of two solutions with respect to their solutes present in the given amount of the solvent. For example, you are given two cups of tea. When you taste them, you feel that one is sweeter than the other. What do you infer from it? The tea which sweet more contains higher amount of sugar than the other. How can you express your observation? You can say that the tea is stronger. But a chemist would say that it is 'concentrated'.

When we compare two solutions having same solute and solvent, the one which contains higher amount of solute per the given amount of solvent is said to be '**concentrated solution**' and the another is said to be '**dilute solution**'. They are schematically represented by Figure 9.5.

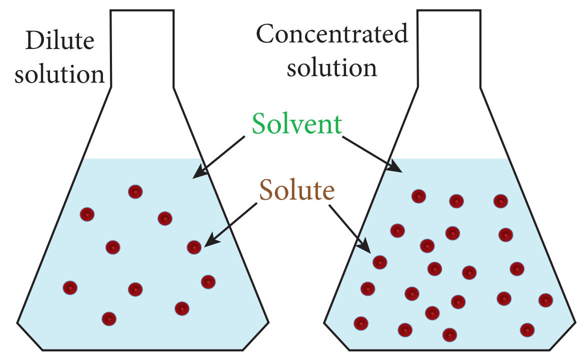
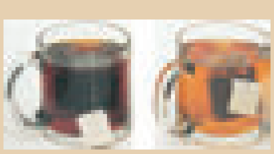


Figure 9.5 Dilute and Concentrated Solution

Differentiating solutions as dilute and concentrated is a qualitative representation. It does not imply the quantity of the solute. This difference is observed by means of some physical characteristics such as colour, density, etc.

Activity 1

Look at the following pictures. Label them as dilute and concentrated solution and justify your answer.



Tea

Copper sulphate

9.4 Solubility

Usually, there is a limit to the amount of solute that can be dissolved in a given amount of solvent at a given temperature. When this limit is reached, we have a saturated solution and any excess solute that is added, simply settle down at the bottom of the solution. The extent of dissolution of a solute in a solvent can be better explained by its solubility. Solubility is a measure of how much of a solute can be dissolved in a specified amount of a solvent.

Solubility is defined as the number of grams of a solute that can be dissolved in 100 g of a solvent to form its saturated solution at a given temperature and pressure. For example, 36 g of sodium chloride need to be dissolved in 100 g of water to form its saturated solution at 25°C. Thus the solubility of NaCl in water is 36 g at 25°C. The solubility is mathematically expressed as

$$\text{Solubility} = \frac{\text{Mass of the solute}}{\text{Mass of the solvent}} \times 100$$

Table 9.2 Solubility's of some common substances in water at 25°C

Name of the solute	Formula of the solute	Solubility g/100 g water
Calcium carbonate	$CaCO_3(s)$	0.0013
Sodium chloride	NaCl (s)	36
Ammonia	$NH_3(g)$	48
Sodium hydroxide	NaOH(s)	80
Glucose	$C_6H_{12}O_6(s)$	91
Sodium bromide	NaBr(s)	95
Sodium iodide	NaI(s)	184

9.4.1 Factors affecting solubility

There are three main factors which govern the solubility of a solute. They are:

(i) Nature of the solute and solvent

(ii) Temperature

(iii) Pressure

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(i) Nature of the solute and solvent

The nature of the solute and solvent plays an important role in solubility. Although water dissolves an enormous variety of substances, both ionic and covalent, it does not dissolve everything. The phrase that scientists often use when predicting solubility is "like dissolves like." This expression means that dissolving occurs when similarities exist between the solvent and the solute. For example: Common salt is a polar compound and dissolves readily in polar solvent like water.

Non-polar compounds are soluble in non-polar solvents. For example, Fat dissolved in ether. But non-polar compounds, do not dissolve in polar solvents; polar compounds do not dissolve in non-polar solvents.

(ii) Effect of Temperature

Solubility of Solids in Liquid:

Generally, solubility of a solid solute in a liquid solvent increases with increase in temperature. For example, a greater amount of sugar will dissolve in warm water than in cold water.

In endothermic process, solubility increases with increase in temperature.

In exothermic process, solubility decreases with increase in temperature.

Solubility of Gases in liquid

Do you know why is it bubbling when water is boiled? Solubility of gases in liquid decrease with increase in temperature. Generally, water contains dissolved oxygen. When water is heated, the solubility of oxygen in water decreases, so oxygen escapes in the form of bubbles.

Aquatic animals live more in cold regions because, more amount of dissolved oxygen is present in the water of cold regions. This shows that the solubility of oxygen in water is more at low temperatures.

(iii) Effect of Pressure

Effect of pressure is observed only in the case of solubility of a gas in a liquid. When the pressure is increased, the solubility of a gas in liquid increases.

The common examples for solubility of gases in liquids are carbonated beverages, i.e. soft drinks, household cleaners containing aqueous solution of ammonia, formalin- aqueous solution of formaldehyde, etc.

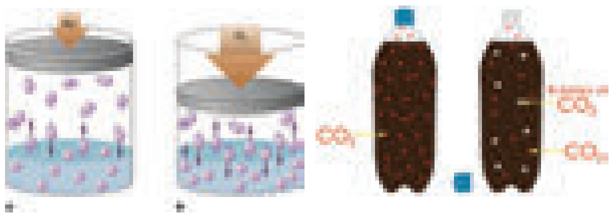

Figure 9.6 Effect of pressure on solubility

More to know

The effect of pressure on the solubility of a gas in liquid is given by **Henry's law**. It states that, the solubility of a gas in a liquid is directly proportional to the pressure of the gas over the solution at a definite temperature.

9.5 Concentration of a Solution

So far, we discussed what is a solution? what does it consist of and its types. Most of the chemical reactions take place in solutions form. So it is essential to quantify the solute in solvent to study the reactions. To quantify the solute in a solution, we can use the term "**concentration**".

Concentration of a solution may be defined as the amount of solute present in a given amount of solution or solvent.

Quantitatively, concentration of a solution may be expressed in different methods. But here, we shall discuss percentage by mass (% mass) and percentage by volume (% volume).

9.5.1 Mass percentage

Mass percentage of a solution is defined as the percentage by mass of the solute present in the solution. It is mostly used when solute is solid and solvent is liquid.

$$\text{Mass Percentage} = \frac{\text{Mass of the solute}}{\text{Mass of the solution}} \times 100$$

$$\text{Mass Percentage} = \frac{\text{Mass of the solute}}{\text{Mass of the solute + Mass of the solvent}} \times 100$$

For example: 5% sugar solution (by mass) means 5 g of sugar in 95 g of water. Hence it is made 100g of solution.

Usually, mass percentage is expressed as w/w (weight / weight); mass percentage is independent of temperature.

9.5.2 Volume percentage

Volume percentage is defined as the percentage by volume of solute (in ml) present in the given volume of the solution.

$$\text{Volume Percentage} = \frac{\text{Volume of the solute}}{\text{Volume of the solution}} \times 100$$

$$\text{Volume Percentage} = \frac{\text{Volume of the solute}}{\text{Volume of the solute + volume of the solvent}} \times 100$$

For example, 10% by volume of the solution of ethanol in water, means 10 ml of ethanol in 100 ml of solution (or 90 ml of water)

Usually volume percentage is expressed as v/v (volume / volume). It is used when both the solute and solvent are liquids. Volume percentage decreases with increases in temperature, because of expansion of liquid.

You can notice that in the commercial products that we come across in our daily life such as a solution of syrups, mouth wash, antiseptic solution, household disinfectants etc., the concentration of the ingredients is expressed as v/v. Similarly, in ointments, antacid, soaps, etc., the concentration of solutions are expressed as w/w.

Ointment

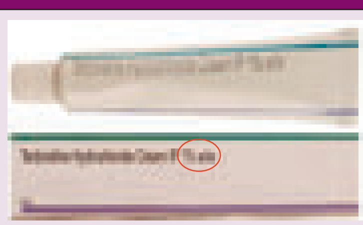

Figure 9.7 Ointment (w/w percent)

9.6 Hydrated salts and Water of Crystallization

When ionic substances are dissolved in water to make their saturated aqueous solution, their ions attract water molecules which then attached chemically in certain ratio. This process is called hydration. These ionic substances crystallize out from their saturated aqueous solution with a definite number of molecules of water. The number of water molecules found in the crystalline substance is called **water of crystallization**. Such salts are called hydrated salts.

Figure 9.8 a) Crystalline hydrated salt
b) Amorphous anhydrous salt

On heating these hydrated crystalline salts, they lose their water of crystallization and become amorphous or lose their colour (if they are coloured). Table 9.3 shows some common hydrated salts:

Table 9.3 Hydrated salts

Common Name	IUPAC Name	Molecular Formula
Blue Vitriol	Copper (II) sulphate pentahydrate	$CuSO_4\cdot 5H_2O$
Epsom Salt	Magnesium sulphate heptahydrate	$MgSO_4\cdot 7H_2O$
Gypsum	Calcium sulphate dihydrate	$CaSO_4\cdot 2H_2O$
Green Vitriol	Iron (II) sulphate heptahydrate	$FeSO_4\cdot 7H_2O$
White Vitriol	Zinc sulphate heptahydrate	$ZnSO_4\cdot 7H_2O$

9.6.1 Copper sulphate pentahydrate $CuSO_4\cdot 5H_2O$ (Blue vitriol)

The number of water molecules in blue vitriol is five. So its water of crystallization is 5. When blue coloured copper sulphate crystals are gently heated, it loses its five water molecules and becomes colourless anhydrous copper sulphate.

$$CuSO_4\cdot 5H_2O \underset{\text{Cooling}}{\overset{\text{Heating}}{\rightleftharpoons}} CuSO_4 + 5H_2O$$

(Copper sulphate pentahydrate) (Anhydrous copper sulphate)

Blue colour colourless

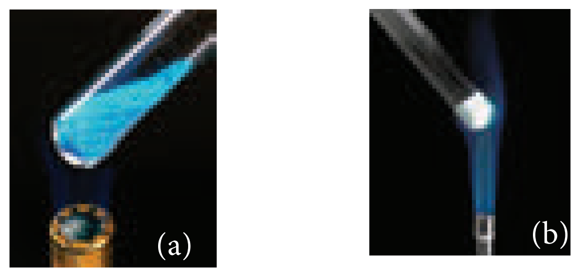

Figure 9.9 a) Copper sulphate before heating
b) Copper sulphate after heating

If you add few drops of water or allow it to cool, the colourless anhydrous salt again turns back into blue coloured hydrated salt.

Figure 9.10 Anhydrous copper sulphate turns to blue when water is added

9.6.2 Magnesium sulphate heptahydrate $MgSO_4\cdot 7H_2O$ (Epsom salt)

Its water of crystallization is 7. When magnesium sulphate heptahydrate crystals are gently heated, it loses seven water molecules, and becomes anhydrous magnesium sulphate.

$$MgSO_4{\cdot}7H_2O \underset{\text{Cooling}}{\overset{\text{Heating}}{\rightleftharpoons}} MgSO_4 + 7H_2O$$

(Magnesium sulphate heptahydrate) (Anhydrous Magnesium sulphate)

If you add few drops of water or allow it to cool, the colourless anhydrous salt again turns back into hydrated salt.

9.7 Hygroscopy

Certain substances, when exposed to the atmospheric air at ordinary temperature, absorb moisture without changing their physical state. Such substances are called **hygroscopic substances** and this property is called hygroscopy.

Hygroscopic substances are used as drying agents.

Example: 1. Conc.Sulphuric acid (H_2SO_4).
2. Phosphorus Pentoxide (P_2O_5).
3. Quick lime (CaO).
4. Silica gel (SiO_2).

9.8 Deliquescence

Certain substances which are so hygroscopic, when exposed to the atmospheric air at ordinary temperatures, absorb enough water and get completely dissolved. Such substances are called **deliquescent substances** and this property is called **deliquescence**.

Deliquescent substances lose their crystalline shape and ultimately dissolve in the absorbed water forming a saturated solution.

Deliquescence is maximum when:

1) The temperature is low
2) The atmosphere is humid

Examples: Caustic soda (NaOH), Caustic potash (KOH) and Ferric chloride ($FeCl_3$).

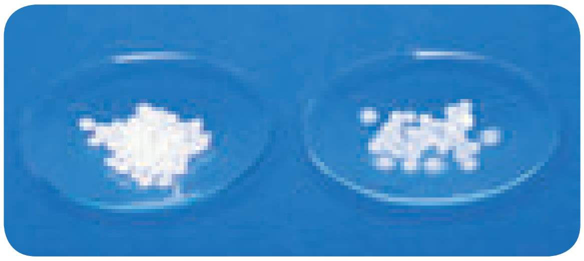

Figure 9.11 Deliquescence in Sodium hydroxide

9.9 Problems Based on Solubility and Percentage by Mass and Volume

I. Problems based on solubility

1) 1.5 g of solute is dissolved in 15 g of water to form a saturated solution at 298K. Find out the solubility of the solute at the temperature.

Table 9.3 Difference between hygroscopic substances and deliquescence.

Hygroscopic substances	Deliquescence substances
When exposed to the atmosphere at ordinary temperature, they absorb moisture and do not dissolve.	When exposed to the atmospheric air at ordinary temperature, they absorb moisture and dissolve.
Hygroscopic substances do not change its physical state on exposure to air.	Deliquescent substances change its physical state on exposure to air.
Hygroscopic substances may be amorphous solids or liquids.	Deliquescent substances are crystalline solids.

Mass of the solute = 1.5 g

Mass of the solvent = 15 g

$$\text{Solubility of the solute} = \frac{\text{Mass of the solute}}{\text{Mass of the solvent}} \times 100$$

$$\text{Solubility of the solute} = \frac{1.5}{15} \times 100$$

$$= 10 \text{ g}$$

2) Find the mass of potassium chloride would be needed to form a saturated solution in 60 g of water at 303 K? Given that solubility of the KCl is 37/100 g at this temperature.

Mass of potassium chloride in 100 g of water in saturated solution = 37 g

$$\text{Mass of potassium chloride in 60 g of water in saturated solution} = \frac{37}{100} \times 60$$

$$= 22.2 \text{ g}$$

3) What is the mass of sodium chloride that would be needed to form a saturated solution in 50 g of water at 30°C. Solubility of sodium chloride is 36 g at 30°C?

At 30°C, 36 g of sodium chloride is dissolved in 100 g of water.

∴ Mass of sodium chloride that would be need for 100 g of water = 36 g

$$\therefore \text{Mass of sodium chloride dissolved in 50 g of water} = \frac{36 \times 50}{100}$$

$$= 18 \text{ g}$$

4) The solubility of sodium nitrate at 50°C and 30°C is 114 g and 96 g respectively. Find the amount of salt that will be thrown out when a saturated solution of sodium nitrate containing 50 g of water is cooled from 50°C to 30°C?

Amount of sodium nitrate dissolved in 100 g of water at 50°C is 114 g

$$\therefore \text{Amount of sodium nitrate dissolving in 50 g of water at 50°C is} = \frac{114 \times 50}{100}$$

$$= 57 \text{ g}$$

$$\text{Similarly amount of sodium nitrate dissolving in 50g of water at 30°C is} = \frac{96 \times 50}{100}$$

$$= 48\text{g}$$

Amount of sodium nitrate thrown when 50g of water is cooled from 50°C to 30°C is

$$57 - 48 = 9 \text{ g}$$

II. Problem based on mass percentage

1) A solution was prepared by dissolving 25 g of sugar in 100 g of water. Calculate the mass percentage of solute.

Mass of the solute = 25 g

Mass of the solvent = 100 g

$$\text{Mass Percentage} = \frac{\text{Mass of the solute}}{\text{Mass of the solution}} \times 100$$

$$\text{Mass Percentage} = \frac{\text{Mass of the solute}}{\text{Mass of the solute + Mass of the solvent}} \times 100$$

$$= \frac{25}{25+100} \times 100$$

$$= \frac{25}{125} \times 100$$

$$= 20\%$$

2) 16 grams of NaOH is dissolved in 100 grams of water at 25°C to form a saturated solution. Find the mass percentage of solute and solvent.

Mass of the solute (NaOH) = 16 g

Mass of the solvent H_2O = 100 g

(i) **Mass percentage of the solute**

$$\text{Mass percentage of solute} = \frac{\text{Mass of the solute}}{\text{Mass of the solute + Mass of the solvent}} \times 100$$

$$= \frac{16 \times 100}{16 + 100}$$

$$= \frac{1600}{116}$$

Mass percentage of the solute = 13.79 %

(ii) Mass percentage of solvent = 100 – (Mass percentage of the solute)

= 100 – 13.79

= 86.21%

3) Find the amount of urea which is to be dissolved in water to get 500 g of 10% w/w aqueous solution?

$$\text{Mass percentage (w/w)} = \frac{\text{Mass of the solute}}{\text{Mass of the solution}} \times 100$$

$$10 = \frac{\text{Mass of the urea}}{500} \times 100$$

Mass of urea = 50g

(iii) **Problem based on volume – volume percentage.**

1) A solution is made from 35 ml of Methanol and 65 ml of water. Calculate the volume percentage.

Volume of the ethanol = 35 ml

Volume of the water = 65 ml

$$\text{Volume percentage} = \frac{\text{Volume of the solute}}{\text{Volume of the solution}} \times 100$$

$$\text{Volume percentage} = \frac{\text{Volume of the solute}}{\text{Volume of the solute + Volume of the solvent}} \times 100$$

$$\text{Volume percentage} = \frac{35}{35+65} \times 100$$

$$\text{Volume percentage} = \frac{35}{100} \times 100$$

= 35%

2) Calculate the volume of ethanol in 200 ml solution of 20% v/v aqueous solution of ethanol.

Volume of aqueous solution = 200 ml

Volume percentage = 20%

$$\text{Volume percentage} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

$$20 = \frac{\text{Volume of ethanol}}{200} \times 100$$

$$\text{Volume of ethanol} = \frac{20 \times 200}{100} = 40 \text{ ml}$$

Points to Remember

- A solution is a homogeneous mixture of two or more substances.
- An aqueous solution is a solution in which the solvent is water.
- A non-aqueous solution is a solution in which the solvent is a liquid, other than water
- A solution in which no more solute can be dissolved in a definite amount of the solvent at a given temperature is called saturated solution.
- An unsaturated solution is one that contains less solute than the saturated solution at a given temperature.
- A supersaturated solution is one that contains more solute than the saturated solution at a given temperature.
- Polar compounds are soluble in polar solvents.
- Non-polar compounds are soluble in non-polar solvents.
- In endothermic process, solubility of solid solute increases with increase in temperature.
- In exothermic process, solubility of solid solute decreases with increase in temperature.

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TEXTBOOK EVALUATION

I. Choose the correct answer.

1. A solution is a ___________ mixture.
   a. homogeneous b. heterogeneous
   c. homogeneous and heterogeneous
   d. non homogeneous

2. The number of components in a binary solution is ___________
   a. 2 b. 3
   c. 4 d. 5

3. Which of the following is the universal solvent?
   a. Acetone b. Benzene
   c. Water d. Alcohol

4. A solution in which no more solute can be dissolved in a definite amount of solvent at a given temperature is called ________
   a. Saturated solution
   b. Un saturated solution
   c. Super saturated solution
   d. Dilute solution

5. Identify the non aqueous solution.
   a. sodium chloride in water
   b. glucose in water
   c. copper sulphate in water
   d. sulphur in carbon-di-sulphide

6. When pressure is increased at constant temperature the solubility of gases in liquid ___________.
   a. No change b. increases
   c. decreases d. no reaction

7. Solubility of NaCl in 100 ml water is 36 g. If 25 g of salt is dissolved in 100 ml of water how much more salt is required for saturation ______________.
   a. 12g b. 11g c. 16g d. 20g

8. A 25% alcohol solution means
   a. 25 ml alcohol in 100 ml of water
   b. 25 ml alcohol in 25 ml of water
   c. 25 ml alcohol in 75 ml of water
   d. 75 ml alcohol in 25 ml of water

9. Deliquescence is due to ___________
   a. Strong affinity to water
   b. Less affinity to water
   c. Strong hatred to water
   d. Inertness to water

10. Which of the following is hygroscopic in nature?
    a. ferric chloride
    b. copper sulphate penta hydrate
    c. silica gel
    d. none of the above

II. Fill in the blanks

1. The component present in lesser amount, in a solution is called ________
2. Example for liquid in solid type solution is ______
3. Solubility is the amount of solute dissolved in ______ g of solvent.
4. Polar compounds are soluble in ______ solvents
5. Volume persentage decreases with increases in temperature because ______

III. Match the following

1. Blue vitriol	– $CaSO_4{\cdot}2H_2O$
2. Gypsum	– CaO
3. Deliquescence	– $CuSO_4{\cdot}5H_2O$
4. Hygroscopic	– NaOH

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IV. True or False: (If false give the correct statement)

1. Solutions which contain three components are called binary solution.
2. In a solution the component which is present in lesser amount is called solvent.
3. Sodium chloride dissolved in water forms a non-aqueous solution.
4. The molecular formula of green vitriol is $MgSO_4 \cdot 7H_2O$
5. When Silica gel is kept open, it absorbs moisture from the air, because it is hygroscopic in nature

V. Short answer

1. Define the term: Solution
2. What is mean by binary solution
3. Give an example each i) gas in liquid ii) solid in liquid iii) solid in solid iv) gas in gas
4. What is aqueous and non-aqueous solution? Give an example.
5. Define Volume percentage
6. The aquatic animals live more in cold region Why?
7. Define Hydrated salt.
8. A hot saturated solution of copper sulphate forms crystals as it cools. Why?
9. Classify the following substances into deliquescent, hygroscopic.

   Conc. Sulphuric acid, Copper sulphate penta hydrate, Silica gel, Calcium chloride, and Gypsum salt.

VI. Long answer:

1. Write notes on i) saturated solution ii) unsaturated solution
2. Write notes on various factors affecting solubility.
3. a) What happens when $MgSO_4 \cdot 7H_2O$ is heated? Write the appropriate equation

   b) Define solubility
4. In what way hygroscopic substances differ from deliquescent substances.
5. A solution is prepared by dissolving 45 g of sugar in 180 g of water. Calculate the mass percentage of solute.
6. 3.5 litres of ethanol is present in 15 litres of aqueous solution of ethanol. Calculate volume percent of ethanol solution.

VII. HOTS

1. Vinu dissolves 50 g of sugar in 250 ml of hot water, Sarath dissolves 50 g of same sugar in 250 ml of cold water. Who will get faster dissolution of sugar? and Why?
2. 'A' is a blue coloured crystaline salt. On heating it loses blue colour and to give 'B'. When water is added, 'B' gives back to 'A'. Identify A and B, write the equation.
3. Will the cool drinks give more fizz at top of the hills or at the foot? Explain

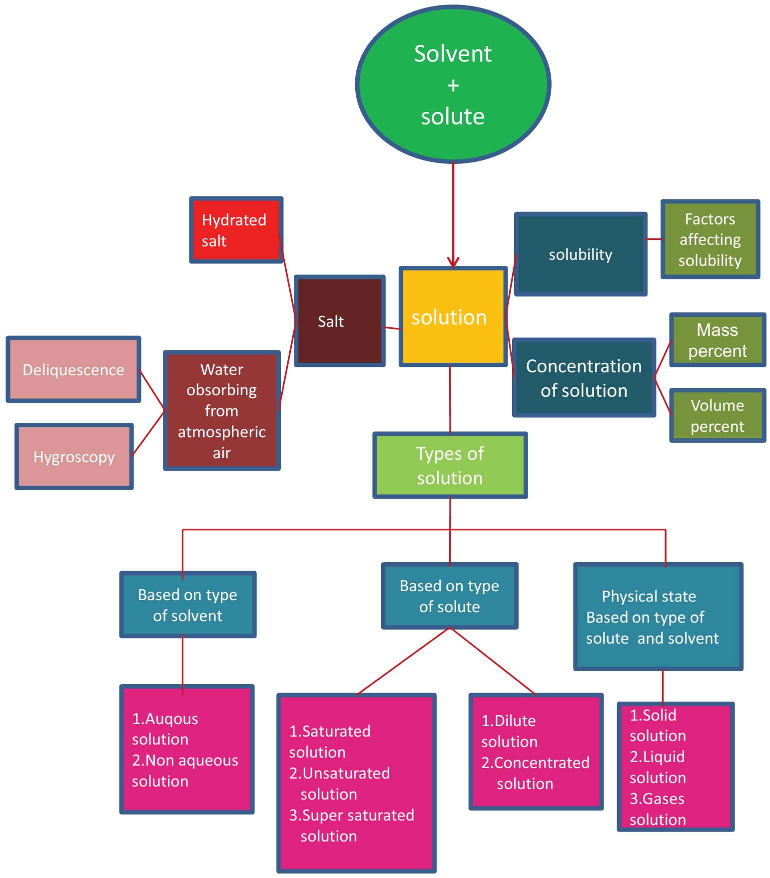
REFERENCE BOOKS

1. Properties Liquids Solutions John Murrell 2nd Edition.
2. Fundamental Interrelationships Between Certain Soluble Salts and Soil Colloids (Classic Reprint) Hardcover, by Leslie Theodore Sharp

INTERNET RESOURCES

1. https://www.cwcboe.org/cms/lib/NJ01001185/Centricity/Domain/203/Solutions%20Suspensions%20and%20Colloids.pdf

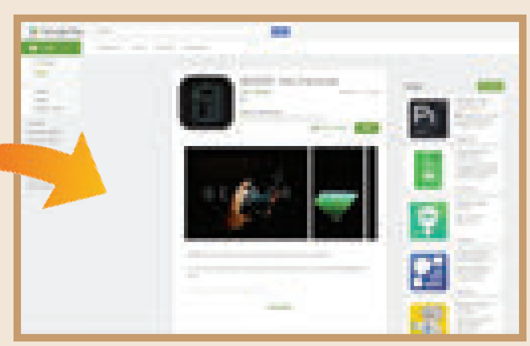
Concept Map



ICT CORNER

Solutions

BEAKER application enable the students to use their mobile as virtual chemistry laboratory and also to do various experiments on their own.



Steps

- Access the application "BEAKER – Mix Chemicals" with help of the URL or QR code, Install it in the mobile. You can see that the screen will act like a beaker after opening the application.
- If you click the round button, you can see many elements and compounds.
- If you click any elements and compounds, it will be added to the beaker in the home screen.
- By clicking Menu at the left side, You can see lid, match stick, burner and chemist. Use those whenever necessary.

URL: https://play.google.com/store/apps/details?id=air.thix.sciencesense.beaker

or Scan the QR Code.

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UNIT 10

TYPES OF CHEMICAL REACTIONS

Learning Objectives

GKWN5A

After completing this lesson learners will be able to

- infer different types of chemical reaction.
- acquire knowledge about combination reaction and skill to perform a combination reaction using quick lime and water.
- identify and differentiate between reversible and irreversible reactions.
- explain the reversible reaction occurring at the equilibrium state.
- list and explain characteristics of equilibrium state.
- define rate of reaction.
- discuss the dependence of rate of reactions on concentration, temperature and catalyst.
- define pH.
- correlate the concentration of hydrogen ions and pH with neutral, acidic and basic nature of aqueous solutions.
- recognize the importance of pH in everyday life.
- explain the term ionic product of water.

INTRODUCTION

As you know from your earlier studies, a chemical reaction involves breaking of old chemical bonds and formation of new chemical bonds. This change may happen spontaneously or it may be facilitated by external forces or energy. Chemistry is all about chemical reactions. In your day to day life, you could observe many chemical reactions. A clear understanding of these reactions is essential in order to manipulate them for the sake of human life and environment. So, chemistry mainly focuses on chemical reactions. Let us try to find the answer for the following questions:

- You need energy to play, walk, run or to perform various physical activities. Where do you get the energy from?
- How do plants grow and get their food?
- How does a car move using fuel?
- Why does iron rust on its exposure to water or air?

You get energy from the digestion of the food you eat. Plants grow by absorbing nutrients from the Earth and get their food by photosynthesis. The combustion of a fuel makes the car to move. Oxidation of iron causes rusting. So, all these processes are chemical changes i.e. the materials, which undergo changes are converted into some other new materials. For example, by burning petrol, the hydrocarbons present in it are converted into carbon dioxide and water. In this chapter, let us discuss the nature and types of chemical reactions.

What happens during a chemical reaction?

- In a chemical reaction, the atoms of the reacting molecules or elements are rearranged to form new molecules.
- Old chemical bonds between atoms are broken and new chemical bonds are formed.
- Bond breaking absorbs energy whereas bond formation releases energy

How are chemical reactions represented?

When methane reacts with oxygen, it forms carbon dioxide and water. How can you represent this reaction? It can be written as a word equation as shown below:

Methane + Oxygen → Carbon dioxide + Water

But, this equation does not give the chemical composition of the reactants and products. So, to learn the characteristics of a chemical reaction, it is represented by a chemical equation. In the chemical equation, the chemicals of the reaction are represented by their chemical formulas.The compounds or elements, which undergo reactions (reactants) are shown to the left of an arrow and the compounds formed (products) are shown to the right of the arrow. The arrow indicates the direction of the reaction. Thus, the aforesaid reaction can be written as follows:

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

But, this is also an incomplete chemical equation. Because, the law of conservation of matter states that matter cannot be created or destroyed. You cannot create new atoms by a chemical reaction. In contrast, they are rearranged in different ways by a chemical reaction to form a new compound. So, in a chemical equation, the number of atoms of the reactants and that of the products must be equal. The number of hydrogen and oxygen atoms in the reactants and the products are not equal in the given equation. On balancing the number of atoms, the following equation can be obtained:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Further, the chemical equation provides information on the physical state of the substances and the conditions under which the reaction takes place.

$$\underset{\text{Methane}}{CH_{4(g)}} + \underset{\text{Oxygen}}{2O_{2(g)}} \rightarrow \underset{\text{Carbon dioxide}}{CO_{2(g)}} + \underset{\text{Water}}{2H_2O_{(g)}}$$

A balanced chemical equation is the simplified representation of a chemical reaction which describes the chemical composition, physical state of the reactants and the products, and the reaction conditions.

MORE TO KNOW

The phases or the physical state of the substances in a chemical reaction are denoted in short form within a bracket, as the subscript of the formula, of the respective substances. For example, when solid potassium reacts with liquid water, it produces hydrogen gas and potassium hydroxide solution. All these information of the reaction is given in the chemical equation as shown below:

$$2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$$

Symbol	Phase or physical state
s	Solid
l	Liquid
g	Gas
aq	Aqueous Solution

10.1 TYPES OF CHEMICAL REACTIONS

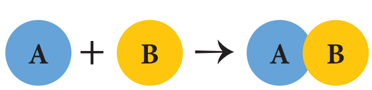
Classification based on the nature of rearrangements of atoms

So far you studied about a chemical reaction and how it can be described as a chemical equation. A large number of chemical reactions are taking place around us every day. Are they taking place in a similar way? No. Each reaction involves different kinds of atoms and hence the way they react also differs. Thus, based on the manner by which the atoms of the reactants are rearranged, chemical reactions are classified as follows.

H1FI6L

(a) Combination reactions

A combination reaction is a reaction in which two or more reactants combine to form a compound. It is otherwise called 'synthesis reaction' or 'composition reaction'. When a reactant 'A' combines with 'B', it forms the product 'AB'. The generalised scheme of a combination reaction is given below:



Example: Hydrogen gas combines with chlorine gas to form hydrogen chloride gas.

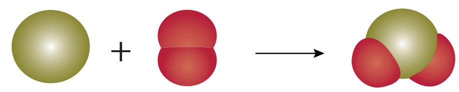
$$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$$

Depending on the chemical nature of the reactants, there are **three classes** of combination reactions:

❖ **Element + Element → Compound**

In this type of combination reaction, two elements react with one other to form a compound. The reaction may take place between a metal and a non-metal or two non-metals.

Example 1: When solid sulphur reacts with oxygen, it produces sulphur dioxide. Here both the reactants are non-metals.



$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$$

Example 2: Sodium, a silvery-white metal, combines with chlorine, a pale yellow green gas, to form sodium chloride, an edible compound. Here one of the reactants is a metal (sodium) and the other (chlorine) is a non-metal.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Test Yourself:

Identify the possible combination reactions between the metals and non-metals given in the following table and write their balanced chemical equations:

Metals	Non-metals
Na, K, Cs, Ca, Mg	F, Cl, Br, I

❖ **Compound + Element → Compound**

In this case, a compound reacts with an element to form a new compound. For instance, phosphorous trichloride reacts with chlorine gas and forms phosphorous pentachloride.

$$PCl_{3(l)} + Cl_{2(g)} \rightarrow PCl_{5(s)}$$

❖ **Compound + Compound → Compound**

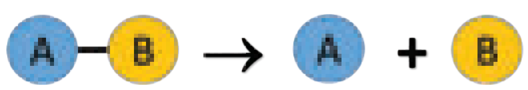
It is a reaction between two compounds to form a new compound. In the following reaction, silicon dioxide reacts with calcium oxide to form calcium silicate.

$$SiO_{2(s)} + CaO_{(s)} \rightarrow CaSiO_{3(s)}$$

Most of the combination reactions are exothermic in nature. Because, they involve the formation of new bonds, which releases a huge amount of energy in the form of heat.

(b) Decomposition reactions

In a decomposition reaction, a single compound splits into two or more simpler substances under suitable conditions. It is the opposite of the combination reaction. The **generalised scheme** of a decomposition reaction is given below:



Breaking of bonds is the major phenomenon in a decomposition reaction and hence it requires energy to break the bonds, depending on the nature of the energy used in the decomposition reaction.

There are three main classes of decomposition reactions. They are

(i) Thermal Decomposition Reactions

(ii) Electrolytic Decomposition Reactions

(iii) Photo Decomposition Reactions

> **DO YOU KNOW?**
>
> A solution of slaked lime 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** $CaCO_3$
>
> $$\underset{\text{Slaked Lime}}{Ca(OH)_{2(aq)}} + \underset{\text{Carbon dioxide}}{CO_{2(g)}} \rightarrow \underset{\text{Calcium Carbonate}}{CaCO_{3(s)}} + \underset{\text{Water}}{H_2O_{(l)}}$$

(i) Thermal Decomposition Reactions

In this type of reaction, the reactant is decomposed by applying heat. For example, on heating mercury (II) oxide is decomposed into mercury metal and oxygen gas. As the molecule is dissociated by the absorption of heat, it is otherwise called **'Thermolysis'.** It is a class of compound to element/element decomposition. i.e. a compound (HgO) is decomposed into two elements (Hg and Oxygen).

$$2HgO_{(s)} \xrightarrow{\text{Heat}} 2Hg_{(l)} + O_{2(g)}$$

Similarly, when calcium carbonate is heated, it breaks down in to calcium oxide and carbon dioxide. It is a type of compound to compound/compound decomposition.

$$CaCO_{3(s)} \xrightarrow{\text{Heat}} CaO_{(s)} + CO_{2(g)}$$

In thermal decomposition reaction, heat is supplied to break the bonds. Such reactions, in which heat is absorbed, are called **'Endothermic reactions'**.

(ii) Electrolytic Decomposition Reactions

In some of the decomposition reactions, electrical energy is used to bring about the reaction. For example, decomposition of sodium chloride occurs on passing electric current through its aqueous solution. Sodium chloride decomposes in to metallic sodium and chlorine gas. This process is termed as 'Electrolysis'.

$$2NaCl_{(aq)} \xrightarrow{\text{Electricity}} 2Na_{(s)} + Cl_{2(g)}$$

Here, a compound (NaCl) is converted into elements (Na and chlorine). So it is a type of compound to element/element decomposition.

(iii) Photo Decomposition Reactions

Light is an another form of energy, which facilitates some of the decomposition reactions. For example, when silver bromide is exposed to light, it breaks down into silver metal and bromine gas. As the decomposition is caused by light, this kind of reaction is also called **'Photolysis'.**

$$2AgBr_{(s)} \xrightarrow{\text{Light}} 2Ag_{(s)} + Br_{2(g)}$$

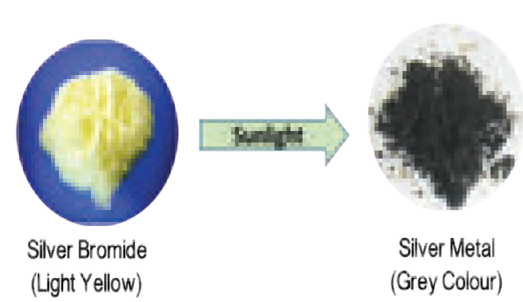
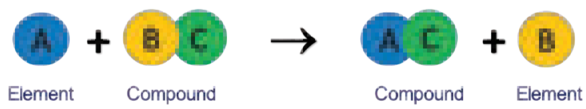


Figure 10.1 Photo decomposition of silver bromide

The yellow coloured silver bromide turns into grey coloured silver metal. It is also a compound to element/element decomposition.

(c) Single Displacement Reactions

It is a reaction between an element and a compound. When they react, one of the elements of the compound-reactant is replaced by the element-reactant to form a new compound and an element. The general schematic representation of a single displacement reaction is given as:



'A' displaces element 'B' from the compound 'BC' and hence a single displacement reaction occurs. If zinc metal is placed in hydrochloric acid, hydrogen gas is evolved. Here, hydrogen is displaced by zinc metal and zinc chloride is formed.

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$$

If an iron nail is placed in an aqueous solution of copper (II) sulphate as shown in Fig. 10.2, the iron displaces copper from its aqueous solution and the so formed copper deposits over the iron nail.

Figure 10.2 Displacement of copper

It is easy to propose so many reactions of this kind with different combinations of reactants. Will they all occur in practice? No. This is most easily demonstrated with halogens. Let us consider the following two reactions:

$$2NaCl_{(aq)} + F_{2(g)} \rightarrow 2NaF_{(aq)} + Cl_{2(g)}$$

$$2NaF_{(aq)} + Cl_{2(g)} \rightarrow 2NaCl_{(aq)} + F_{2(g)}$$

The first reaction involves the displacement of chlorine from NaCl, by fluorine. In the second reaction, chlorine displaces fluorine from NaF. Out of these two, the second reaction will not occur. Because, fluorine is more active than chlorine and occupies the upper position in the periodic table. So, in displacement reactions, the activity of the elements and their relative position in the periodic table are the key factors to determine the feasibility of the reactions. More active elements readily displace less active elements from their aqueous solution.

The activity series of some elements is given below:

To remember	Activity Series	
Please	Potassium (K)	Most reactive
Send	Sodium (Na)	
Lions	Lithium (Li)	
Cats	Calcium (Ca)	
Monkeys	Magnesuim (Mg)	
And	Aluminium (Al)	
Zebras	Zinc (Zn)	
Into	Iron (Fe)	
Lovely	Lead (Pb)	
Hot	Hydrogen (H) non-metal	
Countries	Copper (Cu)	
Signed	Silver (Ag)	
General	Gold (Au)	
Penguin	Platinum (Pt)	Least reactive

By referring the activity series, try to answer the following questions:

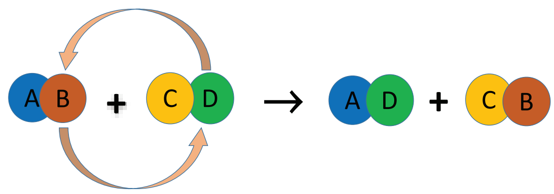
Which of the metals displaces hydrogen gas from hydrochloric acid? Silver or Zinc. Give the chemical equation of the reaction and Justify your answer.

Activity 10.1

- Take about 50 ml of toilet cleaning acid in a beaker
- Place a small iron nail in it
- Wait for about 10 minutes
- Observe what happen in the beaker
- Do you recognize any change?
- Report your observations with the chemical equation.

(d) Double Displacement Reactions

When two compounds react, if their ions are interchanged, then the reaction is called double displacement reaction. The ion of one compound is replaced by the ion of the another compound. Ions of identical charges are only interchanged, i.e., a cation can be replaced by other cations. This reaction is also called '**Metathesis Reaction**'. The schematic representation of a double displacement reaction is given below:



In a double displacement reaction either one of the products must be either a precipitate or water. By this way, there are major classes of double displacement reactions. They are:

(i) Precipitation Reactions

(ii) Neutralization Reactions

(i) Precipitation Reactions

When aqueous solutions of two compounds are mixed, if they react to form an insoluble compound and a soluble compound, then it is called precipitation reaction. Because the insoluble compound, formed as one of the products, is a precipitate and hence the reaction is so called.

Table 10.1 Differences between combination and decomposition reactions

COMBINATION REACTIONS	DECOMPOSITION REACTIONS
One or more reactants combine to form a single product	A single reactant is decomposed to form one or more products
Energy is released	Energy is absorbed
Elements or compounds may be the reactants	Single compound is the reactant

When the clear aqueous solutions of potassium iodide and lead (II) nitrate are mixed, a double displacement reaction takes place between them.

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(s)}\downarrow + 2KNO_{3(aq)}$$

Potassium and lead displace or replace one other and form a yellow precipitate of lead (II) iodide as shown in Fig. 10.3.

Figure 10.3 Precipitation of PbI_2

Activity 10.2

- Take a pinch of silver nitrate crystals.
- Collect about 5 ml of tap water in a test tube.
- Add the silver nitrate crystals to water and shake well.
- Observe what happen in the test tube.
- Report your observations and what you infer from that?

(ii) Neutralization Reactions

In your lower classes, you have learned the reaction between an acid and a base. It is another type of displacement reaction in which the acid reacts with the base to form a salt and water. It is called 'neutralization reaction' as both acid and base neutralize each other.

Acid + Base $\rightarrow$ Salt + Water

Reaction of sodium hydroxide with hydrochloric acid is a typical neutralization reaction. Here, sodium replaces hydrogen from hydrochloric acid forming sodium chloride, a neutral soluble salt.

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Similarly, when ammonium hydroxide reacts with nitric acid, it forms ammonium nitrate and water.

$$HNO_{3(aq)} + NH_4OH_{(aq)} \rightarrow NH_4NO_{3(aq)} + H_2O_{(l)}$$

(e) Combustion Reactions

A combustion reaction is one in which the reactant rapidly combines with oxygen to form one or more oxides and energy (heat). So in combustion reactions, one of the reactants must be oxygen. Combustion reactions are majorly used as heat energy sources in many of our day to day activities. For instance, we use LPG gas for domestic cooking purposes. We get heat and flame from LPG gas by its combustion reaction of its constituent gases. LPG is a mixture of hydrocarbon gases like propane, butane, propylene, etc. All these hydrocarbons burn with oxygen to form carbon dioxide and water.

$$\underset{\text{Propane}}{C_3H_{8(g)}} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)} + \text{Heat}$$

Since heat is evolved, it is an exothermic reaction. As oxygen is added, it is also an oxidation. So, combustion may be called as an exothermic oxidation. If a flame is formed (as shown in Fig. 10.4), then it is called **burning**.

Figure 10.4 Combustion of LPG gas

Which of the following is a combustion?

(i) Digestion of Food

(ii) Rusting of iron

Many thousands of reactions fall under these five categories and further you will learn in detail about these reactions in your higher classes.

10.1.2 Classification based on the direction of the reaction

You know that innumerable changes occur every day around us. Are all they permanent? For example, liquid water freezes into ice, but then ice melts into liquid water. In other words, freezing is reversed. So, it is not a permanent change. Moreover, it is a physical change. Physical changes can be reversed easily. Can chemical changes be reversed? Can the products be converted into reactants? Let us consider the burning of a wood. The carbon compounds present in the wood are burnt into carbon dioxide gas and water. Can we get back the wood immediately from carbon dioxide and water? We cannot. So, it is a permanent change. In most of the cases, we cannot. But, some chemical reactions can be reversed. Our mobile phone gets energy from its lithium ion battery by chemical reactions. It is called discharging. On recharging the mobile, these chemical reactions are reversed. Thus, chemical reactions may be reversed under suitable conditions. Hence, they are grouped into two categories such as reversible and irreversible reactions.

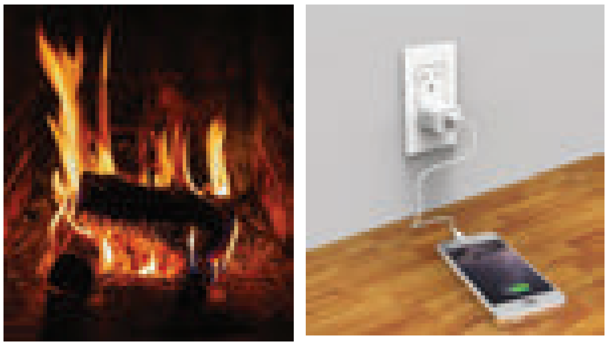
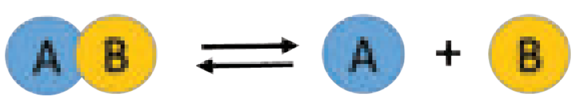


Figure 10.5 Burning of wood and recharging of mobile battery

Reversible Reactions

A reversible reaction is a reaction in which the products can be converted back to the reactants. A reversible reaction is represented by a double arrow with their heads in the direction opposite to each other. Thus, a reversible reaction can be represented by the following equation:



Explanation: Here, the compound 'AB' undergoes decomposition to form the products 'A' and 'B'. It is the **forward reaction**. As soon as the products are formed, they combine together to form 'AB'. It is the **backward reaction**. So, the reaction takes place in both the directions. Do you think then that no products are formed in the aforesaid reaction? If you think so, you are wrong. Because, even though the reaction takes place in both the directions, at the initial stage the rates (speed) of these reactions are not equal. Consider the following decomposition reaction of phosphorous pentachloride into phosphorous trichloride and chlorine.

$$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$$

The forward reaction is the decomposition of PCl_5 and the backward reaction is the combination of PCl_3 and Cl_2. Initially, the forward reaction proceeds faster than the backward reaction. After sometimes, the speed of both the reactions become equal. So, PCl_5 cannot be completely converted into the products as the reaction is reversed. It is a reversible reaction. The actual measurements of the given reaction show that the reaction is at equilibrium, but the amount of PCl_5 is more than that of PCl_3 and Cl_2.

MORE TO KNOW

If hydrogen peroxide is poured on a wound, it decomposes into water and oxygen. The gaseous oxygen bubbles away as it is formed and thus prevent the formation of H_2O_2.

Hydrogen peroxide on a wound

Thus, more amount of products can be obtained in a reversible reaction by the periodical removal of one of the products or the periodical addition of the reactants.

$$2H_2O_{2(aq)} \rightleftharpoons 2H_2O_{(l)} + O_{2(g)}$$

Irreversible Reactions

The reaction that cannot be reversed is called **irreversible reaction**. The irreversible reactions are unidirectional, i.e., they take place only in the forward direction. Consider the combustion of coal into carbon dioxide and water.

$$\underset{\text{Coal}}{C_{(s)}} + \underset{\text{Oxygen}}{O_{2(g)}} \rightarrow \underset{\text{Carbon dioxide}}{CO_{2(g)}} + \text{Heat}$$

In this reaction, solid coal burns with oxygen and gets converted into carbon dioxide gas and water. As the product is a gas, as soon as it is formed it escapes out of the reaction container. It is extremely hard to decompose a gas into a solid. Thus, the backward reaction is not possible in this case. So, it is an irreversible reaction. Table 10.2 provides the main differences between a reversible and an irreversible reaction:

Table 10.2 Differences between reversible and irreversible reactions

REVERSIBLE REACTION	IRREVERSIBLE REACTION
It can be reversed under suitable conditions.	It cannot be reversed.
Both forward and backward reactions take place simultaneously.	It is unidirectional. It proceeds only in forward direction.
It attains equilibrium.	Equilibrium is not attained.
The reactants cannot be converted completely into products.	The reactants can be completely converted into products.
It is relatively slow.	It is fast.

You will learn more about these reactions in your higher classes.

10.2 RATE OF A CHEMICAL REACTION

So far we discussed various types of chemical reactions and the nature of the reactants and products. Let us consider the following reactions:

- Rusting of iron
- Digestion of food
- Burning of petrol
- Weathering of rock

How fast is each reaction? Rank them from the slowest to fastest. How will you determine, which is the fastest and which is the slowest? One of the ways to find out how fast a reaction is as follows: Measure the amount of reactants or products before and after a specific period of time. For example, let us assume that 100 g of a substance 'A' undergoes a reaction and after an hour 50 g of 'A' is left.

$$A \rightarrow \text{Product}$$

In another instance, 100 g of substance 'C' undergoes a reaction and after an hour, 20 g of 'C' is left.

$$C \rightarrow \text{Product}$$

Can you say which is the faster reaction? In the first reaction, 50 g of the reactant (A) is converted into products whereas in the second reaction 80 g of the reactant is converted into products in one hour. So, the second reaction is faster. This measurement is called 'the reaction rate'.

"Rate of a reaction is the change in the amount or concentration of any one of the reactants or products per unit time".

Consider the following reaction

$$A \rightarrow B$$

The rate of this reaction is given by

$$\text{Rate} = -\frac{d[A]}{dt} = +\frac{d[B]}{dt}$$

Where,

[A] – Concentration of A

[B] – Concentration of B

The negative sign indicates the decrease in the concentration of A with time.

The postive sign indicates the increase in the concentration of B with time.

Note: '[ ]' represents the concentration, 'd' represents the infinitesimal change in the concentration.

Why is reaction rate important?

Faster the reaction, more will be the amount of the product in a specified time. So, the rate of a reaction is important for a chemist for designing a process to get a good yield of a product. Rate of reaction is also important for a food processor who hopes to slow down the reactions that cause food to spoil.

10.2.1 Factors influencing the rate of a reaction

Can the rate of a reaction be changed? The rate of a reaction can be changed. For

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example, iron gets rusted faster in an acid than in water. Important factors that affect rate of a reaction are

(i) Nature of the reactants
(ii) Concentration of the reactants
(iii) Temperature
(iv) Catalyst
(v) Pressure
(vi) Surface area of the reactants

(i) Nature of the reactants

The reaction of sodium with hydrochloric acid is faster than that with acetic acid. Do you know why? Hydrochloric acid is a stronger acid than acetic acid and thus more reactive. So, the nature of the reactants influence the reaction rate.

$$2Na_{(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_{2\,(g)} \text{ (fast)}$$
$$2Na_{(s)} + 2CH_3COOH_{(aq)} \rightarrow 2CH_3COONa_{(aq)} + H_{2(g)} \text{ (slow)}$$

(ii) Concentration of the reactants

Changing the amount of the reactants also increases the reaction rate. The amount of the substance present in a certain volume of the solution is called **'concentration'**. More the concentration, more particles per volume exist in it and hence faster the reaction. Granulated zinc reacts faster with 2M hydrochloric acid than 1M hydrochloric acid.

(iii) Temperature

Most of the reactions go faster at higher temperature. Because adding heat to the reactants provides energy to break more bonds and thus speed up the reaction. Calcium carbonate reacts slowly with hydrochloric acid at room temperature. When the reaction mixture is heated the reaction rate increases.

DO YOU KNOW?

Food kept at room temperature spoils faster than that kept in the refrigerator. In the refrigerator, the temperature is lower than the room temperature and hence the reaction rate is less.

(iv) Pressure

If the reactants are gases, increasing their pressure increases the reaction rate. This is because, on increasing the pressure the reacting particles come closer and collide frequently.

(v) Catalyst

A catalyst is a substance which increases the reaction rate without being consumed in the reaction. In certain reactions, adding a substance as catalyst speeds up the reaction. For example, on heating potassium chlorate, it decomposes into potassium chloride and oxygen gas, but at a slower rate. If manganese dioxide is added, it increases the reaction rate.

(vi) Surface area of the reactants

When solid reactants are involve in a reaction, their powdered form reacts more readily. For example, powdered calcium carbonate reacts more readily with hydrochloric acid than marble chips. Because, powdering of the reactants increases the surface area and more energy is available on collision of the reactant particles. Thus, the reaction rate is increased.

You will study more about reaction rate in you higher classes.

10.3 STATE OF EQUILIBRIUM

In a reversible reaction, both forward and backward reactions take place simultaneously. When the rate of the forward reaction becomes equal to the rate of backward reaction, then no more product is formed. This stage of the reaction is called **'equilibrium state'**. After this stage, no net change in the reaction can occur and hence in the amount of the reactants and products. Since this equilibrium is attained in a chemical reaction, it is called 'Chemical Equilibrium'. **Chemical Equilibrium: It is state of a reversible chemical reaction in which no change in the amount of the reactants and products takes place. At equilibrium,**

Rate of forward reaction = Rate of backward reaction

Explanation: Initially the rate of the forward reaction is greater than the rate of the backward reaction. However, during the course of reaction, the concentration of the reactants decreases and the concentration of the products increases. Since the rate of a reaction is directly proportional to the concentration, the rate of the forward reaction decreases with time, whereas the rate of the backward reaction increases.

At a certain stage, both the rates become equal. From this point onwards, there will be no change in the concentrations of both the reactants and the products with time. This state is called as equilibrium state.

Let us consider the decomposition of calcium carbonate into lime and carbon dioxide. It is a reversible reaction. The speed of each reaction can be determined by how quickly the reactant disappears. If the reaction is carried out in a closed vessel, it reaches a chemical equilibrium. At this stage,

$$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$$

The rate of decomposition of $CaCO_3$ = The rate of combination of CaO and CO_2

Not only chemical changes, physical changes also may attain equilibrium. When water kept in a closed vessel evaporates, it forms water vapour. No water vapour escapes out of the container as the process takes place in a closed vessel. So, it builds up the vapour pressure in the container. At one time, the water vapour condenses back into liquid water and when the rate of this condensation becomes equal to that of vapourisation, the process attains equilibrium.

At this stage, the volume of the liquid and gaseous phases remain constant. Since it is a physical change, the equilibrium attained is called **'Physical Equilibrium'.** Physical equilibrium is a state of a physical change at which the volume of all the phases remain unchanged.

$$H_2O_{(l)} \underset{\text{Condensation}}{\overset{\text{Evaporation}}{\rightleftharpoons}} H_2O_{(g)}$$

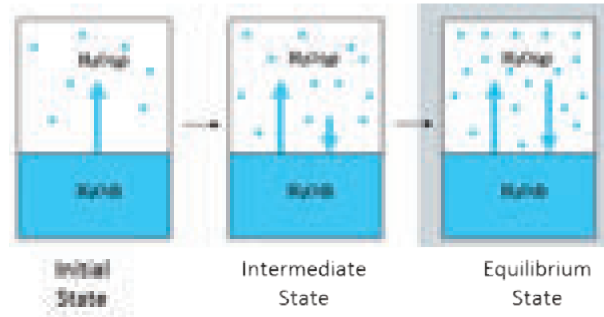


Figure 10.6 State of physical equilibrium

Characteristics of equilibrium

- In a chemical equilibrium, the rates of the forward and backward reactions are equal.
- The observable properties such as pressure, concentration, colour, density, viscosity etc., of the system remain unchanged with time.
- The chemical equilibrium is a dynamic equilibrium, because both the forward and backward reactions continue to occur even though it appears static externally.
- In physical equilibrium, the volume of all the phases remain constant.

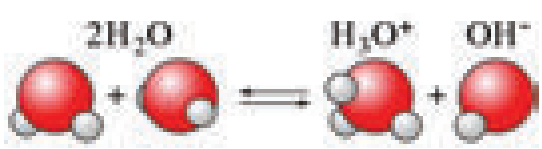
DO YOU KNOW?

Aerated soft drinks contain dissolved carbon dioxide in a pop bottle (Soda). When the bottle is sealed, the dissolved carbon dioxide (in the form of carbonic acid) and gaseous CO_2 are in equilibrium with each other. When you open the bottle, the gaseous CO_2 can escape. So, the dissolved CO_2 begins to undissolve back to the gas phase trying to replace the gas that was lost, when you opened the bottle. That's why if you leave it open long enough, it will goes 'flat'. All the CO_2 will be gone, blown away in the air.

10.4 IONIC PRODUCT OF WATER

Although pure water is often considered as a non-conductor of electricity, precise measurements show that it conducts electricity to a little extent. This conductivity of water has resulted from the self-ionisation of water. Self-ionisation or auto ionisation is a reaction in which two like molecules react to give ions. In the process of ionisation of water, a proton from one water molecule is transferred to another water molecule leaving behind an OH^- ion. The proton gets dissolved in water forming the hydronium ion as shown in the following equation:

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$



The hydronium ion formed is a strong acid and the hydroxyl ion is a strong base. So as fast as they are formed, they react again to produce water. Thus, it is a reversible reaction and attains equilibrium very quickly. So, the extent of ionisation is very little and the concentration of the ions produced is also very less. The product of the concentration of the hydronium ion and the hydroxyl ion is called **'ionic product of water'**. It is denoted as **'K_w'**. It is mathematically expressed as follows:

$$K_w = [H_3O^+] [OH^-]$$

$[H_3O^+]$ may be simply written as $[H^+]$. Thus the ionic product of water may also be expressed as

$$K_w = [H^+] [OH^-]$$

Its unit is $mol^2\ dm^{-6}$. At 25° C, its value is 1.00×10^{-14}.

10.5 pH SCALE

All the aqueous solutions may contain hydrogen and hydroxyl ions due to self-ionisation of water. In addition to this ionisation, substances dissolved in water also may produce hydrogen ions or hydroxyl ions. The concentration of these ions decides whether the solution is acidic or basic. pH scale is a scale for measuring the hydrogen ion concentration in a solution. The 'p' in pH stands for 'Potenz' in German meaning 'power'. pH notation was devised by the Danish biochemist Sorensen in 1909. pH scale is a set of numbers from 0 to 14 which is used to indicate whether a solution is acidic, basic or neutral.

- ✓ Acids have pH less than 7
- ✓ Bases have pH greater than 7
- ✓ A neutral solution has pH equal to 7

The pH is the negative logarithm of the hydrogen ion concentration

i.e $pH = -\log_{10}[H^+]$

COMMON ACIDS	pH	COMMON BASES	pH
HCl (4%)	0	Blood plasma	7.4
Stomach acid	1	Egg white	8
Lemon juice	2	Sea water	8
Vinegar	3	Baking soda	9
Oranges	3.5	Antacids	10
Soda, grapes	4	Ammonia water	11
Sour milk	4.5	Lime water	12
Fresh milk	5	Drain cleaner	13
Human saliva	6-8	Caustic soda 4% (NaOH)	14
Pure water	7	Milk of magnesia	10
Tomato juice	4.2	Coffee	5.6

How can we measure the pH of a given solution using pH Paper

The pH of a solution can be determined by using a universal indicator. It contains a mixture of dyes. It comes in the form of a solution or a pH paper.

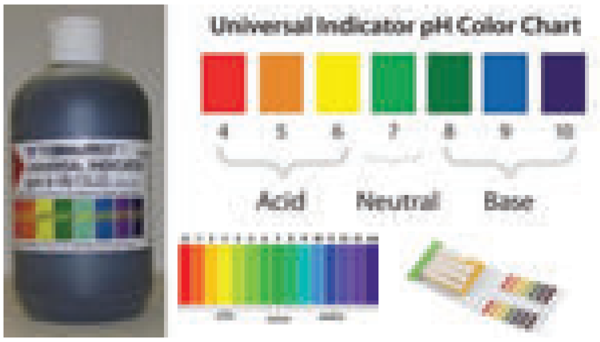


Figure 10.7 pH Indicator

A more common method of measuring pH in a school laboratory is by using the pH paper. A pH paper contains a mixture of indicators. It shows a specific colour at a given pH. A colour guide is provided with the bottle of the indicator or the strips of paper impregnated with it, which are called pH paper strips. The test solution is tested with a drop of the universal indicator, or a drop of the test solution is put on the pH paper. The colour of the solution on the pH paper is compared with the colour chart and the pH value is read from it. The pH values thus obtained are only approximate values.

10.6 ROLE OF pH IN EVERYDAY LIFE

Are plants and animals pH sensitive?

Our body works within the pH range of 7.0 to 7.8. Living organisms can survive only in a narrow range of pH change. Different body fluids have different pH values. For example, pH of blood is ranging from 7.35 to 7.45. Any increase or decrease in this value leads to diseases. The ideal pH for blood is 7.4.

pH in our digestive system

It is very interesting to note that our stomach produces hydrochloric acid. It helps in the digestion of food without harming the stomach. During indigestion the stomach produces too much acid and this causes pain and irritation. pH of the stomach fluid is approximately 2.0.

pH changes as the cause of tooth decay

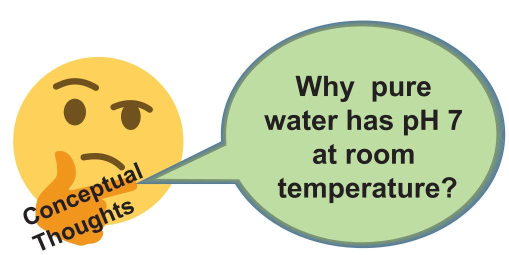
pH of the saliva normally ranges between 6.5 to 7.5. White enamel coating of our teeth is calcium phosphate, the hardest substance in our body. When the pH of the mouth saliva falls below 5.5, the enamel gets weathered. Toothpastes, which are generally basic are used for cleaning the teeth that can neutralise the excess acid and prevent tooth decay.

pH of soil

In agriculture, the pH of the soil is very important. Citrus fruits require slightly alkaline soil, while rice requires acidic soil and sugarcane requires neutral soil.

pH of rain water

The pH of rain water is approximately 7, which means that it is neutral and also represents its high purity. If the atmospheric air is polluted with oxide gases of sulphur and nitrogen, they get dissolved in the rain water and make its pH less than 7. Thus, if the pH of rain water is less than 7, then it is called acid rain. When acid rain flows into the rivers it lowers the pH of the river water also.



The survival of aquatic life in such rivers becomes difficult.

10.7 pH CALCULATION

The pH is the negative logarithm of the hydrogen ion concentration

$$\mathbf{pH = -log_{10} [H^+]}$$

Example: Calculate the pH of 0.01 M HNO_3?

Solution:

$[H^+] = 0.01$

$pH = -\log_{10} [H^+]$

$pH = -\log_{10} [0.01]$

$pH = -\log_{10} [1 \times 10^{-2}]$

$pH = -(\log_{10}1 - 2 \log_{10}10)$

$pH = 0 + 2 \times \log_{10}10$

$pH = 0 + 2 \times 1 = 2$

$pH = 2$

pOH: The pOH of an aqueous solution is realted to the pH.

The pOH is the negative logarithm of the hydroxyl ion concentration

$$\mathbf{pOH = -log_{10}[OH^-]}$$

Example: The hydroxyl ion concentration of a solution is 1×10^{-9}M. What is the pOH of the solution?

Solution

$pOH = -\log_{10} [OH^-]$

$pOH = -\log_{10} [1 \times 10^{-9}]$

$pOH = -(\log_{10} 1.0 + \log_{10} 10^{-9})$

$pOH = -(0-9 \log_{10}10)$

$pOH = -(0 - 9)$

$pOH = 9$

Relationship between pH and pOH

The pH and pOH of a water solution at 25°C are related by the following equation.

$$\mathbf{pH + pOH = 14}$$

If either the pH or the pOH of a solution is known, the other value can be calculated.

Example: A solution has a pOH of 11.76. What is the pH of this solution?

$$pH = 14 - pOH$$
$$pH = 14 - 11.76 = 2.24$$

10.8 PROBLEMS

Example 1: Calculate the pH of 0.001 molar solution of HCl.

Solution: HCl is a strong acid and is completely dissociated in its solutions according to the process:

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

From this process it is clear that one mole of HCl would give one mole of H^+ ions. Therefore, the concentration of H^+ ions would be equal to that of HCl, i.e., 0.001 molar or 1.0×10^{-3} mol $litre^{-1}$.

Thus, $[H^+] = 1 \times 10^{-3}$ mol $litre^{-1}$

$pH = -\log_{10}[H^+] = -\log_{10}10^{-3}$

$= -(-3 \times \log_{10}) = -(3 \times 1) = 3$

Thus, $pH = 3$

Example 2: What would be the pH of an aqueous solution of sulphuric acid which is 5×10^{-5} mol $litre^{-1}$ in concentration.

Solution: Sulphuric acid dissociates in water as:

$$H_2SO_{4(aq)} \rightarrow 2\,H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Each mole of sulphuric acid gives two mole of H^+ ions in the solution. One litre of H_2SO_4 solution contains 5×10^{-5} moles of H_2SO_4 which would give $2 \times 5 \times 10^{-5} = 10 \times 10^{-5}$ or 1.0×10^{-4} moles of H^+ ion in one litre of the solution.

Therefore,

$[H^+] = 1.0 \times 10^{-4}$ mol $litre^{-1}$

$pH = -\log_{10}[H^+] = -\log_{10}10^{-4} = -(-4 \times \log_{10}10)$

$= -(-4 \times 1) = 4$

Example 3: Calculate the pH of 1×10^{-4} molar solution of NaOH.

Solution: NaOH is a strong base and dissociates in its solution as:

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

One mole of NaOH would give one mole of OH^- ions. Therefore,

$[OH^-] = 1 \times 10^{-4}$ mol litre^{-1}

$pOH = -\log_{10}[OH^-] = -\log_{10} \times [10^{-4}]$

$= -(-4 \times \log_{10}10) = -(-4) = 4$

Since, $pH + pOH = 14$

$pH = 14 - pOH = 14 - 4$

$= 10$

Example 4: Calculate the pH of a solution in which the concentration of the hydrogen ions is 1.0×10^{-8} mol litre^{-1}.

Solution: Here, although the solution is extremely dilute, the concentration given is not of an acid or a base but that of H^+ ions. Hence, the pH can be calculated from the relation:

$pH = -\log_{10}[H^+]$

given $[H^+] = 1.0 \times 10^{-8}$ mol litre^{-1}

$pH = -\log_{10}10^{-8} = -(-8 \times \log_{10}10)$

$= -(-8 \times 1) = 8$

Example 5: If the pH of a solution is 4.5, what is its pOH?

Solution:

$pH + pOH = 14$

$pOH = 14 - 4.5 = 9.5$

$pOH = 9.5$

Points to Remember

- A chemical change is a change in which one or more new substances are formed.
- Most combination reactions are exothermic
- All photo decomposition reaction are endothermic reactions.
- Double displacement reaction or metathesis may occur by the mutal exchange of ions.
- Precipitation reaction gives an insoluble salt as the product.
- Neutralisation reactions are reactions between an acid and a base that forms salt and water.
- Neutralisation prevents tooth decay.
- Most reactions in chemistry are irreversible reactions.
- Chemical equilibrium-the rate of the forward reaction is equal to rate of the back ward reactions.
- Equilibrium is possible in a closed system.
- Temperature increases the reaction rate.
- Pressure increases the reaction rate.
- The term pH means power of hydrogen.
- pH plays a vital role in everyday life.
- In humans all bio chemical reactions take place between the pH value of 7.0 to 7.8.
- If pH of rain water is below 5.6 its called acid rain.
- Pure water is a weak electrolyte.

TEXTBOOK EVALUATION

I. Choose the correct answer.

1. $H_{2(g)} + Cl_{29(g)} \rightarrow 2HCl_{(g)}$ is a
   a. Decomposition Reaction
   b. Combination Reaction
   c. Single Displacement Reaction
   d. Double Displacement Reaction

2. Photolysis is a decomposition reaction caused by __________
   a. heat b. electricity
   c. light d. mechanical energy

3. A reaction between carbon and oxygen is represented by $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ + Heat. In which of the type(s), the above reaction can be classified?
   (i) Combination Reaction
   (ii) Combustion Reaction
   (iii) Decomposition Reaction
   (iv) Irreversible Reaction
   a. i and ii b. i and iv
   c. i, ii and iii d. i, ii and iv

HFPH6A

4. The chemical equation
   $Na_2SO_{4(aq)} + BaCl_{2(aq)} \rightarrow BaSO_{4(s)}\downarrow + 2NaCl_{(aq)}$
   represents which of the following types of reaction?
   a. Neutralisation
   b. Combustion
   c. Precipitation
   d. Single displacement

5. Which of the following statements are correct about a chemical equilibrium?
   (i) It is dynamic in nature
   (ii) The rate of the forward and backward reactions are equal at equilibrium
   (iii) Irreversible reactions do not attain chemical equilibrium
   (iv) The concentration of reactants and products may be different
   a. i, ii and iii b. i, ii and iv
   c. ii, iii and iv d. i, iii and iv

6. A single displacement reaction is represented by $X_{(s)} + 2HCl_{(aq)} \rightarrow XCl_{2(aq)} + H_{2(g)}$. Which of the following(s) could be X.
   (i) Zn (ii) Ag (iii) Cu (iv) Mg.
   Choose the best pair.
   a. i and ii b. ii and iii
   c. iii and iv d. i and iv

7. Which of the following is not an "element + element → compound" type reaction?
   a. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$
   b. $2K_{(s)} + Br_{2(l)} \rightarrow 2KBr_{(s)}$
   c. $2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$
   d. $4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$

8. Which of the following represents a precipitation reaction?
   a. $A_{(s)} + B_{(s)} \rightarrow C_{(s)} + D_{(s)}$
   b. $A_{(s)} + B_{(aq)} \rightarrow C_{(aq)} + D_{(l)}$
   c. $A_{(aq)} + B_{(aq)} \rightarrow C_{(s)} + D_{(aq)}$
   d. $A_{(aq)} + B_{(s)} \rightarrow C_{(aq)} + D_{(l)}$

9. The pH of a solution is 3. Its $[OH^-]$ concentration is
   a. 1×10^{-3} M
   b. 3 M
   c. 1×10^{-11} M
   d. 11 M

10. Powdered $CaCO_3$ reacts more rapidly than flaky $CaCO_3$ because of __________.
    a. large surface area
    b. high pressure
    c. high concentration
    d. high temperature

II. Fill in the blanks

1. A reaction between an acid and a base is called ______________.
2. When lithium metal is placed in hydrochloric acid, ________ gas is evolved.

3. The equilibrium attained during the melting of ice is known as ___________.
4. The pH of a fruit juice is 5.6. If you add slaked lime to this juice, its pH ___________ (increse/decrese)
5. The value of ionic product of water at 25^0 C is ___________.
6. The normal pH of human blood is ___________
7. Electrolysis is type of ___________ reaction
8. The number of products formed in a synthesis reaction is ___________
9. Chemical volcano is an example for ___________ type of reaction
10. The ion formed by dissolution of H^+ in water is called ___________

III. Match the following

1. Identify the types of reaction

REACTION	TYPE
$NH_4OH_{(aq)} + CH_3COOH_{(aq)} \rightarrow CH_3COONH_{4(aq)} + H_2O_{(l)}$	Single Displacement
$Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$	Combustion
$ZnCO_{3(s)} + \xrightarrow{Heat} ZnO_{(s)} + CO_{2(g)}$	Neutralisation
$C_2H_{4(g)} + 4O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(g)} + Heat$	Thermal decomposition

IV. True or False: (If false give the correct statement)

1. Silver metal can displace hydrogen gas from nitric acid.
2. The pH of rain water containing dissolved gases like SO_3, CO_2, NO_2 will be less than 7.
3. At the equilibrium of a reversible reaction, the concentration of the reactants and the products will be equal.
4. Periodical removal of one of the products of a reversible reaction increases the yield.
5. On dipping a pH paper in a solution, it turns into yellow. Then the solution is basic.

V. Short answer questions:

1. When an aqueous solution of potassium chloride is added to an aqueous solution of silver nitrate, a white precipitate is formed. Give the chemical equation of this reaction.
2. Why does the reaction rate of a reaction increase on raising the temperature?
3. Define combination reaction. Give one example for an exothermic combination reaction.
4. Differentiate reversible and irreversible reactions

VI. Answer in detail

1. What are called thermolysis reactions?
2. Explain the types of double displacement reactions with examples.
3. Explain the factors influencing the rate of a reaction
4. How does pH play an important role in everyday life?
5. What is a chemical equilibrium? What are its characteristics?

VII. HOT questions

1. A solid compound 'A' decomposes on heating into 'B' and a gas 'C'. On passing the gas 'C' through water, it becomes acidic. Identify A, B and C.

2. Can a nickel spatula be used to stir copper sulphate solution? Justify your answer.

VIII. Solve the following problems

1. Lemon juice has a pH 2, what is the concentration of H^+ ions?
2. Calculate the pH of 1.0×10^{-4} molar solution of HNO_3.
3. What is the pH of 1.0×10^{-5} molar solution of KOH?
4. The hydroxide ion concentration of a solution is 1×10^{-11}M. What is the pH of the solution?

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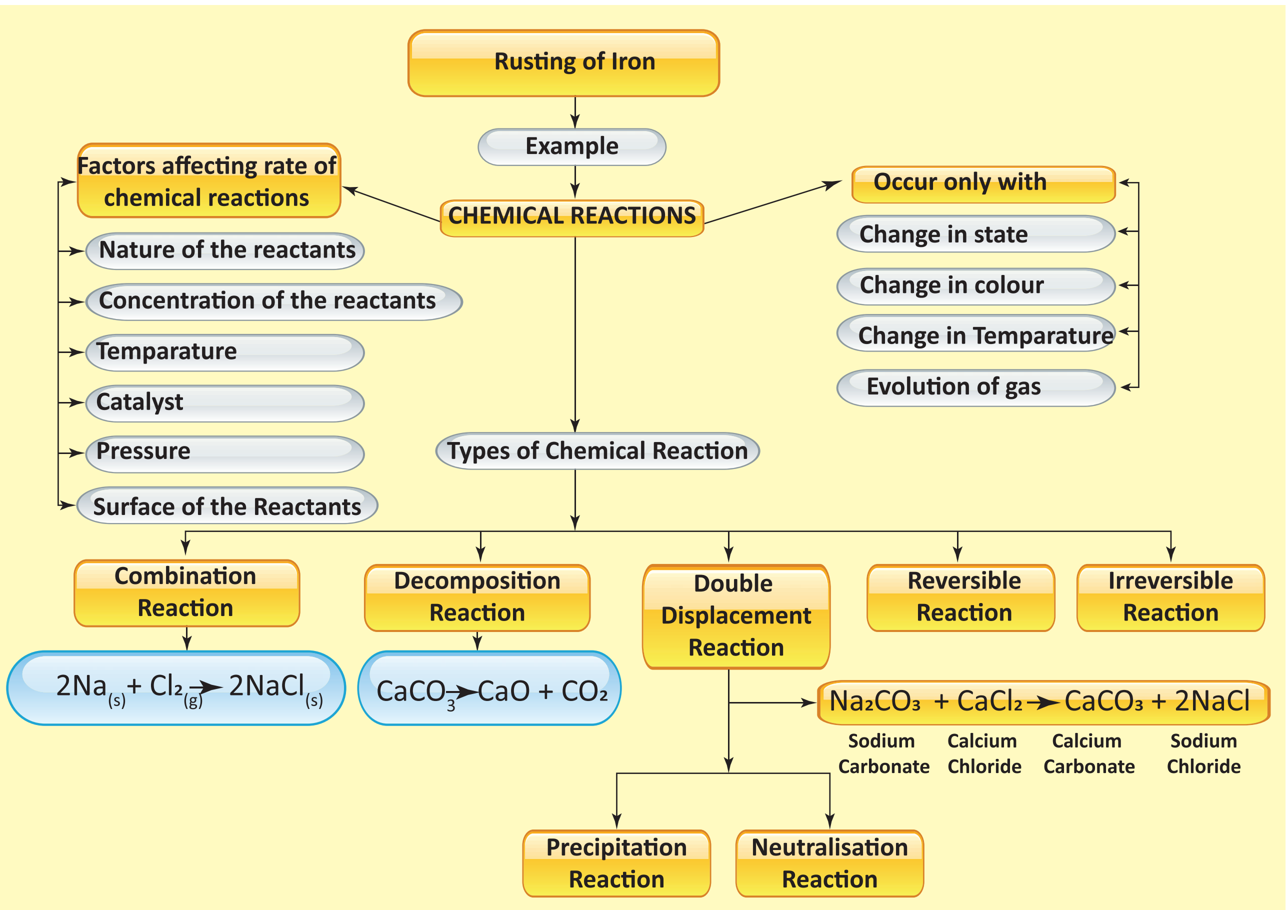
INTERNET RESOURCES

Webliography:www.chem4kids.com

http://aravindguptatoys.com/filims.html

Concept Map



UNIT 11

CARBON AND ITS COMPOUNDS

HV8C7L

Learning Objectives

After studying this lesson, the student will be able to:

- know the importance of organic compounds.
- classify the organic compounds and name them based on IUPAC rules.
- identify the functional groups of organic compounds.
- explain the preparation, properties and uses of ethanol and ethanoic acids.
- know the composition and preparation of soap and detergent.
- understand the cleansing action of soap and detergents.
- differentiate soap and detergents.

INTRODUCTION

You have studied, in your lower classes, that carbon is an inseparable element in human life as we use innumerable number of carbon compounds in our day to day life. Because, the food we eat, medicines we take when ill, clothes we wear; domestic and automobile fuels, paint, cosmetics, automobile parts, etc., that we use contain carbon compounds. The number of carbon compounds found in nature and man-made, is much higher than that of any other element in the periodic table. Infact there are more than 5 million compounds of carbon. The unique nature of carbon, such as catenation, tetravalency and multiple bonding, enables it to combine with itself or other elements like hydrogen, oxygen, nitrogen,sulphur etc., and hence form large number of compounds. All these compounds are made of covalent bonds. These compounds are called **organic compounds.** In this lesson, you will learn about carbon compounds.

11.1 GENERAL CHARACTERISTICS OF ORGANIC COMPOUNDS

Everything in this world has unique character, similarly organic compounds are unique in their characteristics. Some of them are given below:

- Organic compounds have a high molecular weight and a complex structure.
- They are mostly insoluble in water, but soluble in organic solvents such as ether, carbon tetrachloride, toluene, etc.
- They are highly inflammable in nature
- Organic compounds are less reactive compared to inorganic compounds. Hence, the reactions involving organic compounds proceed at slower rates.

- Mostly organic compounds form covalent bonds in nature.
- They have lower melting point and boiling point when compared to inorganic compounds
- They exhibit the phenomenon of isomerism, in which a single molecular formula represents several organic compounds that differ in their physical and chemical properties
- They are volatile in nature.
- Organic compounds can be prepared in the laboratory

11.2 CLASSIFICATION OF ORGANIC COMPOUNDS BASED ON THE PATTERN OF CARBON CHAIN

What is the significance of classification? There are millions of organic compounds known and many new organic compounds are discovered every year in nature or synthesized in laboratory. This may mystify organic chemistry to a large extent. However, a unique molecular structure can be assigned to each compound and it can be listed by using systematic methods of classification and eventually named on the basis of its structural arrangements. In early days, chemists recognised that compounds having similar structural features have identical chemical properties. So they began to classify compounds based on the common structural arrangements found among them.

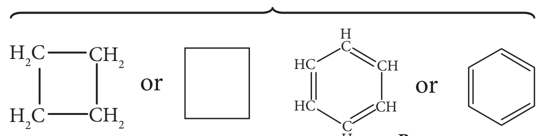
Organic chemistry is the chemistry of catenated carbon compounds. The carbon atoms present in organic compounds are linked with each other through covalent bonds and thus exist as chains. By this way, organic compounds are classified into two types as follows:

1. Acyclic or Open chain compounds: These are the compounds in which the carbon atoms are linked in a linear pattern to form the chain. If all the carbon atoms in the chain are connected by single bonds, the compound is called as **saturated.** If one or more double bonds or triple bonds exist between the carbon atoms, then the compound is said to **unsaturated**.

CH_3-CH_2-CH_3	CH_3-CH=CH_2
Propane	Propene
Saturated compound	Unsaturated compound

2. Cyclic Compounds: Organic compounds in which the chain of carbon atoms is closed or cyclic are called **cyclic compounds.** If the chain contains only carbon atoms, such compounds are called **carbocyclic** (Homocyclic) **compounds.** If the chain contains carbon and other atoms like oxygen, nitrogen, sulphur, etc., these compounds are called **heterocyclic compounds.** Carbocyclic compounds are further subdivided into **alicyclic** and **aromatic compounds.** Alicyclic compounds contain one or more carbocyclic rings which may be saturated or unsaturated whereas aromatic compounds contain one or more benzene rings (ring containing alternate double bonds between carbon atoms). E.g.

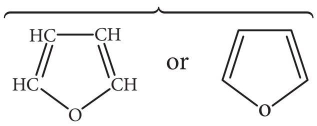
Carbo (Homo) cyclic compounds



Cyclobutane
Alicyclic Compound

Benzene
Aromatic Compound

Hetrocyclic compounds



Furan

Figure 11.1 depicts the classification of organic compounds based on the pattern of carbon arrangements and their bonding in organic compounds:

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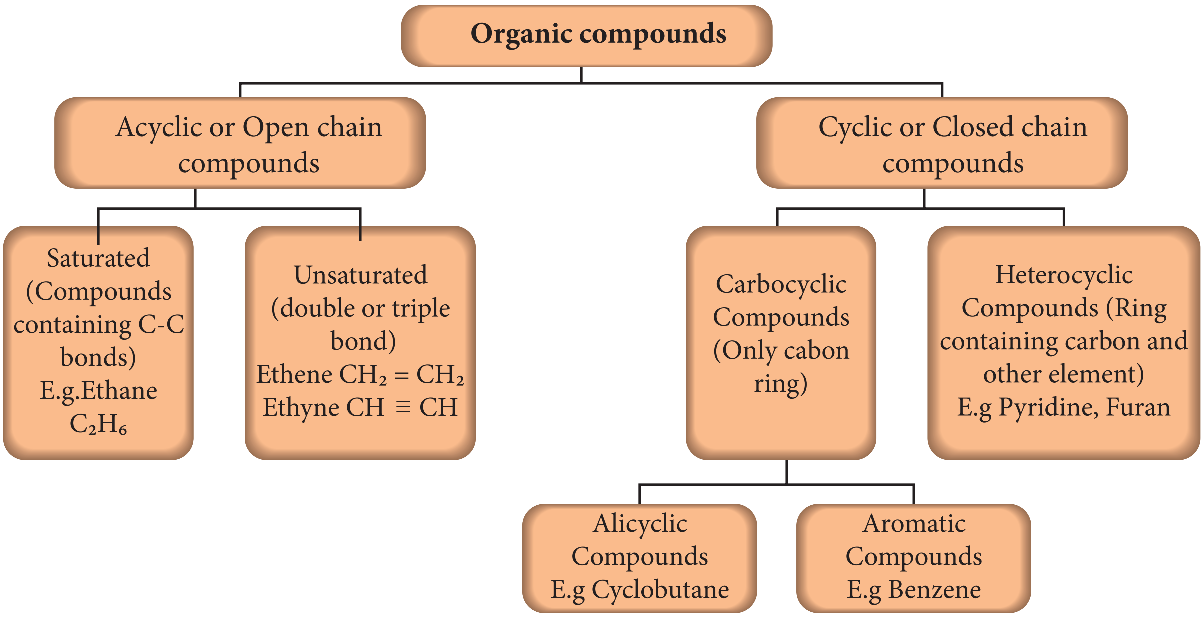

Figure 11.1 Classification of organic compounds

11.3 CLASSES OF ORGANIC COMPOUNDS (BASED ON THE KIND OF ATOMS)

Other than carbon, organic compounds contain atoms like hydrogen, oxygen, nitrogen, etc., bonded to the carbon. Combination of these kinds of atoms with carbon gives different classes of organic compounds. In the following section, let us discuss various classes of organic compounds.

11.3.1 Hydrocarbons

The organic compounds that are composed of only carbon and hydrogen atoms are called **hydrocarbons.** The carbon atoms join together to form the framework of the compounds. These are regarded as the parent organic compounds and all other compounds are considered to be derived from hydrocarbons by replacing one or more hydrogen atoms with other atoms or group of atoms. Hydrocarbons are, further, sub divided into three classes such as:

(a) Alkanes: These are hydrocarbons, which contain only single bonds. They are represented by the general formula $\mathbf{C_nH_{2n+2}}$ (where n = 1,2, 3, ……). The simplest alkane (for n=1) is methane (CH_4). Since, all are single bonds in alkanes, they are saturated compounds.

(b) Alkenes: The hydrocarbons, which contain one or more C=C bonds are called alkenes. These are unsaturated compounds. They are represented by the general formula $\mathbf{C_nH_{2n}}$. The simplest alkene contains two carbon atoms (n=2) and is called ethylene (C_2H_4).

(c) Alkynes: The hydrocarbons containing carbon to carbon triple bond are called **alkynes.** They are also unsaturated as they contain triple bond between carbon atoms. They have the general formula $\mathbf{C_nH_{2n-2}}$. Acetylene (C_2H_2) is the simplest alkyne, which contains two carbon atoms. Table 11.1 lists the first five hydrocarbons of each class:

Table 11.1 **Hydrocarbons containing 1to 5 carbon atoms**

No. of carbon atoms	Alkane (C_nH_{2n+2})	Alkene (C_nH_{2n})	Alkyne (C_nH_{2n-2})
1	Methane (CH_4)	-	-
2	Ethane (C_2H_6)	Ethene (C_2H_4)	Ethyne (C_2H_2)
3	Propane (C_3H_8)	Propene (C_3H_6)	Propyne (C_3H_4)
4	Butane (C_4H_{10})	Butene (C_4H_8)	Butyne (C_4H_6)
5	Pentane (C_5H_{12})	Pentene (C_5H_{10})	Pentyne (C_5H_8)

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11.3.2 Characteristics of hydrocarbons:

- Lower hydrocarbons are gases at room temperature E.g. methane, ethane are gases.
- They are colourless and odourless.
- The boiling point of hydrocarbons increases with an increase in the number of carbon atoms.
- They undergo combustion reaction with oxygen to form CO_2 and water.
- Alkanes are least reactive when compared to other classes of hydrocarbons.
- Alkynes are the most reactive due to the presence of the triple bond.
- Alkanes are saturated whereas alkenes and alkynes are unsaturated.
- They are insoluble in water.

Test to identify saturated and unsaturated compounds:

- Take the given sample solution in a test tube.
- Add a few drops of bromine water and observe any characteristic change in colour.
- If the given compound is unsaturated, it will decolourise bromine water.
- Saturated compounds do not decolourise bromine.

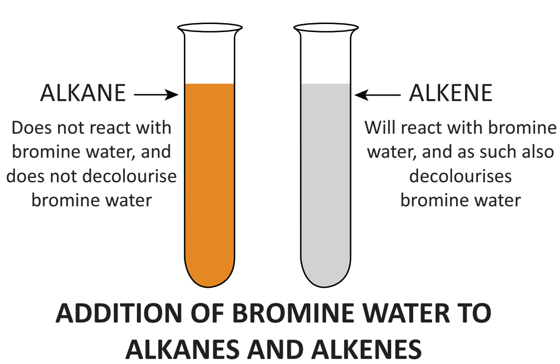


Figure 11.2 Test to identify unsaturated compounds

11.3.3 Classification of organic compounds based on functional groups

The structural frameworks of organic compounds are made of carbon and hydrogen, which are relatively less reactive. But, the presence of some other atoms or group of atoms makes the compounds more reactive and thus determines the chemical properties of the compound. These groups are called **functional groups.**

A functional group is an atom or group of atoms in a molecule, which gives its characteristic chemical properties.

The chemical properties of an organic compound depend on its functional group whereas its physical properties rely on remaining part of the structure. Carbon to carbon multiple bonds (C=C, C≡C) also are considered as functional groups as many of the properties are influenced by these bonds. Other functional groups include, –OH, –CHO, –COOH, etc.

For example, ethane is a hydrocarbon having molecular formula C_2H_6. If one of its hydrogen is replaced by –OH group, you will get an alcohol. Leaving the functional group, the rest of the structure is represented by 'R'. Thus an alcohol is represented by 'R-OH'

CH_3-CH_3	CH_3-CH_2-OH ← Functional Group
Ethane	Ethanol
Hydrocarbon	Alcohol

A series of compounds containing the same functional group is called a **class of organic compounds.** Table 11.2 shows various classes or families of organic compounds and their functional groups:

Table 11.2 **Classes of organic compounds based on functional group**

Class of the compound	Functional group	Common Formula	Examples
Alcohol	–OH	R-OH	Ethanol, CH_3CH_2OH
Aldehyde	O ‖ –C–H	R-CHO	Acetaldehyde, CH_3CHO
Ketone	O ‖ –C–	R–CO-R	Acetone, CH_3COCH_3
Carboxylic acid	O ‖ –C–OH	R-COOH	Acetic acid, CH_3COOH
Ester	O ‖ –C–OR	R-COOR	Methyl acetate, CH_3COOCH_3
Ether	–O–R	R-O-R	Dimethyl ether, CH_3OCH_3

11.4 HOMOLOGOUS SERIES

Homologous series is a group or a class of organic compounds having same general formula and similar chemical properties in which the successive members differ by a -CH_2 group.

Let us consider members of alkanes given in Table 11.1. Their condensed structural formulas are given below:

Methane	-	CH_4
Ethane	-	CH_3CH_3
Propane	-	$CH_3CH_2CH_3$
Butane	-	$CH_3(CH_2)_2CH_3$
Pentane	-	$CH_3(CH_2)_3CH_3$

If you observe the above series. you can notice that each successive member has one methylene group more than the precedent member of the series and hence they are called homologs.

11.4.1 Characteristics of homologous series

- Each member of the series differs from the preceding or succeeding member by one methylene group (–CH_2) and hence by a molecular mass of 14 amu.
- All members of a homologous series contain the same elements and functional group.
- They are represented by a general molecular formula. e.g. Alkanes, C_nH_{2n+2}.
- The members in each homologous series show a regular gradation in their physical properties with respect to their increase in molecular mass.
- Chemical properties of the members of a homologous series are similar.
- All the members can be prepared by a common method.

11.5 NOMENCLATURE OF ORGANIC COMPOUNDS

11.5.1 Why do we need nomenclature?

In ancient days, the names of organic compounds were related to the natural things from which they were obtained. For example, the formic acid was initially obtained by distillation of 'red ants'. Latin name of the red ant is 'Formica'. So, the name of the formic acid was derived from the Latin name of its source Later, the organic compounds were synthesized from sources other than the natural sources. So scientists framed a systematic method for naming the organic compounds based on their structures. Hence, a set of rules was formulated by IUPAC (**International Union of Pure and Applied Chemistry)** for the nomenclature of chemical compounds.

11.5.2 Components of an IUPAC name

The IUPAC name of the any organic compound consists of three parts:

i. Root word
ii. Prefix
iii. Suffix

These parts are combined as per the following sequence to get the IUPAC name of the compound:

Prefix + Root Word + Suffix ⟶ IUPAC Name

(i) Root word: It is the basic unit, which describes the carbon skeleton. It gives the number of carbon atoms present in the parent chain of the compound and the pattern of their arrangement. Based on the number of carbon atoms present in the carbon skeleton, most of the names are derived from Greek numerals (except the first four). Table 11.3 shows the root words for the parent chain of hydrocarbons containing 1to10 carbon atoms:

Table 11.3 Root words of hydrocarbons

No. of carbon atoms	Root word
1	Meth-
2	Eth-
3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-
9	Non-
10	Dec-

(ii) Prefix: The prefix represents the substituents or branch present in the parent chain. Atoms or group of atoms, other than hydrogen, attached to carbon of the parent chain are called substituents. Table 11.4 presents the major substituents of organic compounds and respective prefix used for them:

Table 11.4 Prefix for IUPAC Name

Substituent	Prefix used
-F	Fluoro
-Cl	Chloro
-Br	Bromo
-I	Iodo
$-NH_2$	Amino
$-CH_3$	Methyl
$-CH_2CH_3$	Ethyl

(iii) Suffix

The suffix forms the end of the name. It is divided into two parts such as **(a) Primary suffix** and **(b) Secondary suffix.** The primary suffix comes after the root word. **It represents the nature in carbon to carbon bonding of the parent chain**. If all the bonds between the carbon atoms of the parent chain are single, then suffix **'ane'** has to be used. Suffix **'ene'** and **'yne'** are used for the compounds containing double and triple bonds respectively. **The secondary suffix describes the functional group of the compound**.

Table 11.5 Suffix for IUPAC Name

Class of the Compound	Functional group	Suffix used
Alcohol	$-OH$	-ol
Aldehyde	$-CHO$	-al
Ketone	$\begin{array}{c} O \\ \Vert \\ -C- \end{array}$	-one
Carboxylic acid	$-COOH$	-oic acid

11.5.3 IUPAC rules for naming organic compounds:

- **Rule1:** Identify the longest chain of carbon atoms to get the parent name (root word).
- **Rule 2:** Number the carbon atoms of the parent chain, **beginning at the closest end of the substituent or functional group**. These are called **locant numbers**. *If both functional group and substituent are present, then the priority will be given to the functional group.*
- **Rule 3:** In case of alkenes and alkynes, locate the double bond or triple bond and use its locant number followed by a dash and a primary suffix. The carbon chain is numbered in such a way that the multiple bonds have the lowest possible locant number.
- **Rule 4:** If the compound contains functional group, locate it and use its locant number followed by a dash and a secondary suffix.
- **Rule 5:** When the primary and secondary suffixes are joined, the terminal 'e' of the primary suffix is removed.
- **Rule 6:** Identify the substituent and use a number followed by a dash and a prefix to specify its location and identity.

11.5.4 IUPAC Nomenclature of hydrocarbons – Solved examples

Let us try to name, systematically, some of the linear and substituted hydrocarbons by following IUPAC rules:

Example 1: $CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}CH_3$
Step 1: It is a five- carbon chain and hence the root word is 'Pent'. (Rule 1)
Step 2: All the bonds between carbon atoms are single bonds, and thus the suffix is 'ane'.
So, its name is **Pent + ane = Pentane**

Example 2:

$$\begin{array}{ccccccccc} & & CH_3 & & & & & & \\ & & | & & & & & & \\ CH_3 & - & CH & - & CH_2 & - & CH_2 & - & CH_3 \end{array}$$

Step 1: The longest chain contains five carbon atoms and hence the root word is 'Pent'.

Step 2: There is a substituent. So, the carbon chain is numbered from the left end, which is closest to the substituent. (Rule 2)

$$\begin{array}{ccccccccc} & & CH_3 & & & & & & \\ 1 & & |2 & & 3 & & 4 & & 5 \\ CH_3 & - & CH & - & CH_2 & - & CH_2 & - & CH_3 \end{array}$$

Step 3: All are single bonds between the carbon atoms and thus the suffix is 'ane'.

Step 4: The substituent is a methyl group and it is located at second carbon atom. So, its locant number is 2. Thus the prefix is '2-Methyl'. (Rule 6).

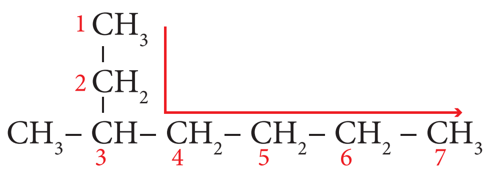
The name of the compound is
2-Methyl + pent +ane = 2-Methylpentane

Example 3:

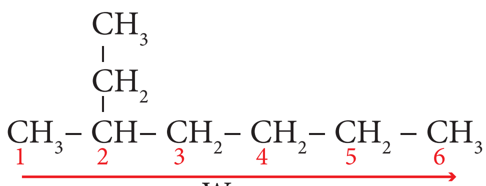
$$\begin{array}{ccccccccccc} & & CH_3 & & & & & & & & \\ & & | & & & & & & & & \\ & & CH_2 & & & & & & & & \\ & & | & & & & & & & & \\ CH_3 & - & CH & - & CH_2 & - & CH_2 & - & CH_2 & - & CH_3 \end{array}$$

Step 1: The longest chain contains seven carbon atoms and hence the root word is 'Hept'.

Step 2: There is a substituent. So, the carbon chain is numbered from the end, which is closest to substituent. (Rule 2)



Correct



Wrong

Step 3: All are single bonds between the carbon atoms and thus the suffix is 'ane'.

Step 4: The substituent is a methyl group and it is located at third carbon. So, its locant number is 3. Thus the prefix is '3-Methyl'. (Rule 6)

Hence the name of the compound is **3-Methyl + hept + ane = 3 –Methylheptane**

Example 4: CH_3-CH_2-CH_2-CH=CH_2

Step 1: It is a 'five- carbon atoms chain' and hence the root word is 'Pent'. (Rule 1)

Step 2: There is a carbon to carbon double bond. The suffix is 'ene'.

Step 3: The carbon chain is numbered from the end such that double bond has the lowest locant number as shown below: (Rule 3):

$$\overset{5}{CH_3}-\overset{4}{CH_2}-\overset{3}{CH_2}-\overset{2}{CH}=\overset{1}{CH_2}$$

Step 4: The locant number of the double bond is 1 and thus the suffix is '-1-ene'.

So, the name of the compound is **Pent + (-1-ene) = Pent-1-ene**

11.5.5 IUPAC Nomenclature of other classes – Solved examples

Example 1: CH_3-CH_2-CH_2-OH

Step1: The parent chain consists of 3 carbon atoms. The root word is 'Prop'.

Step 2: There are single bonds between the carbon atoms of the chain. So, the **primary suffix** is 'ane'.

Step 3: Since, the compound contains –OH group, it is an alcohol. The carbon chain is numbered from the end which is closest to –OH group. (Rule 3)

$$\overset{3}{CH_3}-\overset{2}{CH_2}-\overset{1}{CH_2}-OH$$

Step 4: The locant number of –OH group is 1 and thus the secondary suffix is '1-ol'.

The name of the compound is **Prop + ane + (1-ol) = Propan-1-ol**

Note: Terminal 'e' of 'ane' is removed as per Rule 5

Example 2: CH_3COOH

Step1: The parent chain consists of 2 carbon atoms. The root word is 'Eth'.

Step 2: All are single bonds between the carbon atoms of the chain. So the primary suffix is 'ane'.

Step 3: Since the compound contains the–COOH group, it is a carboxylic acid. The secondary suffix is 'oic acid'

The name of the compound is **Eth + ane + oic acid) = Ethanoic acid**

Table 11.6 lists IUPAC names homologs of various classes of organic compounds

Test yourself:

Obtain the IUPAC name of the following compounds systematically:

(a) CH_3CHO

(b) $CH_3CH_2COCH_3$

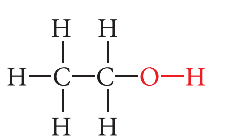
(c) ClCH_2-CH_2-CH_2-CH_3

Table 11.6 IUPAC Name of various classes of compounds

No. of carbons atoms	IUPAC Name			
	Alcohols	Aldehydes	Ketones	Carboxylic acid
1	Methanol (CH_3OH)	Methanal ($HCHO$)	-	Methanoic acid ($HCOOH$)
2	Ethanol (CH_3CH_2OH)	Ethanal (CH_3CHO)	-	Ethanoic acid (CH_3COOH)
3	Propanol ($CH_3CH_2CH_2OH$)	Propanal (CH_3CH_2CHO)	Propanone (CH_3COCH_3)	Propanoic acid (CH_3CH_2COOH)
4	Butanol ($CH_3CH_2CH_2CH_2OH$)	Butanal ($CH_3CH_2CH_2CHO$)	Butanone ($CH_3COCH_2CH_3$)	Butanoic acid ($CH_3CH_2CH_2COOH$)
5	Pentanol ($CH_3CH_2CH_2CH_2CH_2OH$)	Pentanal ($CH_3CH_2CH_2CH_2CHO$)	Pentanone ($CH_3COCH_2CH_2CH_3$)	Pentanoic acid ($CH_3CH_2CH_2CH_2COOH$)

11.6 ETHANOL (CH_3CH_2OH)

Ethanol is commonly known as alcohol. All alcoholic beverages and some cough syrups contain ethanol. Its molecular formula is C_2H_5OH. Its structural formula is



Ethanol (C_2H_5OH)

IAR78X

11.6.1 Manufacture of ethanol

Ethanol is manufactured in industries by the fermentation of molasses, which is a by-product obtained during the manufacture of sugar from sugarcane. Molasses is a dark coloured syrupy liquid left after the crystallization of sugar from the concentrated sugarcane juice. Molasses contain about 30% of sucrose, which cannot be separated by crystallization. It is converted into ethanol by the following steps:

(i) Dilution of molasses

Molasses is first diluted with water to bring down the concentration of sugar to about 8 to 10 percent.

(ii) Addition of Nitrogen source

Molasses usually contains enough nitrogenous matter to act as food for yeast during the fermentation process. If the nitrogen content of the molasses is poor, it may be fortified by the addition of **ammonium sulphate** or **ammonium phosphate**.

(iii) Addition of Yeast

The solution obtained in step (ii) is collected in large 'fermentation tanks' and yeast is added to it. The mixture is kept at about 303K for a few days. During this period, the enzymes invertase and zymase present in yeast, bring about the conversion of sucrose into ethanol.

$$\underset{\text{Sugar}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{invertase}} \underset{\text{glucose}}{C_6H_{12}O_6} + \underset{\text{fructose}}{C_6H_{12}O_6}$$

$$\underset{\text{glucose or fructose}}{C_6H_{12}O_6} \xrightarrow{\text{zymase}} \underset{\text{ethanol}}{2C_2H_5OH} + 2\,CO_2$$

The fermented liquid is technically called **wash**.

(iv) Distillation of 'Wash'

The fermented liquid (i.e. wash), containing 15 to 18 percent alcohol, is now subjected to fractional distillation. The main fraction drawn is an aqueous solution of ethanol which contains

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95.5% of ethanol and 4.5% of water. This is called **rectified spirit**. This mixture is then refluxed over quicklime for about 5 to 6 hours and then allowed to stand for 12 hours. On distillation of this mixture, pure alcohol (100%) is obtained. This is called **absolute alcohol**.

More to know

Yeast and Fermentation: Yeasts are single–celled microorganisms, belonging to the class of fungi. The enzymes present in yeasts catalyse many complex organic reactions. Fermentation is conversion of complex organic molecules into simpler molecules by the action of enzymes. E.g. Curdling of milk

11.6.2 Physical properties

i) Ethanol is a colourless liquid, having a pleasant smell and a burning taste.

ii) It is a volatile liquid. Its boiling point is 78° C (351K), which is much higher than that of its corresponding alkane, i.e. ethane (Boiling Point = 184 K).

iii) It is completely miscible with water in all proportions.

11.6.3 Chemical Properties

(i) Dehydration (Loss of water)

When ethanol is heated with con H_2SO_4 at 443K, it loses a water molecule i.e. dehydrated to form ethene.

$$\underset{\text{Ethanol}}{CH_3CH_2OH} \xrightarrow[443K]{Conc.H_2SO_4} \underset{\text{Ethene}}{CH_2 = CH_2} + H_2O$$

(ii) Reaction with sodium:

Ethanol reacts with sodium metal to form sodium ethoxide and hydrogen gas.

$$2C_2H_5OH + 2Na \longrightarrow \underset{\text{sodium ethoxide}}{2C_2H_5ONa} + H_2 \uparrow$$

(iii) Oxidation:

Ethanol is oxidized to ethanoic acid with alkaline $KMnO_4$ or acidified $K_2Cr_2O_7$

$$CH_3CH_2OH \xrightarrow[2[O]]{K_2Cr_2O_7 / H^+} \underset{\text{Ethanoic acid}}{CH_3COOH} + H_2O$$

During this reaction, the orange colour of $K_2Cr_2O_7$ changes to green. Therefore, this reaction can be used for the identification of alcohols.

(iv) Esterification:

The reaction of an alcohol with a carboxylic acid gives a compound having fruity odour. This compound is called an **ester** and the reaction is called esterification. Ethanol reacts with ethanoic acid in the presence of conc. H_2SO_4 to form ethyl ethanoate, an ester.

$$\underset{\text{Ethanol}}{C_2H_5OH} + \underset{\text{Ethanoic acid}}{CH_3COOH} \xrightarrow{conc.H_2SO_4} \underset{\text{Ethyl ethanoate}}{CH_3COOC_2H_5} + H_2O$$

(v) Dehydrogenation:

When the vapour of ethanol is passed over heated copper, used as a catalyst at 573 K, it is dehydrogenated to acetaldehyde.

$$\underset{\text{Ethanol}}{CH_3CH_2OH} \xrightarrow[573K]{Cu} \underset{\text{Acetadehyde}}{CH_3CHO} + H_2\uparrow$$

(vi) Combustion:

Ethanol is highly inflammable liquid. It burns with oxygen to form carbon dioxide and water.

$$\underset{\text{Ethanol}}{C_2H_5OH} + 3O_2 \longrightarrow \underset{\text{Carbon dioxide}}{2CO_2} + 3H_2O$$

11.6.4 Uses of ethanol

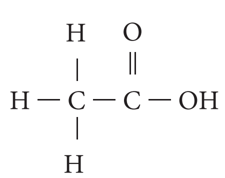
Ethanol is used

- in medical wipes, as an antiseptic.
- as an anti-freeze in automobile radiators.
- for effectively killing micro organisms like bacteria, fungi, etc., by including it in many hand sanitizers.

- as an antiseptic to sterilize wounds in hospitals.
- as a solvent for drugs, oils, fats, perfumes, dyes, etc.
- in the preparation of **methylated spirit** (mixture of 95% of ethanol and 5% of methanol) **rectified spirit** (mixture of 95.5% of ethanol and 4.5% of water), power alcohol (mixture of petrol and ethanol) and denatured spirit (ethanol mixed with pyridine).
- to enhance the flavour of food extracts, for example vanilla extract; a common food flavour, which is made by processing vanilla beans in a solution of ethanol and water.

11.7 ETHANOIC ACID (CH_3COOH)

Ethanoic acid or acetic acid is one of the most important members of the carboxylic acid family. Its molecular formula is $C_2H_4O_2$. Its structural formula is



11.7.1 Manufacture of ethanoic acid

Ethanoic acid is prepared in large scale, by the oxidation of ethanol in the presence of alkaline potassium permanganate or acidified potassium dichromate.

$$\underset{\text{Ethanol}}{CH_3CH_2OH} \xrightarrow[2[O]]{KMnO_4/\,OH^-} \underset{\text{Ethanoic acid}}{CH_3COOH} + H_2O$$

11.7.2 Physical Properties

(i) Ethanoic acid is a colourless liquid having an unpleasant odour.

(ii) It is sour in taste.

(iii) It is miscible with water in all proportions.

(iv) Its boiling point is higher than the corresponding alcohols, aldehydes and ketones.

(v) On cooling, pure ethanoic acid is frozen to form ice like flakes. They look like glaciers, so it is called **glacial acetic acid.**

11.7.3 Chemical Properties

(i) **Reaction with metal:** Ethanoic acid reacts with active metals like Na, Zn, etc., to liberate hydrogen and form their ethanoate.

$$2CH_3COOH + Zn \rightarrow (CH_3COO)_2\,Zn + H_2\uparrow$$

$$2CH_3COOH + 2Na \rightarrow 2CH_3COONa + H_2\uparrow$$

(ii) **Reaction with carbonates and bicarbonates:** Ethanoic acid reacts with sodium carbonate and sodium bicarbonate, which are weaker bases and liberates CO_2, with brisk effervescence.

$$2CH_3COOH + Na_2CO_3 \rightarrow 2CH_3COONa + CO_2\uparrow + H_2O$$

$$CH_3COOH + NaHCO_3 \rightarrow CH_3COONa + CO_2\uparrow + H_2O$$

(iii) **Reaction with base:** Ethanoic acid reacts with sodium hydroxide to form sodium ethanoate and water.

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

(iv) **Decarboxylation (Removal of CO_2):**

When a sodium salt of ethanoic acid is heated with soda lime (solid mixure of 3 parts of NaOH and 1 part of CaO), methane gas is formed.

$$CH_3COONa \xrightarrow{NaOH\,/\,CaO} CH_4\uparrow + Na_2CO_3$$

11.7.4 Uses of ethanoic acid

Acetic acid, in lower concentration, (vinegar) is used as a food additive, a flavoring agent and a preservative.

Ethanoic acid is used

- in the manufacture of plastic.
- in making dyes, pigments and paint.
- in printing on fabrics.
- as a laboratory reagent.
- for coagulating rubber from latex.
- in the production of pharmaceuticals.

11.8 ORGANIC COMPOUNDS IN DAILY LIFE

Organic compounds are inseparable in human life. They are used by mankind or associated at all stages of life right from one's birth to death. Various classes of organic compounds and their uses in our daily life as follows:

Hydrocarbons

- Fuels like LPG, Petrol, Kerosene.
- Raw materials for various important synthetic materials.
- Polymeric materials like tyre, plastic containers.

Alcohols

- As a solvent and an antiseptic agent.
- Raw materials for various important synthetic materials.

Aldehydes

- Formaldehyde as a disinfectant.
- Raw materials for synthetic materials.

Ketones

- As a solvent.
- Stain Remover.

Ethers

- Anaesthetic agents.
- Pain Killer.

Esters

- All the cooking oils and lipids contain esters.

11.9 SOAPS AND DETERGENTS

Soaps and the Detergents are materials that are used by us for cleaning purposes because pure water alone cannot remove all types of dirt or any oily substance from our body or clothes. They contain 'surfactants', which are compounds with molecules that line up around water to break the 'surface tension'. Both of them having a different chemical nature. ***Soap*** is a cleaning agent that is composed of one or more salts of fatty acids. ***Detergent*** is a chemical compound or a mixture of chemical compounds, which is used as a cleaning agent, also. They perform their cleaning actions in certain specific conditions. You will learn more about this in detail, in the following units.

11.9.1 Soap

Soaps are sodium or potassium salts of some long chain carboxylic acids, called fatty acids. Soap requires two major raw materials: i) fat and ii) alkali. The alkali, most commonly used in the preparation of soap is sodium hydroxide. Potassium hydroxide can also be used. A potassium-based soap creates a more water-soluble product than a sodium-based soap. Based on these features, there are two types of soaps:

A. HARD SOAP

Soaps, which are prepared by the ***saponification of oils or fats with caustic soda*** (sodium hydroxide), are known as hard soaps. They are usually used for washing purposes.

B. SOFT SOAP

Soaps, which are prepared by the ***saponification of oils or fats with potassium salts***, are known as soft soaps. They are used for cleansing the body.

Manufacture of soap

KETTLE PROCESS:

This is the oldest method. But, it is still widely used in the small scale preparation of soap. There are mainly, two steps to be followed in this process.

i) Saponification of oil:

The oil, which is used in this process, is taken in an iron tank (kettle). The alkaline solution (10%) is added into the kettle, a little in excess. The mixture is boiled by passing steam through it. The oil gets hydrolysed after several hours of boiling. This process is called Saponification

ii) Salting out of soap:

Common salt is then added to the boiling mixture. Soap is finally precipitated in the tank. After several hours the soap rises to the top of the liquid as a 'curdy mass'. The neat soap is taken off from the top. It is then allowed to cool down.

Effect of hard water on soap

Hard water contains calcium and magnesium ions (Ca^{2+} and Mg^{2+}) that limit the cleaning action of soap. When combined with soap, hard water develops a thin layer (precipitates of the metal ions) called 'scum', which leaves a deposit on the clothes or skin and does not easily rinse away. Over time, this can lead to the deterioration of the fabric and eventually ruin the clothes. On the other hand, detergents are made with chemicals that are not affected by hard water.

DO YOU KNOW?

Why ordinary soap is not suitable for using with hard water?

Ordinary soaps when treated with hard water, precipitate as salts of calcium and magnesium. **They appear at the surface of the cloth as sticky grey scum.** Thus, the soaps cannot be used conveniently in hard water.

11.9.2 Detergents

Development of synthetic detergents is a big achievement in the field of cleansing. These soaps possess the desirable properties of ordinary soaps and also can be used with hard water and in acidic solutions. These are salts of sulphonic acids or alkyl hydrogen sulphates in comparison to soap, which are salts of carboxylic acids. The detergents do not form precipitates with Ca^{2+} and Mg^{2+} present in hard water. So, the cleansing action of detergents is better than that of soaps.

Preparation of detergents

Detergents are prepared by adding sulphuric acid to the processed hydrocarbon obtained from petroleum. This chemical reaction result in the formation of molecules similar to the fatty acid in soap. Then, an alkali is added to the mixture to produce the 'surfactant molecules', which do not bond with the minerals present in the hard water, thus preventing the formation of their precipitates.

In addition to a 'surfactant', the modern detergent contains several other ingredients. They are listed as follows:

i) Sodium silicate, which prevents the corrosion and ensures that the detergent does not damage the washing machine.

ii) Fluorescent whitening agents that give a glow to the clothes.

iii) Oxygen bleaches, such as 'sodium perborate', enable the removal of certain stains from the cloth.

iv) Sodium sulphate is added to prevent the caking of the detergent powder.

v) Enzymes are added to break down some stains caused by biological substances like blood and vegetable juice.

vi) Certain chemicals that give out a pleasant smell are also added to make the clothes fragrant after they are washed with detergents.

11.9.3 Cleansing action of soap

A soap molecule contains two chemically distinct parts that interact differently with water. It has one **polar end**, which is a ***short head*** with a carboxylate group (–COONa) and one **non-polar end** having the ***long tail made of the hydrocarbon chain***.

The polar end is ***hydrophilic (Water loving)*** in nature and this end is attracted towards water. The non-polar end is ***hydrophobic (Water hating)*** in nature and it is attracted towards dirt or oil on the cloth, but not attracted towards water. Thus, the hydrophobic part of the soap molecule traps the dirt and the hydrophilic part makes the entire molecule soluble in water.

When a soap or detergent is dissolved in water, the molecules join together as clusters called '**micelles**'. Their long hydrocarbon chains attach themselves to the oil and dirt. The dirt is thus surrounded by the non-polar end of the soap molecules (Figure 11.3). The charged carboxylate end of the soap molecules makes the micelles soluble in water. Thus, the dirt is washed away with the soap.

Advantages of detergents over soaps

Detergents are better than soaps because they:

- can be used in both hard and soft water and can clean more effectively in hard water than soap.
- can also be used in saline and acidic water.
- do not leave any soap scum on the tub or clothes.
- dissolve freely even in cool water and rinse freely in hard water.
- can be used for washing woollen garments, where as soap cannot be used.
- have a linear hydrocarbon chain, which is biodegradable.
- are active emulsifiers of motor grease.
- do an effective and safe cleansing, keeping even synthetic fabrics brighter and whiter.

Biodegradable and Non-biodegradable detergents:

a) Biodegradable detergents:

They have straight hydrocarbon chains, which can be easily degraded by bacteria.

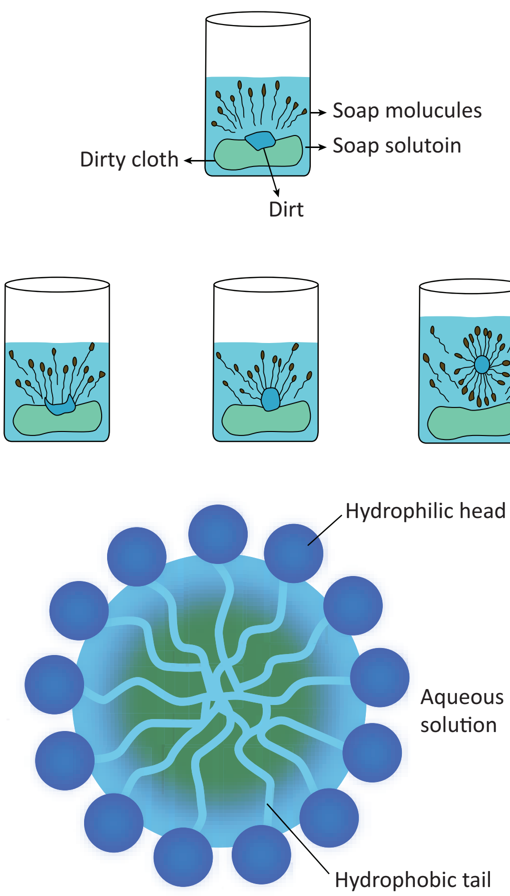


Figure 11.3 Cleansing action of soap

b) Non-biodegradable detergents:

They have highly branched hydrocarbon chains, which cannot be degraded by bacteria.

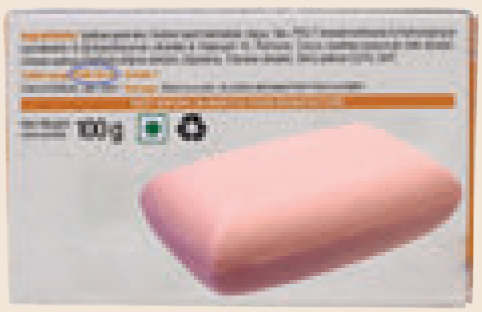
Disadvantages of Detergents

1. Some detergents having a branched hydrocarbon chain are not fully biodegradable by micro-organisms present in water. So, they cause water pollution.
2. They are relatively more expensive than soap.

11.9.4 Comparison between soap and detergents

Soap	Detergent
It is a sodium salt of long chain fatty acids.	It is sodium salts of sulphonic acids.
The ionic part of a soap is $-COO^-Na^+$.	The ionic part in a detergent is $-SO_3^-Na^+$.
It is prepared from animal fats or vegetable oils.	It is prepared from hydrocarbons obtained from crude oil.
Its effectiveness is reduced when used in hard water.	It is effective even in hard water.
It forms a scum in hard water.	Does not form a scum in hard water.
It has poor foaming capacity.	It has rich foaming capacity.
Soaps are biodegradable.	Most of the detergents are non-biodegradable.

Have you noticed the term "TFM" in soap



TFM means TOTAL FATTY MATTER. It is the one of the important factors to be considered to assess the quality of soap. A soap, which has higher TFM, is a good bathing soap.

Points to Remember

- A group or class of organic compounds related to each other by a general molecular formula constitutes homologous series.
- The IUPAC name of the any organic compound consist of three parts. **ROOTWORD, PREFIX and / or SUFFIX**.
- Functional group may be defined as an atom or group of atom or reactive part which is responsible for the characteristic properties of the compounds
- Ethanoic acid is most commonly known as acetic acid and belongs to a group of acids called carboxylic acids.
- Ethanol or ethyl alcohol or simply alcohol is one of the most important members of the family of alcohols.
- The slow chemical change that takes place in complex organic compounds by the action of enzymes leading to the formation of simple molecules is called fermentation.
- Soaps are sodium or potassium salts of some long chain carboxylic acids.
- Detergents are sodium salts of sulphonic acids. Thus instead of $-COOH$ group in soaps, detergents contain $-SO_3H$ group

IQA2A9

TEXTBOOK EVALUATION

I. Choose the best answer.

1. The molecular formula of an open chain organic compound is C_3H_6. The class of the compound is
   a. alkane b. alkene
   c. alkyne d. alcohol

2. The IUPAC name of an organic compound is 3-Methyl butan-1-ol. What type compound it is?
   a. Aldehyde b. Carboxylic acid
   c. Ketone d. Alcohol

3. The secondary suffix used in IUPAC nomenclature of an aldehyde is ____
   a. - ol b. – oic acid
   c. - al d. - one

4. Which of the following pairs can be the successive members of a homologous series?
   a. C_3H_8 and C_4H_{10}
   b. C_2H_2 and C_2H_4
   c. CH_4 and C_3H_6
   d. C_2H_5OH and C_4H_8OH

5. $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$ is a
   a. Reduction of ethanol
   b. Combustion of ethanol
   c. Oxidation of ethanoic acid
   d. Oxidation of ethanal

6. Rectified spirit is an aqueous solution which contains about ________ of ethanol
   a. 95.5 % b. 75.5 %
   c. 55.5 % d. 45.5 %

7. Which of the following are used as anaesthetics?
   a. Carboxylic acids b. Ethers
   c. Esters d. Aldehydes

8. TFM in soaps represents ____________ content in soap
   a. mineral b. vitamin
   c. fatty acid d. carbohydrate

9. Which of the following statements is wrong about detergents?
   a. It is a sodium salt of long chain fatty acids
   b. It is sodium salts of sulphonic acids
   c. The ionic part in a detergent is $-SO_3^-Na^+$
   d. It is effective even in hard water.

II. Fill in the blanks

1. An atom or a group of atoms which is responsible for chemical characteristics of an organic compound is called _____________.
2. The general molecular formula of alkynes is _______________
3. In IUPAC name, the carbon skeleton of a compound is represented by __________ (root word / prefix / suffix)
4. (Saturated / Unsaturated) ____________ compounds decolourize bromine water.
5. Dehydration of ethanol by conc. Sulphuric acid forms ________ (ethene/ ethane)
6. 100 % pure ethanol is called _____________
7. Ethanoic acid turns ________ litmus to __________
8. The alkaline hydrolysis of fatty acids is termed as _______________
9. Biodegradable detergents are made of __________(branched / straight) chain hydrocarbons

III. Match the following

Functional group –OH	-	Benzene
Heterocyclic	-	Potassium stearate
Unsaturated	-	Alcohol
Soap	-	Furan
Carbocyclic	-	Ethene

IV. Assertion and Reason:

Answer the following questions using the data given below:

i) A and R are correct, R explains the A.

ii) A is correct, R is wrong.

iii) A is wrong, R is correct.

iv) A and R are correct, R doesn't explains A.

1. **Assertion**: Detergents are more effective cleansing agents than soaps in hard water.

   Reason: Calcium and magnesium salts of detergents are water soluble.

2. **Assertion**: Alkanes are saturated hydrocarbons.

   Reason: Hydrocarbons consist of covalent bonds.

V. Short answer questions

1. Name the simplest ketone and give its structural formula.
2. Classify the following compounds based on the pattern of carbon chain and give their structural formula: (i) Propane (ii) Benzene (iii) Cyclobutane (iv) Furan
3. How is ethanoic acid prepared from ethanol? Give the chemical equation.
4. How do detergents cause water pollution? Suggest remedial measures to prevent this pollution?
5. Differentiate soaps and detergents.

VI. Long answer questions

1. What is called homologous series? Give any three of its characteristics?
2. Arrive at, systematically, the IUPAC name of the compound: $CH_3–CH_2–CH_2–OH$.
3. How is ethanol manufactured from sugarcane?
4. Give the balanced chemical equation of the following reactions:

   (i) Neutralization of NaOH with ethanoic acid.

   (ii) Evolution of carbon dioxide by the action of ethanoic acid with $NaHCO_3$.

   (iii) Oxidation of ethanol by acidified potassium dichromate.

   (iv) Combustion of ethanol.
5. Explain the mechanism of cleansing action of soap.

VII. HOT questions

1. The molecular formula of an alcohol is $C_4H_{10}O$. The locant number of its –OH group is 2.

   (i) Draw its structural formula.

   (ii) Give its IUPAC name.

   (iii) Is it saturated or unsaturated?
2. An organic compound 'A' is widely used as a preservative and has the molecular formula $C_2H_4O_2$. This compound reacts with ethanol to form a sweet smelling compound 'B'.

   (i) Identify the compound 'A'.

   (ii) Write the chemical equation for its reaction with ethanol to form compound 'B'.

   (iii) Name the process.

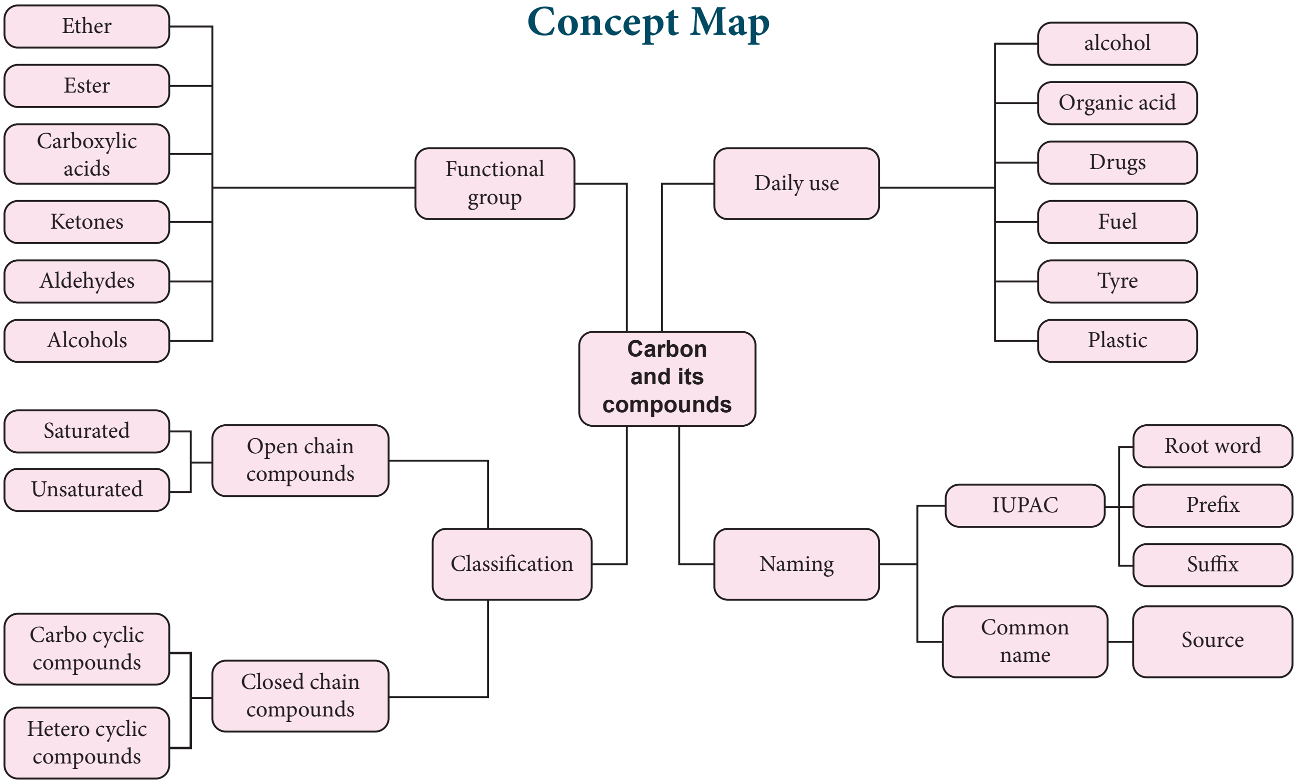
REFERENCE BOOKS

1. Organic chemistry - B.S.Bahl & Arun Bahl S.Chand publishers, New delhi.
2. Organic chemistry - R.T.Morrision & R.MN. Boyd - Prentice Hall Publishers. New Delhi

INTERNET RESOURCES

https://www.tutorvista.com/
https://www.topperlearning.com/
http://www.chem4kids.com/

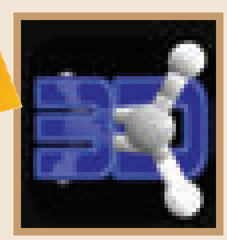
Concept Map



ICT CORNER

Organic Carbon and its Compounds

The students can know about the Hydrocarbons, its formulae, descriptions and also in animated form.



Steps

- Type the URL or use the QR code to access the application "Organic chemistry visualized / OCV" and install in the mobile.
- Click 'Content' to view the list of Hydrocarbon. Explore about the hydrocarbons by clicking those.
- On clicking '**anim**' button, you can view the formulae of hydrocarbon as well as in animated form.
- Click '**Quiz**' to evaluate yourself in this topic.

Step1 **Step2** **Step3** **Step4**

URL: https://play.google.com/store/apps/details?id=com.budgietainment.oc

or Scan the QR Code.

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*Pictures are indicative only