Reading Material for

Applied Sciences – I

(Physics & Chemistry)

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PREFACE

A two years post matric teaching program of <u>Medical Laboratory Technician</u> for the students of Allied Health Sciences. The purpose of this reading material is to provide basic education to the paramedics about <u>physics and chemistry</u>. This reading material attempts to cover almost all the basic theoretical knowledge required by students about <u>physics and chemistry.</u> so that they can perform their work better.

Contents

Chapter # 1

The Nature of Science

Students will be able to understand:

- The nature of Science
- Divisions of Science
- Scientific method

The wonders of nature have always inspired man. He has always been curious to know the secrets of nature and remained in search of the truth and reality. He observes various phenomena and tries to find their answers by logical reasoning. The knowledge gained through observations and experimentation is called **Science**. The word science is derived from the Latin word scientia, which means knowledge. Not until the eighteenth century, various aspects of material objects were studied under a single subject called natural philosophy. But as the knowledge increased, it was divided into two main streams.

Physical sciences: Which deal with the study of non-living things

Biological sciences: which are concerned with the study of living things.

Measurements are not confined to science. They are part of our lives. They play an important role in describing and understanding the physical world. Over the centuries, man has improved the methods of measurement.

1.1 Introduction to physics:

In the nineteenth century, physical sciences were divided into five distinct disciplines.

Physics   Chemistry   Astronomy   Geology   Meteorology

1.1.1 Frontiers of Fundamental Science:

At present, there are three main frontiers of fundamental science.

i. The world of the extremely large, the universe itself, Radio telescopes now gather information from the far side of the universe and have recently detected, as radio waves, the "fire light" of the big bang which probably started the expanding universe nearly 20 billion years ago.

ii. The world of extremely small, that of particles such as electrons, protons, neutrons, mesons and others.

iii. The world of complex matter, it is also the world of "middle-sized" things, from molecules at one extreme to the earth at the other. This is all fundamental physics, which is the heart of science.

The most fundamental of these is Physics. In Physics, we study matter, energy and their interaction. The laws and principles of physics help us to understand nature. The rapid progress in science during recent years has become possible due to the discoveries and inventions in the field of Physics. The technologies are the applications of scientific principles. Most of the technologies of our modern society throughout the world are related to Physics. For example, a car is made on the principles of mechanics and a refrigerator is based on the principles of thermodynamics. In our daily life, we hardly find a device where Physics is not involved. Consider pulleys that make it easy to lift heavy loads.Similarly, the means of communication such as radio, TV, telephone, and computer are the result of applications of Physics. These devices have made our lives much easier, faster, and more comfortable than in the past. For example, think of what a mobile phone smaller than our palm can do. It allows us to contact people anywhere in the world and to get the latest worldwide information. We can take and save pictures and send and receive messages from our friends. We can also receive radio transmissions and can use them as calculators. However, scientific inventions have also caused harm and destruction of serious nature. One of which is environmental pollution and environmental pollution, and the other is deadly weapons.

1.2 BRANCHES OF PHYSICS

Mechanics: It is the study of motion of objects, its causes and effects.

Heat: It deals with the nature of heat, modes of transfer and effects of heat.

Sound: It deals with the physical aspects of sound waves, their production, properties and applications. Light (Optics): It is the study of physical aspects of light, its properties, working and use of optical instruments.

Electricity and Magnetism: It is the study of the charges at rest and in motion, their effects and their relationship with magnetism.

Atomic Physics: It is the study of the structure and properties of atoms.

Nuclear Physics: It deals with the properties and behavior of nuclei and the particles within the nuclei. Plasma Physics: It is the study of production and properties of the ionic state of matter - the fourth state of matter.

Geophysics: It is the study of the internal structure of the Earth.

1.3 Scientific method

''The scientific method is the process of objectively establishing facts through testing and experimentation.''

The Scientific method is a process with the help of which scientists try to investigate, verify, or construct an accurate and reliable version of any natural phenomena. They are done by creating an objective framework for scientific inquiry and analyzing the results scientifically to conclude that either supports or contradicts the observation made at the beginning.

1.3.1 Steps of Scientific Method:

The aim of all scientific methods is the same, that is, to analyse the observation made at the beginning. Still, various steps are adopted per the requirement of any given observation. However, there is a generally accepted sequence of steps in scientific methods.

Question → Collect data → Test hypothesis → Conclusion

1.   **Observation and formulation of a question**: This is the first step of a scientific method. To start, an observation must be made into any observable aspect or phenomena of the universe, and a question needs to be asked about that aspect. For example, you can ask, "Why is the sky black at night? or "Why is air invisible?"

2.   **Data Collection and Hypothesis**: The next step involved in the scientific method is to collect all related data and formulate a hypothesis based on the observation. The Hypothesis could be the cause of the phenomena, its effect, or its relation to any other phenomena.

3.   **Testing the hypothesis**: After the hypothesis is made, it needs to be tested scientifically. Scientists do this by conducting experiments. These experiments aim to determine whether the hypothesis agrees with or contradicts the observations made in the real world. Confidence in the hypothesis increases or decreases based on the results of the experiments.

4.   **Analysis and Conclusion:** This step involves the use of proper mathematical and other scientific procedures to determine the results of the experiment. Based on the

analysis, the future course of action can be determined. If the data found in the analysis is consistent with the hypothesis, it is accepted. If not, then it is rejected or modified and analyzed again.

It must be remembered that a hypothesis cannot be proved or disproved by doing one experiment. It needs to be done repeatedly until there are no discrepancies in the data and the result. When there are no discrepancies and the hypothesis is proved, it is accepted as a 'theory'.

EXERCISE

Write a detailed description of Scientific Method.

Define the following:

    i.   Physical Sciences
    ii.   Plasma Physics
    iii.   Mechanics
    iv.   Scientific Method

What are the frontiers of Fundamental Sciences?

Chapter # 2

The Measurement

Students will be able to understand:

- Measurements.
- Metric System
- Scientific notation
- Units of mass , length and volume.
- Units conversion

2.1 Introduction:

The foundation of physics rests upon physical quantities in terms of which the laws of physics are expressed. Therefore, these quantities must be measured accurately. Among these are mass, length, time, velocity, force, density, temperature, electric current, and numerous others. Physical quantities are often divided into two categories: base quantities and derived quantities. Derived quantities are those whose definitions are based on other physical

quantities. Velocity, acceleration, force, etc. are usually viewed as derived quantities. Base quantities are not defined in terms of other physical quantities. The base quantities are the minimum number of those physical quantities in terms of which other physical quantities can be defined. Typical examples of base quantities are length, mass, and time.

2.2. International System of Units:

In 1960, an international committee agreed on a set of definitions and standards to describe the physical quantities. The system that was established is called the - System International (SI).

Due to the simplicity and convenience with which the units in this system are amenable to arithmetical manipulation, it is in universal use by the world's scientific community and by most nations. The system international (SI) is built up of three kinds of units:

1) Base units
2) Supplementary units
3) Derived units

2.2.1 Base Units:

There are seven base units for various physical quantities namely: "length, mass, time, temperature, electric current, luminous intensity, and amount of a substance (with special reference to the number of particles). The name of base units for these physical quantities together with symbols. are listed in Table 2.1.

Property	Unit	Abbreviation
Length	meter	m
Mass	kilogram	kg
Time	seconds	s
Amount	mole	mol
Temperature	kelvin	K
electric current	ampere	amp
luminous intensity	candella	cd

Table 2.1 : The names of physical quantities with their symbols.

2.2.2 Supplementary Units:

The General Conference on Weights and Measures has not yet classified certain units of the SI under either base units or derived units. These SI units are called supplementary units. For the time being this class contains only two units of _ purely geometrical quantities, which are the plane angle and the solid angle (Table2.2).

SUPPLEMENTARY UNITS		
Quantity	Unit	Symbol
1. Plane Angle	Radian	rad
2. Solid Angle	Steradian	Sr

Table 2.2 : The names of supplementary units with their symbols.

2.2.2.1 Physical Quantities:
Radian:

The radian is the plane angle between two radii of a circle which cut off on the circumference of an arc, equal in length to the radius, as shown in (Fig. 2.3 a)
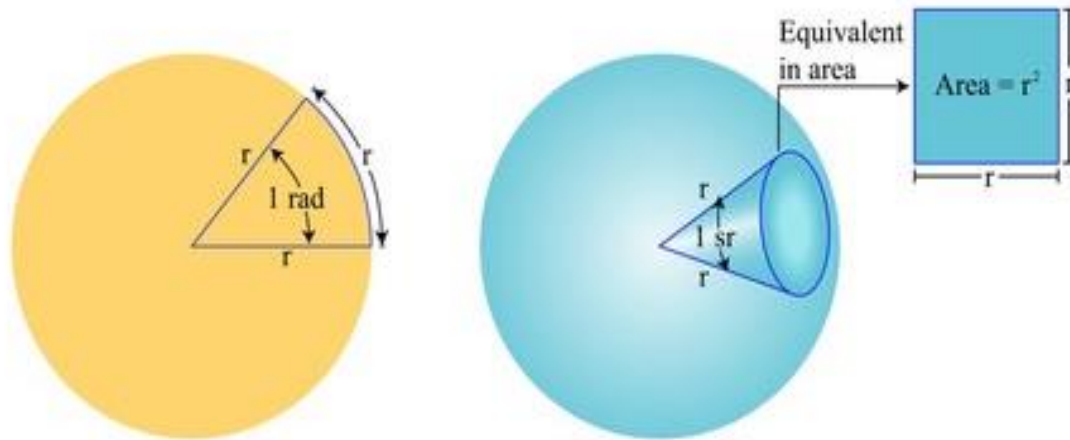
Fig.2.3 Physical quantities (a) Radian (b) Steradian .

Steradian:

The Steradian is the solid angle (three-dimensional angle) subtended at the center of a sphere by, an area of its surface equal to the square of radius of the sphere. (Fig. 1.3 b)

2.2.3 Derived Units:

SI units for__ measuring all other physical quantities are derived from the base and supplementary units. Some of the derived units are given in the table. 1.3.

Quantity	Name	Symbol	Expression
Frequency	Hertz	Hz	$1/s$
Force	Newton	N	$kg \cdot m/s^2$
Pressure, stress	Pascal	Pa	$N/m^2 = kg/m \cdot s^2$
Energy, work	Joule	J	$N \cdot m = kg \cdot m^2/s^2$
Power, radiant flux	Watt	W	$J/s = kg \cdot m^2/s^3$
Electric charge	Coulomb	C	$A \cdot s$
Voltage, electric potential	Volt	V	$W/A = kg \cdot m^2/A \cdot s^3$
Capacitance	Farad	F	$C/V = s^4A^2/m^2kg$
Electric resistance	Ohm	Ω	$V/A = m^2kg/s^3A^2$
Conductance	Siemens or mho	S or Ω	$1/\Omega = s^3A^2/m^2kg$
Magnetic field	Tesla	T	$N/A \cdot m = kg/s^2A$
Magnetic flux	Weber	Wb	$T \cdot m^2 = m^2kg/s^2A$
Inductance	Henry	H	$V \cdot s/A = m^2kg/s^2A^2$

2.3 Scientific notation

Numbers are expressed in standard form called scientific notation, which employs powers of ten. The internationally accepted practice is that there should be only one non-zero digit left of decimal. Thus, the number 134.7 should be written as 1.347×10^2 and 0.0023 should be written as 1347×10^3 and 0.0023 should be expressed as 2.3×10^{-3}.

2.3.1 Conversion for Indicating Units

The use of SI units requires special care, more particularly in writing prefixes.

The following points should be kept in mind while using units.

i. The full name of the unit does not begin with a capital letter even if named after a scientist e.g., Newton.
ii. The symbol of a unit named after a scientist has an initial capital letter such as N for Newton.
iii. The prefix should be written before the unit without any space, such as 1×10^{-3} m is written as 1 mm. Standard prefixes are given in Table 1.4.
iv. A combination of base units is written each with one space apart. For example, Newton meter is written as N m.
v. Compound prefixes are not allowed. For example, 1μμF may be written as 1pF.
vi. A number such as 5.0×10^4 cm may be expressed in scientific notation as 5.0×10^2 m.

vii. When a multiple of a base unit is raised to a power, the power applies to the whole multiple and not the base unit alone. Thus, 1 km2 = 1 $(km)^2$ = 1 x 10^5 m^2.

Measurement in practical work should be recorded immediately in the most convenient unit, e.g., micrometer screw gauge measurement in mm, and the mass of the calorimeter in grams (g). But before calculation for the result, all measurements must be converted to the appropriate SI base units.

Power of 10	Prefix	Symbol	Power of 10	Prefix	Symbol
10^{18}	exa	E	10^{-1}	deci	d
10^{15}	peta	P	10^{-2}	centi	c
10^{12}	tera	T	10^{-3}	milli	m
10^9	giga	G	10^{-6}	micro	μ
10^6	mega	M	10^{-9}	nano	n
10^3	kilo	k	10^{-12}	pico	P
10^2	hecto	h	10^{-15}	femto	f
10^1	deca	da	10^{-18}	atto	a

2.4 Metric System:

Metric System is basically a system used for measuring distance, length, volume, weight, and temperature. It is based on three basic units with which we can measure almost everything in the world.

M- Meter, used to measure the length.

Kg- kilogram, used to measure the mass.

S- Second, used to measure time.

The metric system originated in the year 1799. Over the years, this system has been extended to incorporate many units. Although several variants of this system were discovered in the late 19th and early 20th centuries, the term – metric system is used as another word for SI or the International System of Units.

2.4.1 SI Units Metric System

The SI Units (standard International System of Units) give proper definitions for meter, kilogram, and the second. It also specifies and defines four different additional units:

Kelvin- temperature

Ampere – Electric Current

Candela- luminous Intensity

Mole – Material Quantity

2.4.2 Metric Number Prefixes
To remember the proper movement of units, arrange the prefixes from the largest to the smallest. Thus, the conversion from one unit to the other unit is done by multiplying or dividing the powers of 10.

Some of the units for length, weight, volume, and time are discussed below.

3.5.2.2. **Length:**
The most common units used to measure the length are as follows:

Kilo	hecto	deca	Unit	deci	centi	milli
10^3	10^2	10^1	10^0	10^{-1}	10^{-2}	10^{-3}

Thus, the conversion from one unit to the other unit is done by multiplying or dividing the powers of 10.

The most common units used to measure the length are as follows:

Kilometer (km)	Hectometer (hm)	Decameter (dam)	Meter (m)	Decimeter (dm)	Centimeter (cm)	Millimeter (mm)
1000	100	10	1	1/10	1/100	1/1000

3.5.2.3. **Volume or Capacity:**
The most common units used to measure the capacity or volume of any object are as follows:

Kiloliter (kl)	Hectoliter (hl)	Decaliter(dal)	Liter (l)	Deciliter (dl)	Centiliter (cl)	Milliliter(ml)
1000	100	10	1	1/10	1/100	1/1000

3.5.2.4. **Weight:**
The most common units used to measure the weight of any object are as follows:

Kilogram (kg)	Hectogram (hg)	Decagram (dag)	Gram (g)	Decigram (dg)	Centigram (cg)	Milligram (mg)
1000	100	10	1	1/10	1/100	1/1000

3.5.2.5. **Time:**

To calculate the time, the base unit is second. Now, let us discuss some other units of time.

1 minute = 60 seconds

1 hour = 60 minutes

1 day = 24 hours

1 week = 7 days

1 month = 30 or 31 days (Note: February has 28 days, but in leap year February has 29 days)

1 year = 12 months (or)

1 year = 365 days (But, in a leap year, 1 year = 366 days)

Whether you want to measure very small or very big things, you can go for metric number prefixes—for example- kilo and milli.

Name	The Number	Prefix	Symbol
trillion	1,000,000,000,000	tera	T
billion	1,000,000,000	giga	G
million	1,000,000	mega	M
thousand	1,000	kilo	k
hundred	100	hecto	h
ten	10	deka	da
unit	1		
tenth	0.1	deci	d
hundredth	0.01	centi	c
thousandth	0.001	milli	m
millionth	0.000 001	micro	μ
billionth	0.000 000 001	nano	n
trillionth	0.000 000 000 001	pico	p

Some common units are based on the kilogram, meter, and second.

Area = Square Meter (Area= length by length) Therefore the basic unit of area is equal to meters by meters = m^2 (square meters).

Volume = Cubic Meter. Therefore, the basic unit of volume = m^3 (cubic meters).

Liter = one-thousandth of a cubic meter (1 m^3 = 1,000 Liters) Therefore, 1 Liter = 1/1000 m^3

Time = Hour (1 hour=60 minutes, 1 minutes=60 seconds) Therefore,1 hour= 60×60= 3600 seconds

Day = (1 day= 24 hours) Therefore, 1 day= 24 x 60 x 60 = 86400 seconds

Important Things to Remember

Meter – Length	Kilo – Thousand
Liter – Volume	Milli – Thousand
Gram – Mass/Weight	Centi – Hundred
Celsius – Temperature	Deci – Ten

2.4.3 Metric System Conversion:

Metric conversion chart shows the conversion of Metric Units to Metric Units for Length.

Initial Value	Calculation	Results
millimeters (mm)	Divide by 10 (mm/10)	centimeter (cm)
centimeters (cm)	Multiply by 10 (cm * 10)	Millimeters (mm)
Meters (m)	Multiply by 100 (m * 100)	Centimeters (cm)
Centimeters (cm)	Divide by 100 (cm/100)	Meters (m)
Millimeters (mm)	Divide by 1000 (mm/1000)	Meters (m)

Conversion Ratio:

In the measurement of units, the **metric conversion ratio** is a ratio, which is equal to 1. It is also known as unit factors. It is used to convert the units between the same system, as well as the conversion between the metric and the English systems. It is based on the concept of equivalent values. Similarly, a conversion factor is defined as a number that is used to convert one unit to another unit by dividing or multiplying to get the equivalent values.

For example, to convert 5 meters to centimeters, a conversion factor is required.

We know that 1 meter = 100 centimeters

Thus, the conversion factor is "multiplying by 100". i.e.

5 m = 5 x 100 cm

5 m = 500 cm

Therefore, the conversion from 5 m to cm is 500 cm.

In terms of conversion ratio, it can be written as:

5m / 500 cm = 1

Metric System Examples

The following examples explain the conversion from one unit to the other unit.

Example 2.1:

Convert 356 centimeters to meters.

Solution:

We know that,

1 cm = 0.01 m

Thus,

356 cm = 356 x 0.01

356 cm = 3.56 m.

Therefore,

356 cm is equivalent to 3.56 meters.

Example 2.2:

Convert 4000 grams to kilograms.

Solution:

We know that,

1 gram = 0.001 kilogram

Thus,

4000 g = 4000 x 0.001

4000 g = 4 kilogram

Therefore,

4000 grams is equivalent to 4 kilogram

Example 2.3:

Convert 50 kiloliters to liters.

Solution:

We know that.

1 kiloliter = 1000 liters

Thus,

50 kiloliters = 50 x 1000 liters

50 kiloliters = 50000 liters

Therefore,50 kiloliters are equivalent to 50000 liters.

EXERCISE

Q1. What is meant by the metric system?

    A. The metric system is defined as the system of measurements to calculate the mass, distance, and volume of any object. We generally use the metric system to measure smaller or larger quantities.

Q2. What are the base units for length, weight, and volume in a metric system?

    A. In the metric system, the base units for length, weight, and volume are meters, grams, and liters respectively.

Q3. What is meant by metric number prefixes?

    A. The metric number prefix is defined as a unit prefix, which precedes a base unit to indicate the multiple or the submultiple of a unit. Some of the metric prefixes are:
Kilo (1000 or 10^3)
Hecto (100 or 10^2)
Deca (10 or 10^1)
Base unit (1)
Deci (1/10 or 10^{-1})
Centi (1/100 or 10^{-2})
Milli (1/1000 or 10^{-3})

Q4. Convert 80 grams to milligrams.

    A. We know that 1 gram = 1000 milligrams
therefore, 80 grams = 80 x 1000 milligrams
80 grams = 80000 milligrams
thus, 800 grams is equivalent to 80000 milligrams.

Chapter # 3

Gravity

Students will be able to understand:

- Speed
- Velocity
- Acceleration
- Center of gravity,
- Weight and mass.
- Work, Power, Energy.

We live in a universe of continual motion. In every piece of matter, the atoms are in a state of never-ending motion. We move around the Earth's surface, while the Earth moves in its orbit around the Sun. The Sun and the stars, too, are in motion; everything in the vastness of space is in a state of perpetual motion.

Every physical process involves motion of some sort. Because of its importance in the physical world around us.

3.1 Displacement:

-Whenever a body moves from one position to another, the change in its position is called displacement. The displacement can be represented as a vector that describes how far and in what direction the body has been displaced from its original position. The tail of the displacement vector is located at the position where the displacement started, and its tip or arrowhead is located at the final position where the displacement ended. For example, if a body is moving along a curve as shown in Fig. 3.1 with A as its initial position and B as its final position then the displacement d of the body is represented by AB. Note that although the body is moving along a curve, the displacement is different from the path of motion. ' f If r1 is the position vector of and r2 that of point B then by head and tail rule it can be seen from the figure that.

$$d = r_2 - r_1$$

Fig.3.1   Fig 3.2 (b)
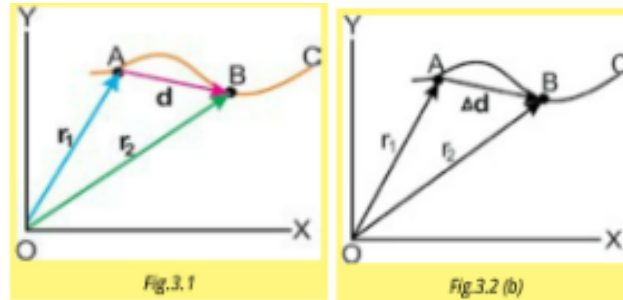
> **The displacement is thus a change in the position of body from its initial position to its final position.**

Its magnitude is the straight-line distance between the initial position and the final position of the body. When a body moves along a straight line, _ the. displacement coincides with the path of motion as shown in Fig. 3.2. _(a)

3.2 Velocity:

Time rate of change of displacement is known as velocity. Its direction.is along the direction of displacement. So if d is the total displacement of the body in time t, then its average velocity during the interval tis defined as

$$V_{av} = \frac{d}{t} \qquad \text{...........} \qquad (3.1)$$

Average velocity does not tell us about the motion between A and B. The path may be straight or curved and the / motion may be steady or variable. For example, if a squash. ball comes back to its starting point after bouncing off the wall several times, its total displacement is zero and so also is its" average velocity.

In such cases, the motion is described by instantaneous velocity.

To understand the concept of instantaneous velocity, consider a body moving along a path ABC in xy-plane. At any time t, let the body be at point A Fig_.3.2(b).Its position is given by position vector r_1". After a. short time the instant t, the body reaches point B which is described by the position vector T_2. The displacement of the body during this short time interval is given by:

$$\Delta d = r_2 - r_1$$

The notation Δ(delta) is used to represent a very small change.

The instantaneous velocity at point A, can be found making vt smaller and smaller. in this case, Δd will also become smaller and point B will approach A. If we continue this process, letting B Approach A, thus, allowing delta t and Δd F to decrease but never disappear completely, the ratio Δd/Δt approaches a definite limiting value which is the' instantaneous 1 velocity. Although At and Ad become extremely small in this process, yet their ratio is necessarily a small quantity. Moreover, while decreasing B the displacement vector, Δd approaches a limiting direction s r along the tangent at A. Therefore,

Fig. 3.3
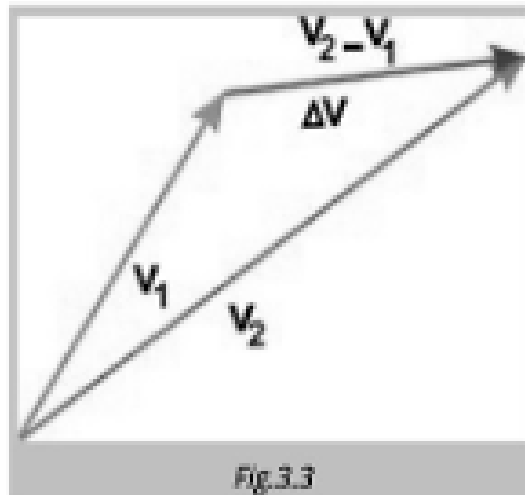
The instantaneous velocity is defined as the limiting value of △d/△t as the time interval △t following the time t, approaches zero.

Using the mathematical language, the definition of instantaneous velocity v_{ms} is expressed as

$$V_{ins} = \lim_{\Delta t \to 0} \frac{\Delta d}{\Delta t} \qquad \dotsb \qquad (3.2)$$

3.3 Acceleration:

If the velocity of an object changes, it is said to be moving with acceleration.

The time rate of change of velocity of a body is called acceleration.

As velocity is a vector any change in velocity may be due to a change in its magnitude or change in its direction or both.

Consider a body whose velocity v_1 at any instant t changes to v_2 in further small-time interval Δt. The two velocity vectors v_1 and v_2 and the change in velocity, delta v =v_2 — v_1 are represented in Fig: 3.3. The average acceleration a_{av} during time interval Δt is given by:

$$a_{av} = \frac{v_2 - v_1}{\Delta t} = \frac{\Delta v}{\Delta t} \quad \cdots\cdots\cdots\cdots \quad (3.3)$$

As a_{av} is the difference of two vectors divided by a scalar At, a_{av} must also be a vector. Its direction is the same as that of ΔV. Acceleration of a body at a particular instant is known as instantaneous acceleration and it is the value obtained from the average acceleration as Δt is made smaller and smaller till it approaches zero. Mathematically, it is expressed as

$$\text{Instantaneous acceleration} = a_{ns} = \lim_{\Delta t \to 0} \frac{\Delta v}{\Delta t} \quad \cdots\cdots\cdots\cdots \quad (3.4)$$

If the velocity of a body is increasing, its acceleration is positive but if the velocity is decreasing the (acceleration is negative. if the velocity of the body changes by an equal amount in equal intervals of time, the body is said to have uniform acceleration. For a body moving with uniform acceleration, it average acceleration is equal to that of instantaneous acceleration.

3.4 Velocity Time Graph:

Graphs may be used to illustrate the variation of velocity of an object with time. Such graphs are called velocity-time graphs. The Velocity-Time graphs of an object making three different journeys along a straight road are shown in figures 3,4 to 3.6. When the velocity of the car is constant, its velocity-time graph is a horizontal straight line (Fig

3.4). When the car moves with constant acceleration, the velocity-time graph is a straight line that rises the same height for equal intervals of time (Fig 3.5).

Fig.3.4
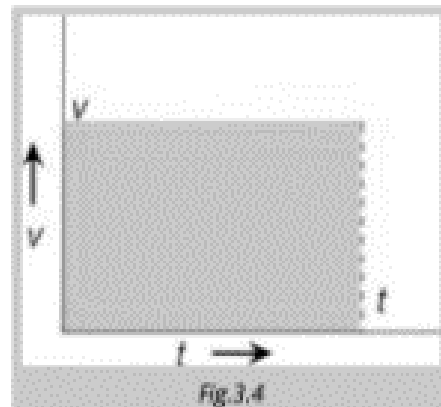
The average acceleration of the car during the interval t is given by the slope of its velocity-time graph.

When the car moves with increasing acceleration, the velocity-time graph is a curve (Fig 3.6). The point A on the graph corresponds to time t._ The magnitude of the instantaneous acceleration at this instant is numerically equal to the slope of the tangent at the point _A on the velocity-time graph of the object as shown in Fig 3.6.

Fig.3.5
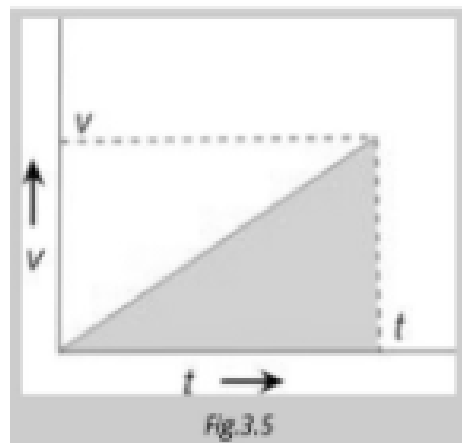
The distance moved by an object can also be determined by using its velocity-time graph. For example, Fig 3.4 shows that the object moves at constant velocity v for time it. The distance covered: by the object given by Eq. 3.1 is v x t. This distance. can also be found by calculating the area under the velocity-time graph. This area is shown shaded in Fig 3.4 and is equal to v x t. We now give another example shown in Fig 3.5.

Here the velocity of the object increases uniformly from 0 to v in time t. The magnitude of its average velocity is given by:

Fig 3.6
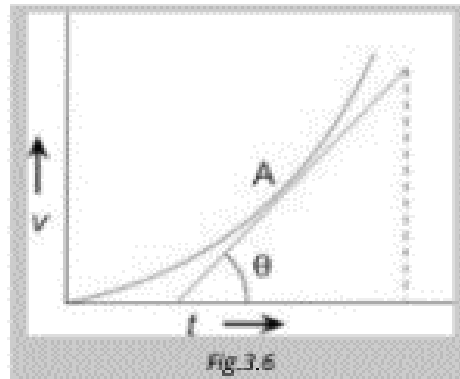
$$V_{av} = \frac{0 + v}{2} = \frac{1}{2}v$$

Distance covered = average velocity x time = $\frac{1}{2}v \times t$

Now we calculate the area under velocity-time graph which is equal to the area of the triangle shaded in Fig 3.5. Its value is equal to 1/2 base x height = 1/2 v x t. Considering the above two examples it is a general conclusion that

The area between the velocity-time graph and the time axis is numerically equal to the distance covered by the object.

3.5 Centre Of Gravity

One way to look at gravity is to look at it as a force that pulls things downward more precisely towards the centre of the Earth. But it doesn't always work like that.Sometimes gravity causes things to topple and turn over, especially if they are high up and unbalanced. No one understands this better than tightrope walkers. While tiptoeing on the high wire, they often wobble from side to side to entertain us, yet they hardly ever fall.

3.5.1 What is the Centre of Gravity?

The Centre of gravity is a theoretical point in the body where the body's total weight is thought to be concentrated. It is important to know the centre of gravity because it predicts the behavior of a moving body when acted on by gravity. It is also useful in designing static structures such as buildings and bridges.

In a uniform gravitational field, the centre of gravity is identical to the centre of mass. Yet, the two points do not always coincide. For the Moon, the centre of mass is very close to its geometric centre. However, its centre of gravity is slightly towards the Earth due to the stronger gravitational force on the Moon's near side.

In a symmetrically shaped object formed of homogenous material, the centre of gravity may match the body's geometric centre. However, an asymmetrical object composed of various materials with different masses is likely to have its centre of gravity located at some distance away from its geometric centre. In hollow bodies or irregularly shaped bodies, the centre of gravity lies at a point external to the physical material.

3.5.2 Calculating the Centre of Gravity

In the next few sections, we will learn how to determine the centre of gravity of objects. In general, determining the centre of gravity may be complicated because the mass and weight may not be uniformly distributed throughout the object.

3.5.2.1 Determining the Centre of Gravity for Symmetrical Bodies
For simple shaped object, the centre of gravity can be determined by the following ways:

When we balance an object with the help of a string or an edge, the point at which the object balances is the centre of gravity.

Another simple physical procedure employed to determine the centre of gravity is as follows:

Suspending the plate given in the figure by attaching a cord to point A and then by attaching the cord at point C, the centre of gravity of the plate can be located.
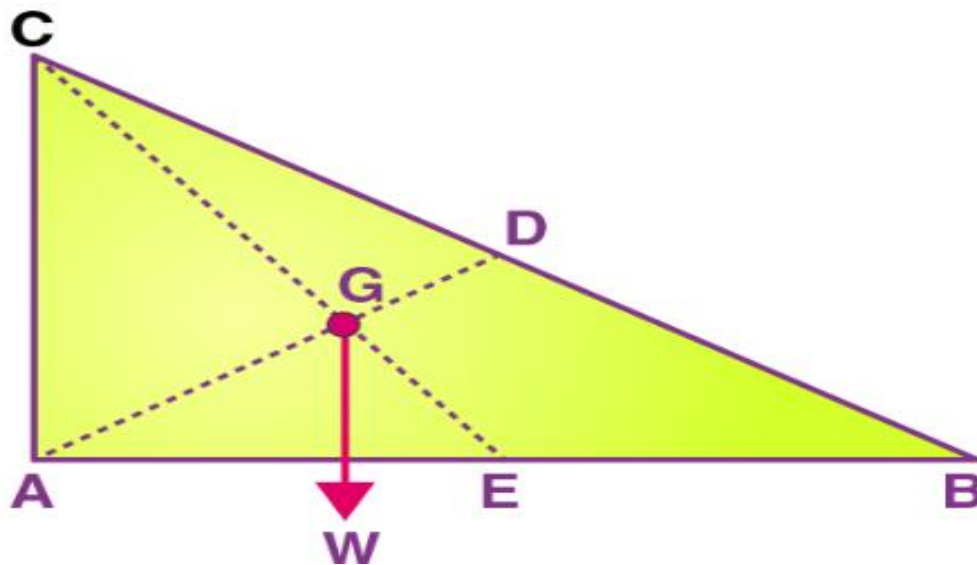
The line AD is vertical when the plate is suspended from A; likewise, the line CE is vertical when suspended from C. The centre of gravity of the plate lies at the intersection of AD and CE. When an object is suspended from any single point, its centre of gravity lies directly beneath that point.

3.5.2.6. Determining the Centre of Gravity for Asymmetrical Bodies

We make use of calculus to determine the centre of gravity when the mass is not distributed uniformly. We denote the integration of a continuous function with respect to weight using the symbol **S dw.** Then, the centre of gravity is determined as follows:

$cg \times W = SxdW$

Here, **x** is the distance from a reference line, **dw** is an increment of weight, and **W** is the total weight of the object.

We know that

$W = mg$

But, the mass of an object is equal to

$m = \rho V$

Combining both the equations, we get

$W = g\rho V$

Solving,

$dW = g\rho dV$
$dW = g * \rho(x, y, z) * dxdydz$

If we have a functional form of the mass distribution, we can solve the equation for the centre of gravity.

$cg * W = g * SSSx * \rho(x, y, z)dxdydz$

SSS in the equation indicates a triple integral over **dx. dy.** and **dz.**

3.5.2.7. Difference between Centre of Mass and Centre of Gravity

It is common to assume the centre of mass and the centre of gravity to be the same, but this is not the case.

The centre of mass is a point at which mass distribution is equal in all directions, and it doesn't depend upon the gravitational field. The centre of gravity is a point in an object where the distribution of weight is equal in all directions, and it does depend on the gravitational field.

However, an object's centre of mass and centre of gravity lies at the same point in a uniform gravitational field.

In the above figure, the object on the left is in a uniform gravitational field and has an overlapping centre of gravity and mass. For the object on the right, the gravitational field is stronger towards its base, and the centre of gravity is below the centre of mass.

How can a ring be balanced if its centre of gravity lies outside?

The centre of gravity of a uniform circular ring does not lie in its centre but at its geometric centre. Although it lies outside the mass, the centre of gravity of a circular ring balances itself when placed on another object, as shown in the figure. The gravitational force acts through the supporting object.
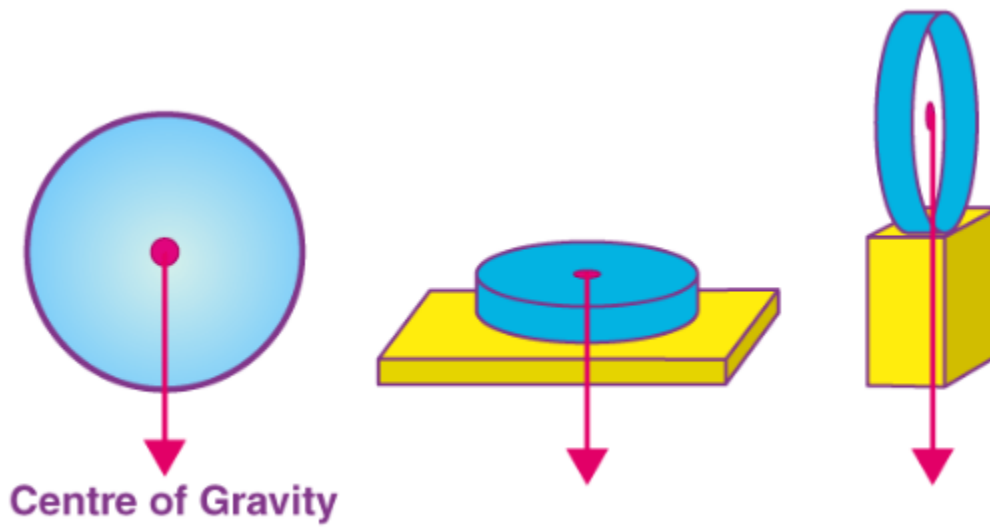
Centre of Gravity

Exercise

i.   Who discovered the centre of gravity?
ii.   How does the Centre of gravity affect the balance?
iii.   Why is the centre of gravity important?
iv.   Where is the centre of gravity in a human body?
v.   How does centre of gravity affect stability?
vi.   What is the difference between uniform and variable velocity? From explanation or variable velocity, define acceleration. Give SI units of velocity and acceleration.
vii.   An object is thrown vertically upward. Discuss the sign of acceleration due to gravity, relative to velocity, while the object is in air.
viii. Can the velocity of an object reverse the direction when acceleration is constant? If so, give an example.
ix.   Specify the correct statements:

a. An object can have a constant velocity even its speed is changing.

b. An object can have a constant speed even its velocity is changing.

c. An object can have a zero velocity even its acceleration is not zero.

d. An object subjected to a constant acceleration can reverse its velocity.

x.   A man standing on the top of a tower throws a ball straight up with initial velocity v,and at the same time throws a second ball straight downward with the same speed. Which ball will have a larger speed when it strikes the ground? Ignore air friction.

NUMERICALS:

1. A helicopter is ascending vertically at the rate of 19.6 m/s. When it is at a height of 156.8 m above the ground, a stone is dropped. How long does the stone take to reach the ground?

(Ans: 8.0s)

2. Using the following data, draw a velocity-time graph for a short journey on a straight road of a motorbike.

Velocity (ms⁻¹)	0	10	20	20	20	20	0
Time (s)	0	30	60	90	120	150	180

Use the graph to calculate
   a.   the initial acceleration
   b.   the final acceleration
   c.   the total distance travelled by the motorcyclist

[Ans:(8) 0.33 ms^{-2} (b) -0 .67ms^{-2}. (c) 2.7k m]

3. A proton moving with speed of 1.0×10^7 m/s passes through a 0.020 cm thick sheet of paper and emerges with a speed. of 2.0×10^6 m/s. Assuming uniform deceleration, find retardation and time taken to pass through the paper.

(Ans: - 2.4×10^{17} ms^{-2}, 3.3×10^{-11} S)

4. Two masses m1 and m2 are initially at rest with a spring compressed between them. What is the ratio of the magnitude of their velocities after the spring has been released?

$$\left(\text{Ans} : \frac{v_1}{v_2} = \frac{m_2}{m_1}\right)$$

Chapter # 4

Mechanics

Students will be able to understand:

- Force
- Equation of motion
- Laws of motion

In school physics, we have studied some useful equations for objects moving at constant acceleration.

4.1 Equation of Motion:

Suppose an object is moving with uniform acceleration along a straight line. if its initial velocity is v_i, and the final velocity after a time interval t is v_f. Let the distance covered during this interval be S then we have:

$$v_f = v_i + at \qquad \text{.... 4.1}$$

$$S = \frac{(v_f + v_i)}{2} \times t \qquad \text{.... 4.2}$$

$$S = v_i t + \frac{1}{2} at^2 \qquad \text{.... 4.3}$$

$$v_f^2 = v_i^2 + 2aS \qquad \text{.... 4.4}$$

These equations are useful only for linear motion with uniform acceleration. When the object moves along a straight line, the direction of motion does not change. In this case all the vectors can be manipulated like scalars. In such problems, the direction of initial velocity is taken as positive. A negative sign is assigned to quantities where the direction is opposite to that of the initial velocity.

In the absence of air resistance, all objects in free fall near the surface of the Earth move towards the Earth with a uniform acceleration. This acceleration, known as acceleration due to gravity, is denoted by the letter g and its average value near the Earth's surface is taken as 9.8 ms^{-2} in the downward direction. The equations for uniformly accelerated motion can also be applied to the free fall motion of the objects by replacing a by g.

4.2 NEWTON'S LAW OF MOTION

Newton's laws are empirical laws, deduced from experiments. They were clearly stated for the first time by Sir Isaac Newton, who published them in 1687 in his famous 'book called "Principia". Newton's laws are adequate for speeds that are low compared with the speed of light.

Distance covered = average velocity × time = $\frac{1}{2}v \times t$ …. 4.5

Now we calculate the area under the velocity-time graph which is equal to the area of the triangle shaded in Fig 3.5. Its value is equal to 1/2 base x height= 1/2 v x t. T Considering the above two examples it is a general conclusion that

The area between the velocity -time graph and the time axis is numerically equal to the distance covered by the object.

4.2.1 Newton's First Law of Motion

A body at rest will remain at rest, and a body moving with uniform velocity will continue to do so, unless acted upon by some unbalanced external force. This is also known as the law of inertia. The property of an object tending to maintain the state of rest or state of uniform motion is referred to as object's inertia. The more inertia, the stronger is this tendency in the presence of a force. Thus,

The mass of the object is a quantitative measure of its Inertia.

The frame of reference in which Newton's first law of motion' holds, is known as inertial frame of reference. A frame of reference stationed on Earth is approximately an inertial frame of reference.

4.2.2 Newton's Second Law of Motion

A force applied on a body produces acceleration in its own direction. The acceleration produced varies directly with the applied force and inversely with the mass of the body. Mathematically, it is expressed as

$$F = ma$$

4.2.3 Newton's Third Law of Motion

Newton's Third Law of Motion states that "for every action, there is an equal and opposite reaction." This law is one of the fundamental principles of classical physics, and it describes the relationship between the forces two objects exert on each other when they interact. For example, whenever an interaction occurs between two objects, each object exerts the same force on the other, but in the opposite direction and for the same

length of time. Each force in the action-reaction pair acts only on one of the two bodies, the action. and reaction forces never act on the same body.

4.2.4 Applications of Newton laws:

Newton's laws of motion have practical applications in transportation, engineering, sports, aerospace, robotics, and more. They help design vehicles, structures, sports techniques, and machinery. These laws are crucial in understanding fluid dynamics, civil engineering, physics, and astronomy. They're used in medical devices, entertainment, and environmental analysis. Newton's laws are fundamental in physics education and scientific research, guiding our understanding of motion and forces in the world.

4.3 MOMENTUM

All moving objects possess a quality by which they exert force on anything that tries to stop it. The faster the object is traveling, the harder is to stop it. Similarly, if two objects move with the same velocity, then it is more difficult to stop the massive of the two.

This quality of the moving body was called the quantity of motion of the body, by Newton. This term is now called linear momentum of the body and is defined by the relation.

Linear momentum= P=mv   …. 4.6

In this expression v is the velocity of the mass m. Linear momentum is, therefore, a vector quantity and has the direction of velocity.

The SI unit of momentum is kilogram meter - per second (kg m s^{-1}). It can also be expressed as newton second" (N s).

4.3.1 Momentum and Newton's Second Law of Motion

Consider a body of mass m moving with an initial velocity v_i, suppose an external for F acts upon it for time t after which velocity becomes v_f. The acceleration produced by this force is given by.

$$a = \frac{v_f - v_i}{t} \quad \text{…. 4.7}$$

By Newton's second law, the acceleration is given as:

$$a = \frac{F}{m} \quad \text{…. 4.8}$$

Equating the two expressions of acceleration, we have

$$\frac{F}{m} = \frac{v_f - v_i}{t}$$

$$F \times t = mv_f - mv_i \qquad \text{.... 4.9}$$

where mv_i is the initial momentum and mv_f is the final momentum of the body.

The Above equation shows that change in momentum is equal to the product of force and the time for which force is applied. This form of the second law is more general than the form F=ma because this can easily extend to account for changes as the body accelerates when its mass also changes. For example, as a rocket accelerates, it loses mass because its fuel is burnt and ejected to provide greater thrust.

From the above equation,

$$F = \frac{mv_f - mv_i}{t} \qquad \text{.... 4.10}$$

Thus, second law of motion can also be stated in terms of momentum as follows;

"Time rate of change of momentum of a body equals the applied force".

4.3.2 Impulse:

Sometimes we wish to apply the concept of momentum to cases where the applied force is not constant, it acts for a very short time. For example, when a bat hits a cricket ball, the force certainly varies from instant to instant during the collision. In such cases, it is more convenient to deal with the product of force and time (F x t) instead of either quantity alone. The quantity (F x t) is called the impulse of the force, where F can be regarded as the average force that acts during the time t. From Eq. 4.11

$$F \times t = mv_f - mv_i \qquad \text{…. (4.11)}$$

We get;

$$\text{Impulse} = F \times t = mv_f - mv_i \qquad \text{….. (4.12)}$$

4.3.3 Law of Conservation of Momentum

Let us consider an isolated system. It is a system on which no external agency exerts any force. For example, the molecules of a gas enclosed in a glass vessel at constant temperature constitute an isolated system. The molecules can collide with one another because of their random motion but being enclosed by a glass vessel, no external

agency can exert a force on them. Consider an isolated system of two smooth hard interacting balls of masses m1 and mg, moving along the same straight line, in the same "direction, with velocities v1 and v2 respectively. Both the balls collide and after the collision, ball of mass m1 moves with velocity v'1 and mg moves with velocity v'2 in the same direction as shown in Fig 4.1.
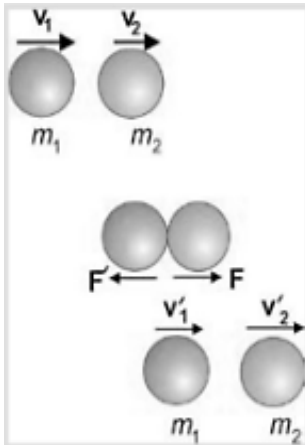
Figure. 4.1

To find the change in momentum of mass m_1, using Eq 3.11 we have,

$$F' \times t = m_1 v'_1 - m_1 v_1$$

Similarly for the ball of mass m_2, we have

$$F \times t = m_2 v'_2 - m_2 v_2$$

Adding these two expressions, we get

$$(F + F')t = (m_1 v'_1 - m_1 v_1) + (m_2 v'_2 - m_2 v_2)$$

Since the action force F is equal and opposite to the reaction force F', we have F'= - F, so the left-hand side of the equation is zero.

Hence,

$$0 = (m_1 v'_1 - m_1 v_1) + (m_2 v'_2 - m_2 v_2)$$

In other words, change in momentum of 1st ball + change in momentum of the 2 ball = 0

$$\text{Or} \quad (m_1 v_1 + m_2 v_2) = (m_1 v'_1 + m_2 v'_2) \quad \text{.... 4.13}$$

Which means that total initial momentum of the system before collision is equal to the total final momentum of the system after collision. Consequently, the total change in momentum of the isolated two ball system is zero. For such a group of objects, if one object within the group experiences a force, there must exist an equal but opposite reaction force on some other object in the same group. As a result, the change in momentum of the group of objects is always zero. This can be expressed in the form of law of conservation of momentum which states that

The total linear momentum of an isolated system remains constant.

The momentum of a body is a vector quantity.

4.4 Inelastic and elastic collision:

When two tennis balls collide then, after collision, they will rebound with velocities less than the velocities before the impact. During this process, a portion of K.E is lost, partly due to friction as the molecules in the ball move past one another when the balls distort and partly due to its change into heat and sound energies.

"A collision in which the K.E of the system is not conserved, is called inelastic collisions"

Under certain special conditions no kinetic energy is lost in the collision.

For example, when a hard ball is dropped onto a marble floor, it bounces to very nearly the initial height, and loses the negligible amount of energy in the collision with the floor. It is to be noted that momentum and total energy are consent/ed in all types of collisions. However, the K.E. is conserved only in elastic collisions.

4.4.1 **Elastic Collision in One Dimension**

Consider two smooth, non-rotating balls of masses m, and mg, moving initially with velocities v1 and v2 respectively, in the same direction. They collide and after collision, they move along the same straight line without rotation. Let their velocities after the collision be v1' and v5 respectively, as shown in Fig. 3.2. We take the positive direction of the velocity and momentum to the right. By applying the law of conservation of momentum we have
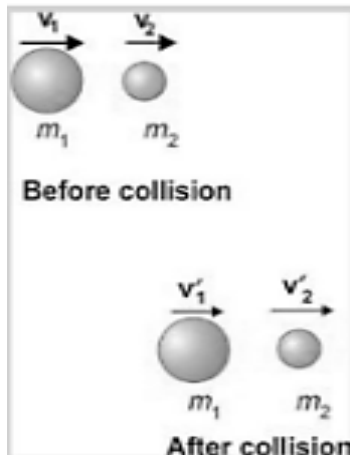
$$m_1 v_1 + m_2 v_2 = m_1 v'_1 + m_2 v'_2$$

$$m_1(v_1 - v'_1) = m_2(v'_2 - v_2) \qquad .... 4.14$$

Before collision

After collision

Fig.4.2

As the collision is elastic,so the K.E is also conserved. From the conservation of K.E we have

$$\frac{1}{2}m_1 v_1^2 + \frac{1}{2}m_2 v_2^2 = \frac{1}{2}m_1 v_1'^2 + \frac{1}{2}m_2 v_2'^2$$

or $$m_1(v_1^2 - v_1'^2) = m_2(v_2'^2 - v_2^2)$$

or $$m_1(v_1 + v'_1)(v_1 - v'_1) = m_2(v'_2 + v_2)(v'_2 - v_2) \qquad .... 4.15$$

$$(v_1 + v'_1) = (v'_2 + v_2)$$

or $$(v_1 - v_2) = (v'_2 - v'_1) = -(v'_1 - v'_2) \qquad .... 4.16$$

We noted that, before. collision (v_1 - v_2) is the velocity of the first ball relative to the second ball.

Similarly, ($v_1 - v_2$) is the velocity of the first ball relative to the second ball after collision. It means that relative velocities before and after the collision have the same magnitude but are reversed after the collision. In other words, the magnitude of relative velocity of approach is equal to the magnitude of the relative velocity of the separation. - In equations 4.14 and 4.16, m_1, m_2, v_1 and v_2 are known quantities. We solve these equations to find the values of v'_1 and v'_2, which are unknown.

The results are

$$v'_1 = \frac{m_1 - m_2}{m_1 + m_2} v_1 + \frac{2m_2}{m_1 + m_2} v_2 \qquad \text{---------} \qquad 4.17$$

$$v'_2 = \frac{2m_1}{m_1 + m_2} v_1 + \frac{m_2 - m_1}{m_1 + m_2} v_2 \qquad \text{---------} \qquad 4.18$$

There are some cases discussed below:

When $m_1 = m_2$

From equations 4.17 and 4.18 we find that of special interest, which are

$V'_1 = V_2$

$V'_2 = V_1$

as shown in Fig 4.3

When $m_1 = m$, and $v_2 = 0$

In this case the mass m_2 be at rest,

then $v_2 = 0$ the equations

4.17 and 4.18 give.
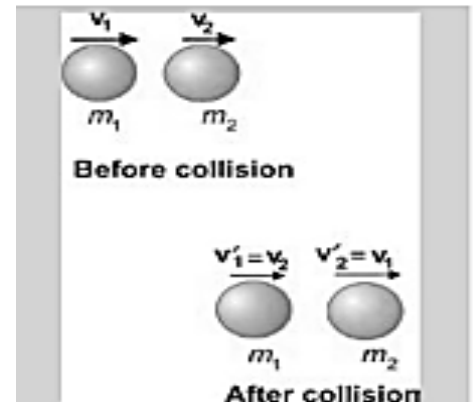
Fig 4.3
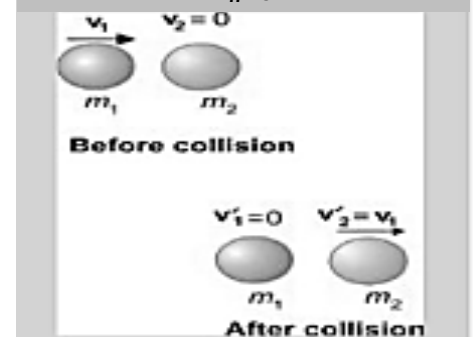
Fig 4.4
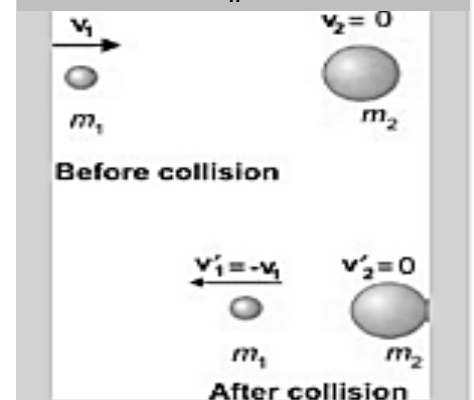
Fig 4.5

$$v'_1 = \frac{m_1 - m_2}{m_1 + m_2} v_1 \quad ; \quad v'_2 = \frac{2m_1}{m_1 + m_2} v_1$$

When a light body collides with a massive body at rest

In this case initial velocity $V_2 = 0$, and $m_2 \gg m_1$. Under these conditions, m_1 can be neglected as compared to m_2. From equations 4.17 and 4.18, we have $v'_1 = -vi$ and $v'_2 = 0$.

The result is shown in Fig 4.4. This means that mi will bounce back with the same velocity while mg will remain stationary. This fact is made use of by the squash player.

When a massive body collides with a light stationary body

In this case $m_1 \gg m_2$ and $v_2 = 0$ so m_2 can be neglected in equations 4.17 and 4.18. This gives $v'_2 = v_1$ and $v'_2 = 2vi$. Thus, after the collision, there is practically no change in the velocity of the massive body, 'but the lighter one bounces off in the forward direction with approximately twice the velocity of the incident body, as shown in Fig. 4.5.

Forces due to water flow:

When water from a horizontal pipe strikes a wall normally, a. force is exerted on the wall. (Suppose the water strikes the wall normally with velocity v and comes to rest on striking the wall, the change in velocity is then 0. V = — V. From the second law, the force equals the momentum change per second of water. if mass m of the water strikes the wall in time t then force F on the water is

$$F = -\frac{m}{t}v = - \text{mass per second} \times \text{change in velocity} \quad \text{......} \quad \textbf{4.19}$$

From third law of motion, the reaction force exerted by water on the wall is equal but opposite.

Hence, $\qquad F = -\left(-\frac{m}{t}v\right) = \frac{m}{t}v$

Thus, force can be calculated from the product of mass of water striking normally per second and change in velocity. Suppose the water flows out from a pipe at 3 kgs^{-1} and its velocity changes from 5 ms^{-1} to zero on striking the ball, then,

Force = $3 \text{ kgs}^{-1} \times (5 \text{ ms}^{-1} - 0) = 15 \text{ kgms}^{-2} = 15\text{N}$

EXERCISE

1. Explain the circumstances in which the velocity v and acceleration a of a car are.
    i. Parallel
    ii. Anti-parallel
    iii. Perpendicular to one another
    iv. v is zero but a is not zero
    v. a is zero but v is not zero

2. Motion with constant velocity is a special case of motion with constant acceleration. Is this statement true? Discuss.

3. Find the change in momentum for an object subjected to a given force for a given time and state law of motion in terms of momentum.

4. Define impulse and show how it is related to linear momentum?

5. State the law of conservation of linear momentum, pointing out the importance of isolated system. Explain, why under certain conditions, the law is useful even though the system is not completely isolated?

6. Explain the difference between elastic and inelastic collisions. Explain how a bouncing ball would behave in each case.

Numericals

1. An amoeba of mass 1.0×10^{-10} kg propels itself through water by blowing a jet of water through a tiny orifice. The amoeba ejects water with a speed. of 1.0×10^{-4} m/s and at a rate of 1.0×10^{-13} kg/s. Assume that the water is being continuously replenished so that the mass of the amoeba remains the same.
   a. If there were no force on amoeba other than the reaction force caused by 1 the emerging jet, what would be the acceleration of the amoeba?
   b. If amoeba moves with constant velocity through water, what is force of surrounding water (exclusively of jet) on the amoeba?

[Ans: (a) 1.0×10^{-5} ms^{-2} (b) 1.0×10^{-17} N]

2. A boy places a firecracker of negligible mass in an empty can of 40 g mass. He plugs the end with a wooden block of mass 200 g. After igniting the firecracker, he throws the can straight up. It explodes at the top of its path. If the block shoots out with a speed of 3.0 m/s, how fast will they be going?

(Ans: 15 m/s)

3. A truck weighing 2500 kg and moving with a velocity of 21 m/s collides with the stationary car weighing 1000 kg. The truck and the car move together after the impact. Calculate their common velocity.

(Ans: 15 m/s)

4.Two blocks of masses 2.0 kg and 0.50 kg are attached at the two ends of a ' compressed spring. The elastic potential energy stored in the spring is 10 J. Find the velocities of the blocks if the spring delivers its energy to the blocks when released.

(Ans: 1.4 m/s, -5.6 m/s)

Chapter # 5

Work, Power, Energy

Students will be able to understand:

- Work
- Power
- Energy
- Interconversion of different forms of energy i.e . K.E & P.E.

5.1 Work:

Work is often thought of in terms of physical or mental effort. Physics, however, the term work involves two things

(i)   force
(ii) displacement.
       We shall begin with a simple situation in which work is done by constant force.

5.1.1 WORK DONE BY A CONSTANT FORCE

Let us consider an object that is being pulled by a constant force F at an angle 9 to the direction of motion. The force displaces the object from position A to B through a displacement d (Fig. 5.1).
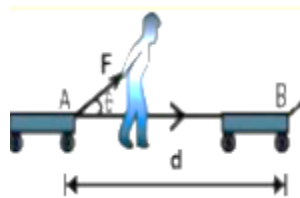
Fig. 5.1: work done by a constant force.

We define work W done by the force F as the scalar product of F and d.

W=F.d=Fd cos θ ….…..   5.1

The quantity (Fcosθ) is the component of the force in the direction of the displacement d.

> The work done on a body by a constant force is defined as the product of the magnitudes of the displacement and the component of the force in the direction of the displacement.

Can you tell how much work is being done?

On the pail when a person holding the pail by the force F is moving forward (Fig. 5.2 a).

On the wall (Fig. 5.2 b)?
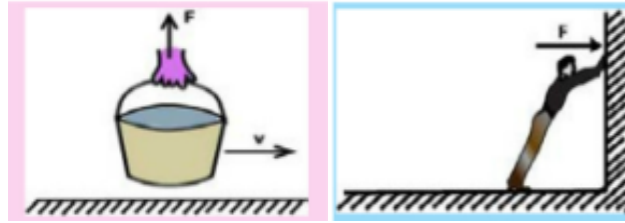
5.2 (a)   5.2 (b)

When a constant force acts through a distance d, the event can be plotted on a simple graph (Fig. 4.3). The distance is normally plotted along x-axis and the force along y-axis. In this case-as the force does not vary, the graph _will be a horizontal straight line. If the constant force F (newton) and the displacement d (meter) are in the same direction then the work done is Fd- (joule). The clearly shaded area in Fig. 5.3 is also Fd. Hence the area under a force- displacement curve can be taken to represent the work done by the force. In case the force F is not in the direction of displacement, the graph is plotted between F $\cos\theta$ and d.
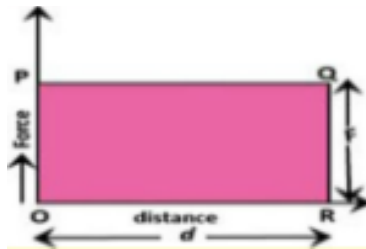
Fig. 5.3

From the definition of work, we find that:

Work is a scalar quantity.

If 9 < 90°, work is done and it is said to be positive work.

If e = 90°, no work is done.

If 0 > 90°, the work done is said to be negative.

SI unit of work is N m known as joule (J).

5.1.1 WORK DONE BY A VARIABLE FORCE

In many cases, the force does not remain constant during the process of doing work. For example, as a rocket moves away from the Earth, work is done against the force of

gravity, which varies as the inverse square of the distance from the Earth's centre. Similarly, the force exerted by a spring increases with the amount of stretching.

How can we calculate the work done in such a situation?

Fig. 5.4 A particle acted upon by a variable force, moves along the path shown from point a to point b.
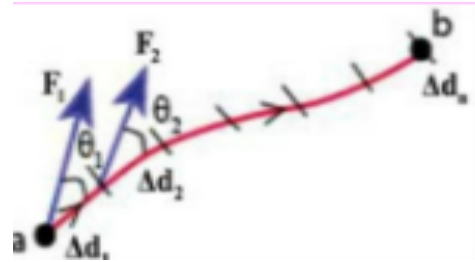
Fig. 5.4 shows the path of a particle in the x-y plane as it moves from point a to point b. The path has been divided into n short intervals of displacements Δd_1, Δd_2 , ... Δd_n , and F_1, F_2, ... Fn , are the forces acting during these intervals. During each small interval, the force is supposed to be approximately constant. So the work done for the first interval can then be written as

$$\Delta W_1 = F_1 . \Delta d_1 = F_1 \cos \theta_1 \, \Delta d_1$$

and in the second interval

$$\Delta W_2 = F_2 . \Delta d_2 = F_2 \cos \theta_2 \, \Delta d_2$$

and so on. The total work done in moving the object can be calculated by adding all these terms.

$$W_{total} = \Delta W_1 + \Delta W_2 + ... + \Delta W_n$$
$$= F_1 \cos \theta_1 \, \Delta d_1 + F_2 \cos \theta_2 \, \Delta d_2 + ... + F_n \cos \theta_n \, \Delta d_n$$
$$W_{total} = \sum_{i=1}^{n} F_i \cos \theta_i \Delta d_i \qquad .... \quad 5.2$$
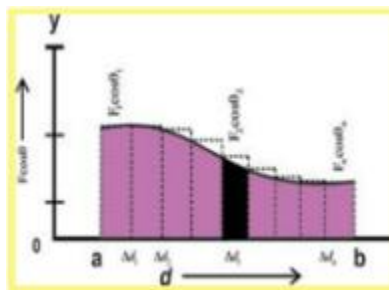
Fig .5.5

We can examine this graphically by plotting F cosθ verses d,'as shown in Fig. 5.5. The displacement d has been subdivided into n equal intervals. The value of Fcosa at the beginning of each interval is indicated in figure.

Now the ith shaded rectangle has an area F, - cosθ Δd, which is the work done during the ith interval. Thus, the work done given by Eq. 5.2 equals the sum of the areas of all the rectangles. If we subdivide the distance into many intervals so that each Ad becomes very small, the work done given by Eq. 5.2 becomes more accurate. If we let each Δd to approach zero then we obtain an exact result for the work done, such as

$$W_{total} = \underset{\Delta d \to 0}{limit} \sum_{i=1}^{n} F_i \cos \theta_i \Delta d_i \qquad \cdots\cdots\cdots 5.3$$

in this limit Δd approaches zero, the total area of the rectangles (Fig. 5.5) approaches the area between the F cosθ curve and d-axis from a to b as shown shaded in Fig. 5.6.

Thus, the work done by a variable force in moving a particle between two points is equal to the area under the F cosθ verses d curve between the two points a and b as shown in Fig. 5.6
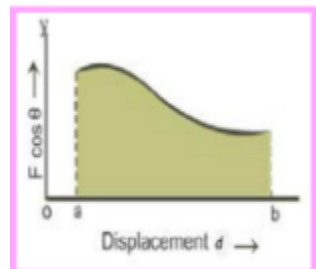
Fig .5.6

5.1.2 WORK DONE BY GRAVITATIONAL FIELD

The space around the Earth in which its gravitational force acts on a body is called the gravitational field. When an object is moved in the gravitational field, the work is done by the gravitational force. If displacement is in the direction of gravitational force, the work is positive. If the displacement is against the -gravitational force, the work is negative. Let us consider an object of mass m being displaced with constant velocity from point A to B -along various paths in the presence of a gravitational force (Fig.5.7). In this case the gravitational force is equal to the weight mg of the object.
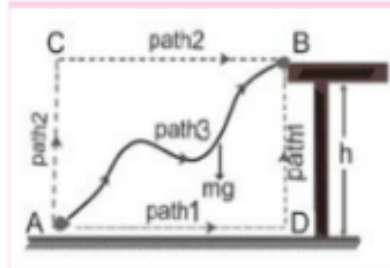
Fig .5.7

The work done by the gravitational force along the path ADB can be split into two parts. The work done along AD is zero, because the weight mg is perpendicular to this path, the work done along DB is (-mgh) because the direction of mg is opposite to that of displacement i.e; θ=180°Hence, the work done is displacing a body from A to B through path 1 is

$$W_{ADB} = 0 + (-mgh) = -mgh$$

If we consider the path ACB, the work done along AC is also (-mgh). Since the work done along CB is zero, therefore,

$$W_{ACB} = -mgh + 0 = -mgh$$
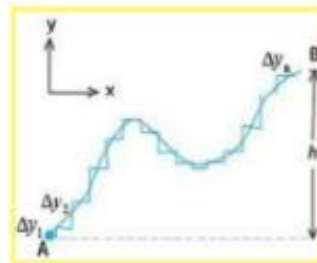
Fig.5.8:A smooth path may be a series of infinitesimal x and y displacements. Work is done only during the displacements.

Let us now consider path 3, i.e. a curve one. Imagine the v curved path, to be broken down into a series of horizontal and vertical steps as shown in Fig.5.8. There is no work done along the horizontal steps because mg is perpendicular to the displacement for these steps. Work is done by the force of gravity only along vertical displacements.

$$W_{AB} = -mg(\Delta y_1 + \Delta y_2 + \Delta y_3 + \ldots\ldots + \Delta y_n)$$

as

$$(\Delta y_1 + \Delta y_2 + \Delta y_3 + \ldots\ldots + \Delta y_n) = h$$

Hence,

$$W_{AB} = -mgh$$

The net amount of work done along AB path is still (-mgh). We conclude from the above discussion that:

Work done in the Earth's gravitational field is independent of the path followed.

Can you, prove that the work done along a closed path such as ACBA or ADBA (Fig. 5.7), in a gravitational field is zero?

The field in which the work done be independent of the path followed or work done in a closed path be zero, is called a conservative field.

The frictional force is a non-conservative force, because if an object is moved over a rough surface between two points along different paths, the work done against the frictional force certainly depends on the path followed.

5.2 POWER:

Power is the measure of the rate at which work is being done.

If work ΔW is done in a time interval Δt, then the average power P, during the interval At is defined as

$$P_{av} = \frac{\Delta W}{\Delta t} \qquad \text{..........} \quad 5.4$$

If work is expressed as a function of time, the instantaneous power P at any instant is defined as

$$P_{av} = \underset{\Delta t \to 0}{Limit} \frac{\Delta W}{\Delta t} \qquad \text{..........} \quad 5.5$$

Where ΔW is the work done in short interval of time Δt following the instant t.

5.2.1 Power and Velocity

It is, sometimes, convenient to express power in terms of a constant force F acting on an object moving at constant velocity v. For example, when the propeller of a motorboat A causes the water to exert a constant force F on the boat, it moves with a constant velocity v. The power delivered by the motor at any instant is, then, given by

$$P = \underset{\Delta t \to 0}{Limit} \frac{\Delta W}{\Delta t}$$

we know $\quad\quad \Delta W = \mathbf{F}.\,\Delta\mathbf{d}$

so $\quad\quad\quad P = \underset{\Delta t \to 0}{Limit} \frac{\mathbf{F}.\,\Delta\mathbf{d}}{\Delta t}$

Since $\quad\quad \underset{\Delta t \to 0}{Limit} \frac{\Delta\mathbf{d}}{\Delta t} = \mathbf{v}$

Hence, $\quad\quad P = \mathbf{F}.\,\mathbf{v}$ 5.6

The SI unit of power is watt, defined as one joule of work done in one second.

Sometimes, for example, in electrical measurements, the unit of work is expressed as watt second. However, a commercial unit of electrical energy is kilowatt-hour. One kilowatt hour is the work done in one hour by an agency whose power is one kilowatt.

Therefore, $\quad\quad$ 1kWh = 1000 W x 3600 s.

or $\quad\quad\quad\quad$ 1kWh = 3.6 x 10^6 J = 3.6 M J

5.3 ENERGY:

Energy of a body is its capacity to do work. There are two basic forms of energy.

i.   Kinetic energy
ii.   Potential energy

The **kinetic energy** is possessed by a body due to its motion and is given by the formula:

$$K.E. = \frac{1}{2}mv^2 \quad\quad .... \quad 5.7$$

where m is, the mass of the body moving with velocity v.

The **potential energy** is possessed by a body because of its position in a force field e.g. gravitational field or because of its constraint state. The potential energy due to gravitational field near the surface of the Earth at a-height h is given by the formula:

$$P.E. = mgh \quad\quad .... \quad 5.8$$

This is called gravitational potential energy. The gravitational P.E. is always determined relative to some arbitrary position which is assigned the value of zero P.E. In the present case, this reference level is the surface of the Earth as position of zero P.E. In some cases a point at infinity from the Earth can also be chosen as zero reference level.

The energy stored in a compressed spring is the potential energy possessed by the spring due to its compressed or stretched state. This form of energy is called elastic potential energy.

5.3.1 Work-Energy Principle:

Work is done on a body; it increases its energy. For example, a body of mass m is moving with velocity v_i, A force F acting through a distance d increases the velocity to v_f, then from equation of motion.

$$2ad = v_f^2 - v_i^2$$

or $\quad d = \dfrac{1}{2}(v_f^2 - v_i^2) \qquad$ 5.9

From second law of motion

$$F = ma \qquad$$ 5.10

Multiplying equations 5.9 and 5.10, we have

$$Fd = \dfrac{1}{2}m(v_f^2 - v_i^2)$$

or $\quad Fd = \dfrac{1}{2}mv_f^2 - \dfrac{1}{2}mv_i^2 \qquad$ 5.11

Where the left-hand side of the above equation gives the work done on the body and right hand side gives the increase ore change in kinetic energy of the body. Thus

> *Work done on the body equals the change in its kinetic energy.*

This is known as the work-energy principle. If a body is raised up from the Earth's surface, the W0l'k done changes the gravitational potential energy. Similarly, if at ' spring is compressed, the work done on it equals the increase in its elastic potential energy.

5.3.2 INTERCONVERTION OF POTENTIAL ENERGY AND KINETIC ENERGY:

Consider a body of mass m at rest, at a height h above the surface of the Earth as shown in Fig. 5.9. At position, A the body has P.E. = mgh and K.E. = 0. We release the

body and as it falls, we can examine how kinetic and potential energies associated with it interchange.

Let us calculate P.E and K.E. at position B when the body has fallen through a distance x, ignoring air friction.

$$P.E. = mg\,(h-x)$$

and $$K.E. = \frac{1}{2}mv_B^2$$

Velocity v_B, at B, can be calculated from the relation,

$$v_f^2 = v_i^2 + 2gS$$

$$v_f = v_B \quad , \quad v_i = 0 \qquad S = x$$

$$v_B^2 = 0 + 2gx = 2gx$$

$$K.E. = \frac{1}{2}m\,(2gx) = mgx$$

Total energy at B = P.E. + K.E.

$$= mg(h-x) + mgx = mgh \quad \ldots\ldots\ 5.12$$
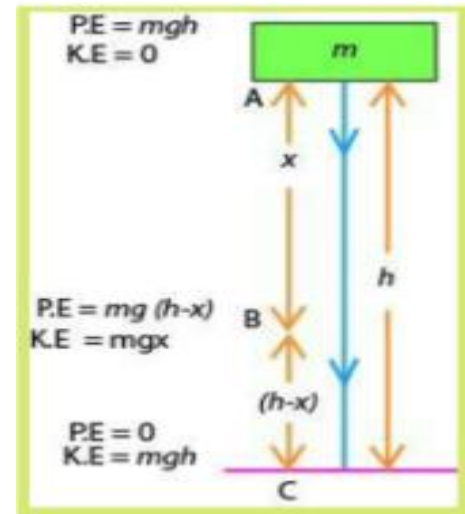
At position C, just before the body strikes the Earth, P.E. = 0- and K.E. = $1/2mv_c^2$ of, where v_c can be found out by the following expression.

$$v_c^2 = v_i^2 + 2gh = 2gh \quad as\ v_i = 0$$

i.e. $$K.E. = \frac{1}{2}mv_c^2 = \frac{1}{2}m \times 2gh = mgh$$

Thus at point C, kinetic energy is equal to the original value. of the potential energy of the body. Actually when a body falls, its velocity increases i.e., the body is being accelerated under the action of gravity. The increase of velocity results in the increase in its kinetic energy. On the other hand, as the body falls, its height decreases and hence, its potential energy also

Fig. 5.9
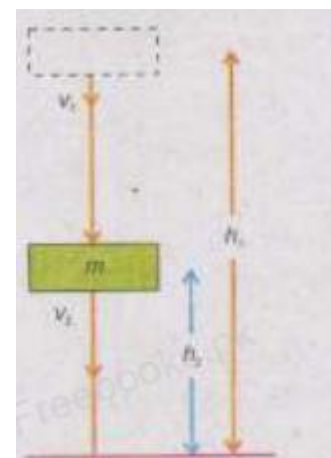
Fig. 5.10

decreases. Thus we see (Fig. 5.10) that,

Loss in P.E. = Gain in K.E.

$$mg(h_1 - h_2) = \frac{1}{2}m(v_2^2 - v_1^2)$$ ·········· 5.13

Where v_1 and v_2 are velocities of bodies at heights of h_1 and h_2 respectively. This result is true only when frictional force is not considered. If we assume that a frictional force f is present during the downward motion, then a part of P.E. is used in doing Work against friction equal to fh. The remaining P.E. = mgh -fh is converted into K.E.

Hence, $$mgh - fh = \frac{1}{2}mv^2$$

or $$mgh = \frac{1}{2}mv^2 + f\,h$$ ········ 5.14

Thus,

Loss in P.E. = Gain in K.E. + Work done against friction.

5.3.3 CONSERVATION OF ENERGY

The kinetic and potential energies are different forms of same basic quantities i.e; mechanical energy of a body is the sum of the kinetic energy and potential energy.in our previous discussion of a falling body, potential energy may change into kinetic energy and vice versa, but the total energy remains constant. Mathematically,

Total Energy =P.E + K.E= Constant

This is a special case of the law of conservation of energy which states that

Energy cannot be destroyed. It can be transformed from one kind into another, but the total amount of energy remains constant.

This is one of the basic laws of physics. We observe many energy transformations from one form to another. Some forms, such as electrical arid chemical energy, are more easily transferred than others, such as heat. Ultimately all energy transfers result in the heating of the environment and energy is wasted. For example, the P.E. of the falling object changes to K.E., but on striking the ground, the K.E. changes into heat and sound. If it seems in an energy transfer that some energy has disappeared, the lost" energy is

often converted into heat. This appears to be the fate of all available energies and is one reason why new sources of useful energy must be developed.

EXERCISE

i. A Person holds a bag of groceries while standing still and talking to a friend. A car is stationary with its engine running. From the standpoint of work, how are these two situations similar?

ii. Calculate the work done in kilo joules in lifting a mass of 10 kg (at a steady velocity) through a vertical height of 10 m.

iii. A force F acts through a distance L. The force is then increased to 3 F and then acts through a further distance of 2 L. Draw the work diagram to scale.

iv. In which case is more work done? When a 50 kg bag of books is lifted through 50 cm, or when a 50 kg crate is pushed through 2m across the floor with a force of 50 N?

v. An object has 1 J of potential energy. Explain what does it mean.

vi. A ball of mass m is held at a height of h_1 above a table. The tabletop is at a height of h_2 above the floor. One student says that the ball has potential energy mgh_1 but another says that it is $mg(h_1 + h_2)$. Who is correct?

vii. When a rocket re-enters the atmosphere, its nose cone becomes very hot. Where does this heat energy come from?

viii. What sort of energy is in the following:
a) Compressed spring
b) Water in a high dam
c) A moving car '

(ix) A girl drops a cup from a certain height, which breaks into pieces. What energy changes are involved?

Chapter # 6

Simple machines

Students will be able to understand:

- Simple machines
- Examples of simple machines along with their principals.
- Principles of machines
- Friction

6.1 Simple machine:

Any of several devices with few or no moving parts that are used to modify motion and the magnitude of a force in order to perform work. They are the simplest mechanisms known that can use leverage (or mechanical advantage) to increase force. The simple machines are the inclined plane, lever, wedge, wheel and axle, pulley, and screw.

6.1.1The inclined plane:
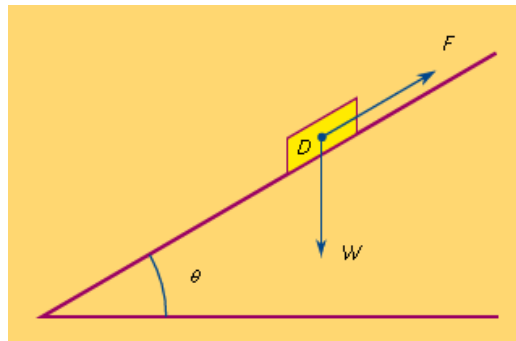
In this representation of an inclined plane, *D* represents a block to be moved up the plane, *F* represents the force required to move the block, and *W* represents the weight of the block. Expressed mathematically, and assuming the plane to be without friction, $F = W \sin \theta$.(more)

An inclined plane consists of a sloping surface; it is used for raising heavy bodies. The plane offers a mechanical advantage in that the force required to move an object up the incline is less than the weight being raised (discounting friction). The steeper the slope, or incline, the more nearly the required force approaches the actual weight. Expressed mathematically, the force *F* required to move a block *D* up an inclined plane without friction is equal to its weight *W* times the sine of the angle the inclined plane makes with the horizontal (θ). The equation is $F = W \sin \theta$.

The principle of the inclined plane

It is used widely—for example, in ramps and switchback roads, where a small force acting for a distance along a slope can do a large amount of work.

6.1.2The lever:
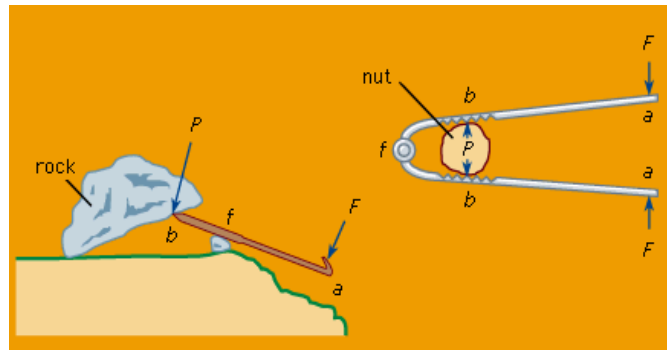
Fig: Lever

Two examples of levers (Left) A crowbar, supported and turning freely on a fulcrum *f*, multiplies a downward force *F* applied at point *a* such that it can overcome the load *P* exerted by the mass of the rock at point *b*. If, for example, the length *af* is five times *bf*, the force *F* will be multiplied five times. (Right) A nutcracker is essentially two levers connected by a pin joint at a fulcrum *f*. If *af* is three times *bf*, the force *F* exerted by hand at point *a* will be multiplied three times at *b*, easily overcoming the compressive strength *P* of the nutshell.

A lever is a bar or board that rests on a support called a fulcrum. A downward force exerted on one end of the lever can be transferred and increased in an upward direction at the other end, allowing a small force to lift a heavy weight.

6.1.3 The wedge:

Fig: Wedge

Wedge used for splitting wood.

A wedge is an object that tapers to a thin edge. Pushing the wedge in one direction creates a force in a sideways direction. It is usually made of metal or wood and is used for splitting, lifting, or tightening, as in securing a hammer head onto its handle.

The wedge was used in prehistoric times to split logs and rocks; an ax is also a wedge, as are the teeth on a saw. In terms of its mechanical function, the screw may be thought of as a wedge wrapped around a cylinder.

6.1.4 The wheel and axle:

A wheel and axle is made up of a circular frame (the wheel) that revolves on a shaft or rod (the axle). In its earliest form it was probably used for raising weights or water buckets from wells.
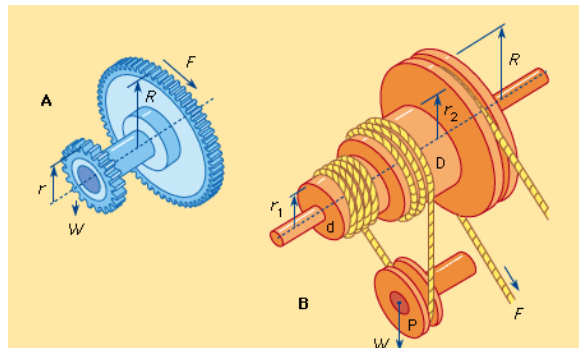
Fig: Wheel and axle

3.5.2.8.   Wheel and axle arrangements:
Two wheel and axle arrangements (A) with a large gear and a small gear attached to the same shaft, or axle, a force F applied at the radius R on the large gear is sufficient to overcome the larger force W at the radius r on the small gear, turning the axle. (B) In a drum and rope arrangement capable of rising weights, a large drum of radius R can be used to turn a small drum. An increase in mechanical advantage can be obtained by using the large drum to turn a small drum with two radii as well as a pulley block. When a force F is applied to the rope wrapped around the large drum, the rope wrapped around the small two-radius drum winds from d (radius r_1) and onto D (radius r_2). The force W on the radius of the pulley block P is easily overcome, and the attached weight is lifted.

PRINCIPLE:

Its principle of operation is best explained by way of a device with a large gear and a small gear attached to the same shaft. The tendency of a force, F, applied at the radius R on the large gear to turn the shaft is sufficient to overcome the larger force W at the radius r on the small gear. The force amplification, or mechanical advantage, is equal

to the ratio of the two forces (W:F) and also equal to the ratio of the radii of the two gears (R:r).

If the large and small gears are replaced with large- and small-diameter drums that are wrapped with ropes, the wheel and axle become capable of raising weights. The weight being lifted is attached to the rope on the small drum, and the operator pulls the rope on the large drum. In this arrangement the mechanical advantage is the radius of the large drum divided by the radius of the small drum. An increase in the mechanical advantage can be obtained by using a small drum with two radii, r_1 and r_2, and a pulley block. When a force is applied to the large drum, the rope on the small drum winds onto D and off of d.

A measure of the force amplification available with the pulley-and-rope system is the velocity ratio, or the ratio of the velocity at which the force is applied to the rope (V_F) to the velocity at which the weight is raised (V_W). This ratio is equal to twice the radius of the large drum divided by the difference in the radii of the smaller drums D and d. Expressed mathematically,

The equation is $V_F/V_W = 2R/(r_2 - r_1)$

The actual mechanical advantage W/F is less than this velocity ratio, depending on friction. A very large mechanical advantage may be obtained with this arrangement by making the two smaller drums D and d of nearly equal radius.

6.1.5 The pulley:
A pulley is a wheel that carries a flexible rope, cord, cable, chain, or belt on its rim. Pulleys are used singly or in combination to transmit energy and motion. Pulleys with grooved rims are called sheaves. In belt drive, pulleys are affixed to shafts at their axes, and power is transmitted between the shafts by means of endless belts running over the pulleys.
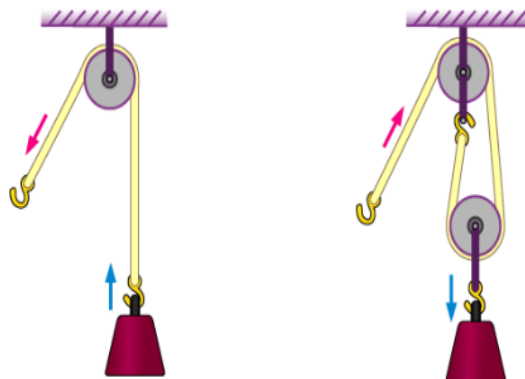
Fig : The Pulley

6.1.6 Block and tackle:

One or more independently rotating pulleys can be used to gain mechanical advantage, especially for lifting weights. The shafts about which the pulleys turn may affix them to frames or blocks, and a combination of pulleys, blocks, and rope or other flexible material is referred to as a block and tackle. The Greek mathematician Archimedes (3rd century BCE) is reported to have used compound pulleys to pull a ship onto dry land.

6.1.7 The screw:

A screw is a usually circular cylindrical member with a continuous helical rib, used either as a fastener or as a force and motion modifier.
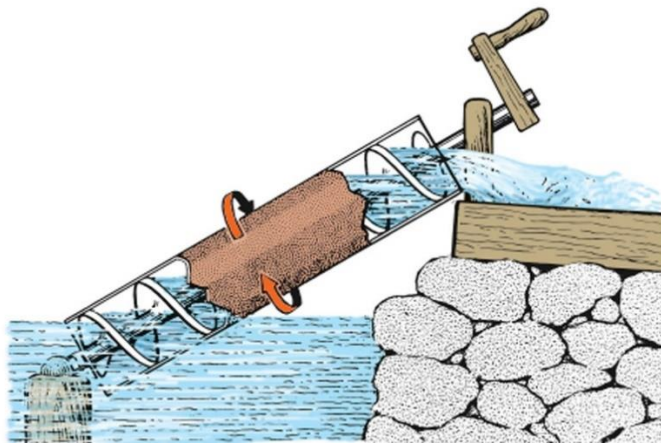
6.1.8 Archimedes screw:

The exact period of its first appearance as a useful mechanical device is obscure. The invention of the water screw is usually ascribed to Archimedes, but evidence exists of a similar device used for irrigation in Egypt at an earlier date. The screw press, probably invented in Greece in the 1st or 2nd century BCE, has been used since the days of the Roman Empire for pressing clothes. In the 1st century CE, wooden screws were used in wine and olive-oil presses, and cutters (taps) for cutting internal threads were in use.
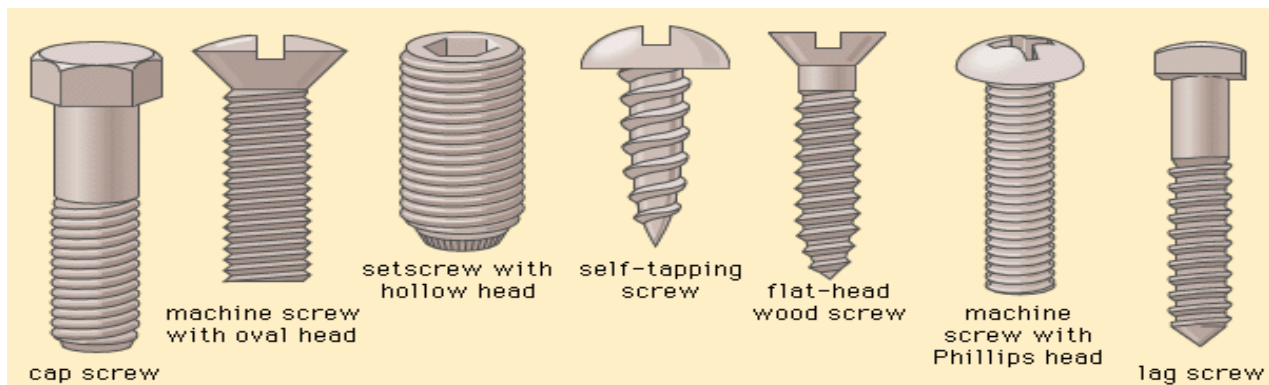
6.1.8.1   Types:

(A) Cap screw

(B) Machine screw with oval head

(C) Setscrew with hollow head

(D) Self-tapping screw

(E) flat-head wood screw

(F) Machine screw with Phillips head

(G) Lag screw.

(H)Screws and screw heads

Cap and machine screws are used to clamp machine parts together, either when one of the parts has a threaded hole or in conjunction with a nut. These screws stretch when tightened, and the tensile load created clamps the parts together. The setscrew fits into a threaded hole in one member; when tightened, the cup-shaped point is pressed into a mating member (usually a shaft) and prevents relative motion. Self-tapping screws form or cut mating threads by displacing material adjacent to a pilot hole so that it flows around the screw.

Wood screws are made in a wide variety of diameters and lengths; when using the larger sizes, pilot holes are drilled to avoid splitting the wood.

Lag screws are large wood screws used to fasten heavy objects to wood. Heads are either square or hexagonal.

Screws that modify force and motion are known as power screws. A screw jack converts torque (turning moment) to thrust. The thrust (usually to lift a heavy object) is created by turning the screw in a stationary nut. By using a long bar to turn the screw, a

small force at the end of the bar can create a large thrust force. Workpiece tables on machine tools are moved linearly on guiding ways by screws that rotate in bearings at the ends of the tables and mate with nuts fixed to the machine frame. A similar torque-to-thrust conversion can be obtained by either rotating an axially fixed screw to drive a rotationally fixed nut along the screw or by rotating an axially fixed nut to drive a rotationally fixed screw through the nut.

6.2 Principle of a Machine:

The work output of a machine is equal to the work input. All machines work on the principle that when the effort is smaller than the load, it must move a greater distance to achieve the same work.

6.3 Friction:

Friction is an external force that opposes the relative motion of two contact areas. Friction occurs at the point of contact between the two bodies.

Relative motion occurs whenever one item travels relative towards another.

6.3.1 Causes of friction:

Friction is caused by the surface imperfections of the two objects throughout contact: While seen at a microscopic level, all surfaces feature hills and valleys that interconnect whenever moving as well as rubbing against one another. Surface imperfections are the name given to such surface abnormalities. Rough surfaces have more irregularities, whereas smooth surfaces have fewer irregularities.

Adhesive forces across surfaces of contact: Whenever two surfaces come into touch, bonds develop, and the surfaces start to stay collectively. It is known as the Adhesion phenomenon. Whether individuals try to stack items on top of one another, we eventually weaken the bonds as well as defeat the adhesive forces.

Plowing effect: When delicate or readily deformed coatings come into touch with another substance, they get distorted.

Two simple experimental facts characterize the friction of sliding solids.

First, the amount of friction is nearly independent of the area of contact. If a brick is pulled along a table, the frictional force is the same whether the brick is lying flat or standing on end.

Second, friction is proportional to the load or weight that presses the surfaces together. If a pile of three bricks is pulled along a table, the friction is three times greater than if one brick is pulled.

Thus, the ratio of friction F to load L is constant. This constant ratio is called the coefficient of friction and is usually symbolized by the Greek letter mu (μ). Mathematically,

$\mu = F/L$.

Because both friction and load are measured in units of force (such as pounds or newtons), the coefficient of friction is dimensionless. The value of the coefficient of friction for a case of one or more bricks sliding on a clean wooden table is about 0.5, which implies that a force equal to half the weight of the bricks is required just to overcome friction in keeping the bricks moving along at a constant speed. The frictional force itself is directed oppositely to the motion of the object. Because the friction thus far described arises between surfaces in relative motion, it is called kinetic friction.

6.3.2 Types of Friction:

There are four types of friction, and they are classified as follows:

a) Static friction
b) Sliding friction
c) Rolling friction
d) Fluid friction

Static friction, in contrast, acts between surfaces at rest with respect to each other. The value of static friction varies between zero and the smallest force needed to start motion. This smallest force required to start motion, or to overcome static friction, is always greater than the force required to continue the motion, or to overcome kinetic friction.

Rolling friction occurs when a wheel, ball, or cylinder rolls freely over a surface, as in ball and roller bearings. The main source of friction in rolling appears to be dissipation of energy involved in deformation of the objects.

If a hard ball is rolling on a level surface, the ball is somewhat flattened, and the level surface somewhat indented in the regions in contact. The elastic deformation or compression produced at the leading section of the area in contact is a hindrance to motion that is not fully compensated for as the substances spring back to normal shape at the trailing section. The internal losses in the two substances are similar to those that keep a ball from bouncing back to the level from which it is dropped. Coefficients of sliding friction are generally 100 to 1,000 times greater than coefficients of rolling friction for corresponding materials. This advantage was realized historically with the transition from sledge to wheel.

6.3.3 Applications of Friction:

i. Friction finds application when matchsticks are ignited.
ii. Motion of pistons in a cylinder is an application of friction.

iii.   It is possible to write on books and board as there is friction between pen and the board.

Friction Problem with Solution

Example 6.1. Block A is kept on top of block B; the coefficient of static friction between A and B is 0.6 and between B and ground is 0.5. Also, the coefficient of kinetic friction between A and B is 0.4 and between B and ground is 0.3. If a force of 60 N is applied on block B, Find the acceleration of both the blocks. Mass of block A is 5 kg, and Block B is 10 Kg.

Solution:
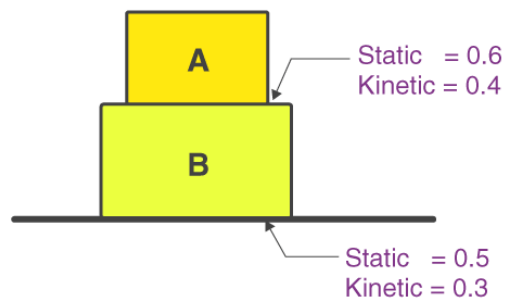
The first step should always be to draw the F.B.D of the given setup:
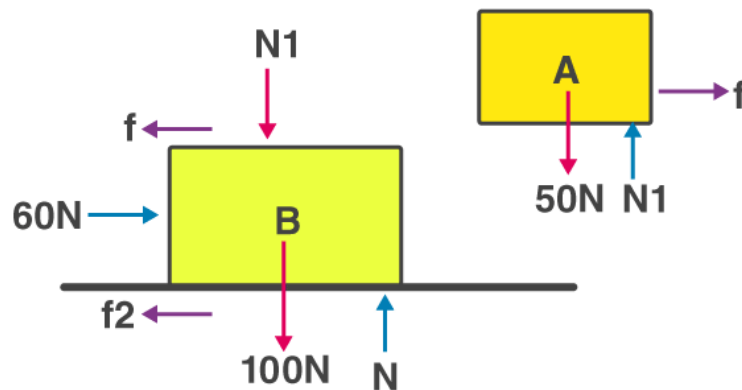
Let's consider that both blocks are moving with the same acceleration, then A + B can be considered as a system, so the force equation in the horizontal direction will be,

60–f1 = 15a ――― (1)

Since block B is moving over the surface, hence it is sliding friction.

Therefore, in this case we will use coefficient of friction as 0.3, f1 = 0.3 × 15 × 10 N f1 = 45 N —— (2)

From (1) and (2)

we have, a = 1m/s^2 from F.B.D of block A, f = 5.a

Putting the value of a, we get, f = 5 N

Maximum value of static friction between A and B = 0.6 × 5 × 10 = 30 N

As f lies within the range, the two blocks can move together with an acceleration of 1 m/s^2. We can also be asked to find the maximum force applied to block B so that both blocks move together. In that case, since both blocks will move together, there will be static friction between the two, and hence the maximum value that can act is 30 N.

Therefore, acceleration of B, a = 305 m/s^2 = 6 m/s^2

For Block A,   F – 45 – 30 = 10 × 6 F = 135 N

So, if a force of more than 135 N is applied, both blocks will have different accelerations.

EXERCISE

Give short answers.

i. How does friction produce heat?
ii. How is friction useful?
iii. Why is friction a non-conservative force?
iv. Will friction increase as the speed increases?
v. Can friction be zero?
vi. What is a simple machine?
vii. Name the six simple machines.
viii. A bottle cap is an example of which simple machine?
ix. Is Broom a simple machine?
x. What type of simple machine is a doorknob?
xi. What are compound machines?

Chapter # 7

Density

Students will be able to understand: • Density • Units of density • Specific gravity • Archimedes' Principle. • Different applications

7.1 Density:

Density is a measurement that compares the amount of matter an object has to its volume. An object with much matter in a certain volume has a high density. In this article, we will learn about the definition of density and the units of density.

The density of material shows the denseness of that material in a specific given area. A material's density is defined as its mass per unit volume. Density is essentially a measurement of how tightly matter is packed together. It is a unique physical property of a particular object. The principle of density was discovered by the Greek scientist Archimedes. It is easy to calculate density if you know the formula and understand the related units The symbol ρ represents density or it can also be represented by the letter D.

Definition: Density is the measurement of how tightly a material is packed together. It is defined as the mass per unit volume.
Density Symbol: D or ρ
Density Formula: $\rho = m/V$, Where ρ is the density, m is the mass of the object and V is the volume of the object.

7.1.1.Examples:

Iron, platinum, and lead are examples of dense materials. Many types of rock and minerals are examples of dense material. Dense materials are most likely to 'feel' heavy or hard. The opposite of dense is sparse and a few examples of sparse material are glass, bamboo, aluminum, and Styrofoam.

In general, liquids are less dense than solids and gases are less dense than liquids. This is because solids have densely packed particles, liquids are materials where particles can slide around one another, and gases have particles that are free to move all over the place.

7.1.2 How Is Density Calculated?

Mathematically, the density of an object is expressed as follows:

$$Density = \frac{Mass}{Volume}$$

$$\rho = \frac{m}{v}$$

7.1.3 Unit of Density:

Though the SI unit of density is kg/m³, for convenience we use g/cm³ for solids, g/ml for liquids, and g/L for gases.

Density can be explained as the relationship between the mass of the substance and the volume it takes up.

In a qualitative term, it shows how heavy an object is at constant volume.

Different substances have different densities, which means for the same volume of different substances weigh differently.

7.1.4 SI Unit of Density:

Each substance has a specific density. Generally, the density of water (which is approximately about 1 gram/cubic centimetre) is taken as the standard value for calculating the density of substances. However, the SI unit of Density is measured using kilograms per cubic metre (*kg/m³*).

7.1.5 Other Density Units:

Talking about other density units, metric tons and liters are also used even though they are not part of the SI. Some other units include:

    i.   gram per milliliter (g/mL)
    ii.   metric ton per cubic meter (t/m^3)
    iii.   kilogram per liter (kg/L)
    iv.   mega gram (metric ton) per cubic meter (mg/m^3)
    v.   gram per cubic centimeter (g/cm^3)
         1 g/cm^3 = 1000 kg/m^3
    vi.   kilogram per cubic decimeter (kg/dm^3)

In addition to this, in the cgs system density is measured in grams per cubic centimeter (g/cm^3).

7.1.6 Applications of Density in Real Life:

Many applications of density are there in our real-life, life a few examples are:

    i.   In pipe design, shipbuilding, helium balloons, weight distribution in the Aero plan, and the fact that ice floats on water.
    ii.   The knowledge of the densities of two substances helps you in separation techniques. For example, the separation of oil from water. If there is a leakage of an oil tank in the ocean, then oil drops start to float on the water due to less density than the water.
    iii.   Another well-known application of density is determining whether an object will float on water or not. The floating of ships and diving of submarines are due to their density difference.

7.2 Specific Gravity:

Have you wondered why some objects sink while others float? Many factors come into play in determining if an object will sink or float, and specific gravity is one of the main deciding factors.

Specific Gravity or relative gravity is a dimensionless quantity that is defined as the ratio of the density of a substance to the density of the water at a specified temperature and is expressed as

$$SG = \frac{\rho_{substance}}{\rho_{H_2O}}$$

It is common to use the density of water at 4 °C as a reference point as water at this point has the highest density of 1000 kg/m^3.

Definition:

Specific gravity or relative gravity is a dimensionless quantity that is defined as the ratio of the density of a substance to the density of water at a specified temperature and pressure

$$\text{Specific gravity} = \frac{\text{Density of the object}}{\text{Density of water}} = \frac{\rho_{object}}{\rho_{H_2O}}$$

Given below is a table that lists the value of specific gravity for some common materials.

Material	Specific Gravity
Dry Air	0.0013
Alcohol	0.82
Carbon dioxide	0.00126
Cast Iron	7.20
Petrol	0.72
Rubber	0.96
Wood Oak	0.77
Nylon	1.12

7.2.1 Calculating Specific Gravity:

Specific gravity is determined by dividing the density of a material by the density of water at 4 degrees Celsius. For the calculation, the density of the material and that of the water must be expressed in the same units.

Density of Water in Various Units.

• 1000 kilograms per cubic meter.

• 1 gram per cubic centimeter.

• 62.43 pounds per cubic foot.

• 0.036 pounds per cubic inch.

7.2.2 Specific Gravity of Gas:

For gases, the specific gravity is normally calculated with reference to air. Specific gravity for gases is defined as the ratio of the density of the gas to the density of air at a

specified temperature and pressure. The density of air at room temperature is 1.20 kg/m^3.

$$SG = \frac{\rho_{gas}}{\rho_{air}}$$

7.2.3 Factors Affecting Specific Gravity:

As discussed, specific gravity is the measure of the density of an object in comparison to the density of water at an exact temperature and pressure. Changes in the pressure and temperature affect the density of objects and liquids, thereby affecting the specific gravity of objects and liquids. Standard temperature and pressure are used in the calculation of specific gravity. If failed to regulate the outside influences, then the specific gravity will change.

7.2.4 Applications of Specific Gravity:

Specific gravity has a wide range of applications, following are a few of the applications:

i.   The gem purity can be determined by comparing its specific gravity with the already measured high purity level of another gem. This allows the gem's value to be determined very quickly.
ii. Specific gravity is used by mineralogists and geologists to determine the mineral content of the rock.
iii. Urinary specific gravity is a measure of the concentration of solutes in the urine. The test is a routine part of a urinalysis.

7.3 Archimedes Principle

Archimedes' principle deals with the forces applied to an object by fluids surrounding it. This applied force reduces the net weight of the object submerged in a fluid. In this article, let us familiarize ourselves with Archimedes' principle.

Archimedes' principle states that:

"The upward buoyant force that is exerted on a body immersed in a fluid, whether partially or fully submerged, is equal to the weight of the fluid that the body displaces and acts in the upward direction at the center of mass of the displaced fluid".

The value of thrust force is given by the Archimedes law which Archimedes of Syracuse of Greece discovered. When an object is partially or fully immersed in a liquid, the apparent loss of weight is equal to the weight of the liquid displaced by it.

If you look at the figure, the weight due to gravity is opposed by the thrust provided by the fluid. The object inside the liquid only feels the total force acting on it as the weight. Because the actual gravitational force is decreased by the liquid's up thrust, the object feels as though its weight is reduced. The apparent weight is thus given by:
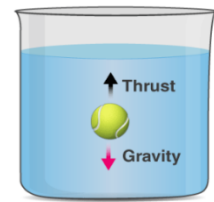
Archimedes' Principal Explanation:

Archimedes' principle tells us that the weight loss is equal to the weight of liquid the object displaces.

7.3.1 Archimedes' Principle Formula:

In simple form, the Archimedes law states that the buoyant force on an object is equal to the weight of the fluid displaced by the object. Mathematically written as:

Apparent weight= Weight of object (in the air) – Thrust force (buoyancy)

$$F_b = \rho \times g \times V$$

Where F_b is the buoyant force, ρ is the density of the fluid, V is the submerged volume, and g is the acceleration due to gravity.

7.3.1.1 Archimedes' Principle Derivation:
We know that the density is defined as

$$Density\ (\rho) = \frac{Mass\ (M)}{Volume\ (V)}$$

Therefore, the mass of the displaced liquid can be written as follows:

$$Mass\ (M) = Density\ (\rho) Volume\ (V)$$

Now, the weight of the displaced liquid can be calculated as follows:

$$Weight = Mass \times Acceleration\ due\ to\ gravity$$
$$Weight = Mass \times g = \rho \times V \times g$$

From Archimedes' principle, we know that the apparent loss of weight is equal to the weight of the water displaced therefore the thrust force is given by the following equation:

$$Thrust\ Force = \rho \times V \times g$$

Where ρ is the density of the liquid, V is the volume of liquid displaced and g is the acceleration due to gravity.

The thrust force is also called the buoyant force because it is responsible for objects floating. Thus, this equation is also called the law of buoyancy.

Archimedes' Principle Experiment:

Take a container filled with water to the brim.

Now take any solid object you like and measure its

weight by using a spring balance. Note this down.

Keep the object attached to the spring balance and s sure the spring balance is not submerged.
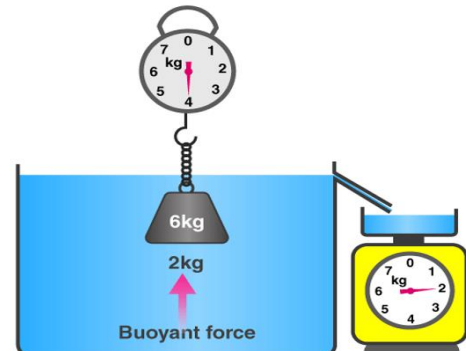
Now, note down the weight shown by the spring balance. You will notice that it is less. Some water will be displaced into the bowl.

Collect this water and weigh it. You will find that the weight of the water will be exactly equal to the loss of weight of the object!

7.3.2 Archimedes' Principle Applications:

Following are the applications of Archimedes' principle:

i. Submarine:

The reason why submarines are always underwater is that they have a component called ballast tank which allows the water to enter making the submarine be in its position underwater as the weight of the submarine is greater than the buoyant force.

ii. Hot-air balloon:

The reason why hot-air balloons rise and float in mid-air is because the buoyant force of the hot-air balloon is less than the surrounding air. When the buoyant force of the hot-air balloon is more, it starts to descend. This is done by varying the quantity of hot air in the balloon.

iii. Hydrometer:

A hydrometer is an instrument used for measuring the relative density of liquids. Hydrometer consists of lead shots which makes them float vertically on the liquid. The lower the hydrometer sinks, the lesser is the density of the liquid.

EXERCISE

1. Take two boxes that have the same volume. Fill the first box with x balls and the second box with 6x balls. If the mass of each ball is the same, which box would weigh more?

2. Calculate the density of water if it has a mass of 1160 Kg and a volume of 1 m^3?

<div align="right">(Ans:1160 kg/m^3)</div>

3. If you find a shiny rock, a carbon allotrope with a volume of 0.042 cm^3 and a mass of 0.14 g, is it graphite or diamond? The density of graphite is 2.266 g/cm^3 and the density of diamond is 3.51g/cm^3.

<div align="right">Ans:(0.0951g for graphite)
(0.1474 g for diamond)</div>

4. What is Density?
5. Who discovered the principle of Density?
6. How would you find the density of a human body?
7. How will we know if a substance is less dense than water?
8. What happens to the least dense of two immiscible liquids?
9. If the density of gold is 19300 kg/m^3, what is the specific gravity of gold?
10. If the specific gravity of ice is 0.92, then what is the density of ice?
11. What does Specific Gravity tell us?
12. What is the difference between specific gravity and specific weight?
13. Which metal has the highest specific gravity?
14. Which metal has the lowest specific gravity?
15. Which instrument measures the specific gravity of liquids?
16. What is a urine specific gravity test?
17. What is specific gravity in a gemstone?
18. Which is the temperature at which water has the highest density?
19. What is the specific gravity of rubber?
20. What is the main application of specific gravity?
21. What does Archimedes Principle state?
22. Who discovered the Archimedes' Principle?
23. How does the Archimedes' principle apply to ships?
24. Where is the Archimedes' principle used?
25. How can the Archimedes' Principle be used to determine the density?

Chapter # 8

Pressure

Students will be able to understand:

- Pressure
- Factors effecting pressure
- Water pressure
- Bernoulli's Equation:

8.1 Pressure:

Pressure can be defined as the force acting normally on unit area of a surface.

Mathematically;

$$Pressure = Force / Area$$

If pressure is denoted by P, force is denoted by F and area is denoted by A, then the above relation can be expressed as:

$$P = F / A$$

8.1.1 Units of Pressure

There are various units to describe Pressure, some of which we will discuss further in this article.

The SI unit of pressure is the Pascal (Pa).

A Pascal can be defined as a force of one newton applied over a surface area of a one-meter square.

8.1.2 Factors Affecting Pressure

    I.   **Force:**

Pressure is dependent on the area over which the force is acting, the pressure can be increased and decreased without any change in the force. The force applied to be constant if the surface becomes smaller the pressure increases and vice versa.

For example, a brick sitting on a surface exerts a force equal to its weight on the object it is resting on. Now we know that a rectangular brick has a wide surface and a thin surface on the sides.

II.   Surface area:

By changing the orientation of the brick resting on a surface, we are effectively changing the pressure acting on the surface by the same brick. See the image below for more information.
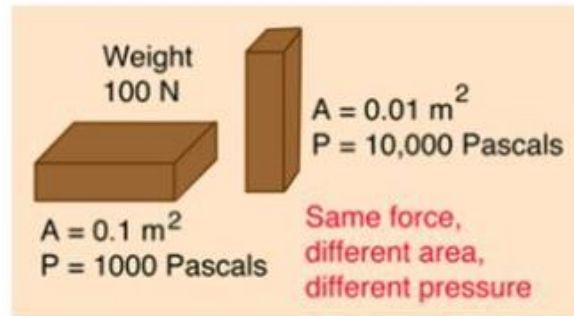
In other words, **If the surface becomes smaller, the pressure becomes larger**.

It is for this reason that our knives and nails are so sharp. A knife distributes the force over its entire cutting edge. The sharper the edge, higher the pressure, and consequently the cutting with a sharp knife is easy. In a blunt knife, the force is distributed over its blunt surface with a larger surface area. Therefore, we need to put more force in order to cut. Therefore, a knife is best when at its sharpest.

For the same reason—that is, that reduction of surface area increases net pressure—an expertly delivered karate chop is much more damaging and deadly than an open-handed slap. When you slap someone, the force you apply in slapping the surface is distributed all over the palm of your hand. In contrast, a karate chop concentrates all the force on the sides of your hand which has significantly lesser surface area than your palms. This leads to a greater application of pressure on the surface thereby rendering a karate chop deadlier than a slap.

8.2 Water Pressure

Pressure is defined as the force applied which is perpendicular to the surface of the object per area over which the force is distributed. The Gauge pressure is the pressure relative to ambient pressure. water pressure is the term used to describe the strength of water flow through a channel or pipe.

The water pressure formula and SI unit are given as:

Formula	Water pressure= ρ g h
SI unit	Pa

Where,

P = water pressure in Pa

ρ = density of water in kg.m^{-3}

g = gravitational force in 9.81 m.s^{-2}

h = height in m

The loss of water pressure can also be calculated.

The water pressure loss formula due to height h is given as:

Pressure loss = 0.4335 x h

Where,

h denotes the height

8.2.1 Causes of fluid to flow:
There are three factors that cause fluid to flow:

  i. Viscosity, density, and compressibility
  ii. Speed of the flow
  iii. Shape of the surface

8.2.2 Pressure in hydrostatic fluids
Consider the person in the last row from the bottom. He is clearly carrying far more weight than the man in the front row. The fluid pressure in a column is the same way. The pressure in a fluid column rises in proportion to the depth. And why is there a rise in pressure? Because fluid at a lesser depth must also sustain fluid above it as you go deeper, to define it, we may thus state,

"The pressure at a place within a fluid caused by the fluid's weight is known as fluid pressure".

Calculation of fluid pressure
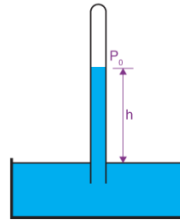
- $P_{fluid} = P + \rho gh$
- where,
  - P = Pressure at the reference point
  - P_{fluid} = Pressure at a point in a fluid
  - ρ = Density of the fluid
  - g = Acceleration due to gravity (considering earth g = 9.8 m/s)
  - h = Height from the reference point
- The density of a fluid may be estimated by dividing its mass by the volume of fluid taken into account.
- $\rho = m/V$

- where,
  - m = Mass of the fluid
  - V = Volume of fluid considered
- If the fluid is subject to atmospheric pressure, then the total pressure on the system is given by
- $P_{fluid} = P_o + \rho gh$
- where,
  - P_o = Atmospheric pressure

8.2.3 Conditions for the Consideration of Fluid Pressure:

Fluid pressure is examined in two situations:

i.   In an open condition or open channel flow
ii. In a closed condition or closed conduit

The fluid pressure listed above is hydrostatic pressure, often known as static fluid pressure. Because the pressure created by fluid movement may be regarded as minor, the pressure is simply taken into consideration here, depending on the depth of the fluid. The static fluid pressure is independent of the liquid's volume, mass, total surface area, and container form.

8.3 Hydraulic lift:

Hydraulic Lift is a device for moving objects using force created by pressure on a liquid inside a cylinder that moves a piston upward. Incompressible oil is pumped into the cylinder, which forces the piston upward. When a valve opens to release the oil, the piston lowers by gravitational force.

The principle for hydraulic lifts is based on Pascal 's law for generating force or motion, which states that pressure change on an incompressible liquid in a confined space is passed equally throughout the liquid in all directions.

The concept of Pascal's law and its application to hydraulics can be seen in the example below, where a small amount of force is applied to an incompressible liquid on the left to create a large amount of force on the right.

Hydraulic systems are used for precision control of large force applications, are economical, and make excellent use of energy resources. A hydraulic system works by applying force at one point to an incompressible liquid, which sends force to a second point. The process involves two pistons that are connected by an oil filled pipe.

The image below is a representation of the two pistons and how they are connected by a pipe.

Two Cylinders of a Hydraulic Device

F 1

A 2

P   F 2

P   A 1

$$\frac{F2 = A2}{F1 = A1}$$
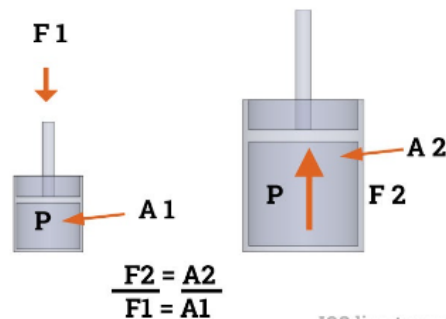
The diagram below represents a simple version of the working mechanism of a hydraulic device. The handle on the right moves the incompressible oil, under pressure, from the reservoir to the high-pressure chamber in the middle of the diagram. The ram moves up as the oil is pumped in.

Force Generated

The force generated in a hydraulic system depends on the size of the pistons. If the smaller of the two pistons is two inches and the larger piston is six inches, or three times as large, the amount of force created will be nine times greater than the amount of force from the smaller piston. One hundred pounds of force by a small piston will be able to lift 900 pounds.

In this diagram, the piston on the left has a one-pound load and an area of one inch. When it moves down ten inches, it can move the ten-pound load on the piston on the right.

Example of Hydraulic Force

F 1

F 2

W1=1 pound

W2=10 pounds

A1=1 sq inch

A2=10 sq inchs

D1=10 inchs

D2=1 inch
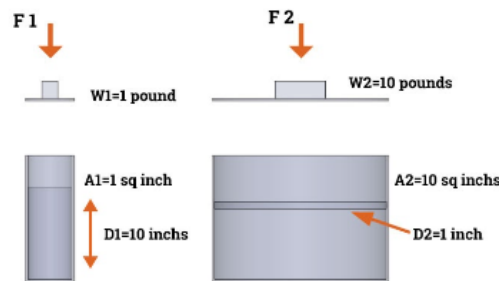
8.4 Bernoulli's Equation:

Bernoulli's principle formulated by Daniel Bernoulli states that as the speed of a moving fluid increases (liquid or gas), the pressure within the fluid decreases. Although Bernoulli deduced the law, it was Leonhard Euler who derived Bernoulli's equation in its usual form in the year 1752.

Bernoulli's principle states that:

"The total mechanical energy of the moving fluid comprising the gravitational potential energy of elevation, the energy associated with the fluid pressure and the kinetic energy of the fluid motion, remains constant".

Bernoulli's principle can be derived from the principle of conservation of energy.

Bernoulli's equation formula is a relation between pressure, kinetic energy, and gravitational potential energy of a fluid in a container.

The formula for Bernoulli's principle is given as follows:

$$p + \tfrac{1}{2}\rho v^2 + \rho g h = constant$$

Where p is the pressure exerted by the fluid, v is the velocity of the fluid, ρ is the density of the fluid and h is the height of the container.

Bernoulli's equation gives great insight into the balance between pressure, velocity and elevation.

8.4.1 Bernoulli's Equation Derivation

Consider a pipe with varying diameter and height through which an incompressible fluid is flowing. The relationship between the areas of cross-sections A, the flow speed v, height from the ground y, and pressure p at two different points 1 and 2 are given in the figure below.
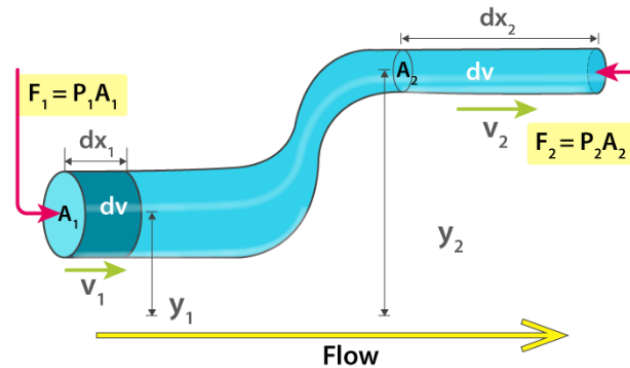
Assumptions:

The density of the incompressible fluid remains constant at both points.

The energy of the fluid is conserved as there are no viscous forces in the fluid.

Therefore, the work done on the fluid is given as:

$$dW = F_1 dx_1 - F_2 dx_2$$

$$dW = p_1 A_1 dx_1 - p_2 A_2 dx_2$$

$$dW = p_1 dv - p_2 dv = (p_1 - p_2) dv$$

We know that the work done on the fluid was due to the conservation of change in gravitational potential energy and change in kinetic energy. The change in kinetic energy of the fluid is given as:

$$dK = \tfrac{1}{2}m_2v_2^2 - \tfrac{1}{2}m_1v_1^2 = \tfrac{1}{2}\rho dv(v_2^2 - v_1^2)$$

The change in potential energy is given as:

$$dU = m_2gy_2 - m_1gy_1 = \rho dvg(y_2 - y_1)$$

Therefore, the energy equation is given as:

$$dW = dK + dU$$

$$(p_1 - p_2)dv = \tfrac{1}{2}\rho dv(v_2^2 - v_1^2) + \rho dvg(y_2 - y_1)$$
$$p_1 - p_2) = \tfrac{1}{2}\rho(v_2^2 - v_1^2) + \rho g(y_2 - y_1)$$

Rearranging the above equation, we get

$$p_1 + \tfrac{1}{2}\rho v_1^2 + \rho gy_1 = p_2 + \tfrac{1}{2}\rho v_2^2 + \rho gy_2$$

This is Bernoulli's equation.

8.4.2 Applications of Bernoulli's Principle and Equation

Bernoulli's principle is used for studying the unsteady potential flow which is used in the theory of ocean surface waves and acoustics. It is also used for approximation of parameters like pressure and speed of the fluid.

The other applications of Bernoulli's principle are:

I.   **Venturi meter:** It is a device that is based on Bernoulli's theorem and is used for measuring the rate of flow of liquid through the pipes. Using Bernoulli's theorem, Venturi meter formula is given as:

$$V = A_1A_2\sqrt{\frac{2hg}{A_1^2 - A_2^2}}$$
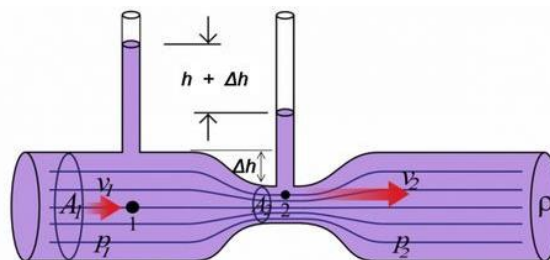
II.   **Working of an aeroplane:** The shape of the wings is such that the air passes at a higher speed over the upper surface than the lower surface. The

difference in airspeed is calculated using Bernoulli's principle to create a pressure difference.

III.   When we are standing at a railway station and a train comes, we tend to fall towards the train. This can be explained using Bernoulli's principle as the train goes past, the velocity of air between the train and us increases. Hence, from the equation, we can say that the pressure decreases. So, the pressure from behind pushes us towards the train. This is based on Bernoulli's effect.

EXERCISE

1. What does hydrostatic pressure mean in simple terms?
2. What causes fluids to exert pressure?
3. What causes fluid to flow?
4. What will increase gas pressure?
5. How does gas pressure change with temperature?

Chapter # 9

Heat

Students will be able to understand:

- Nature and measurements of heat.
- Effects of heat
- Methods of transfer of heat. i.e. conduction ,convection and radiation.

Heat is an important form of energy. It is necessary for our survival. We need it to cook our food and to maintain our body temperature. Heat is also needed in various industrial processes. How to protect ourselves from high as well as low temperatures needs knowledge of how heat travels. In this unit, we will study various ways of heat transfer.

9.1 TRANSFER OF HEAT:
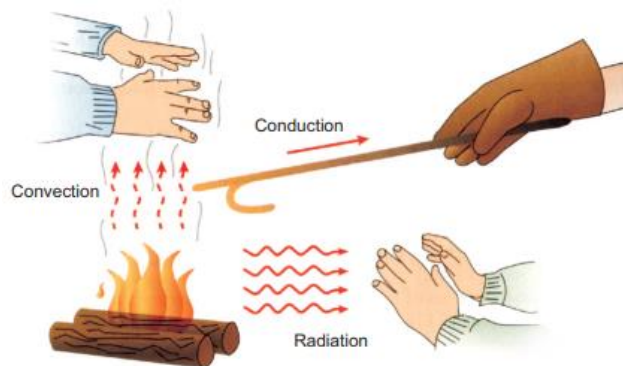
Fig 9.1

When two bodies at different temperatures are in thermal contact with each other . Thermal energy from a hot body flow to a cold body in the form of heat . This is called as transfer of heat. Transfer of heat is a natural process. It continues all the time if the bodies in thermal contact are at different temperatures. There are three ways by which the transfer of heat takes place.

These are:

1. Conduction
2. Convection
3. Radiation

9.1.1 CONDUCTION:

The handle of the metal spoon held in hot water soon gets warm. But in the case of a wooden spoon, the handle does not get warm. Both materials behave differently regarding the transfer of heat. Both metals and non-metals conduct heat. Metal is generally a better conductor than non-metals. In solids, atoms and molecules are packed close together as shown in Figure 9.2. They continue to vibrate about their mean position. What happens when one of its ends is heated? The atoms or molecules present at that end begin to vibrate more rapidly. They also collide with their neighboring atoms or molecules. In doing so, they pass some of their energy to neighboring atoms or molecules during collisions with them with the increase in their vibrations. These atoms or molecules in turn pass on a part of the energy to their neighboring particles. In this way, some heat reaches the other parts of the solids. This is a slow process, and a very small transfer of heat takes place from hot to cold parts in solids. How does the heat flow from hot to cold parts in metals so rapidly than in non-metals? Metals have free electrons as shown in figure 9.3. These free electrons move with very high velocities within the metal objects. They carry energy at a very fast rate from hot to cold parts of the object as they move. Thus, heat reaches the cold parts of the metal objects from its hot part much more quickly than non-metals.
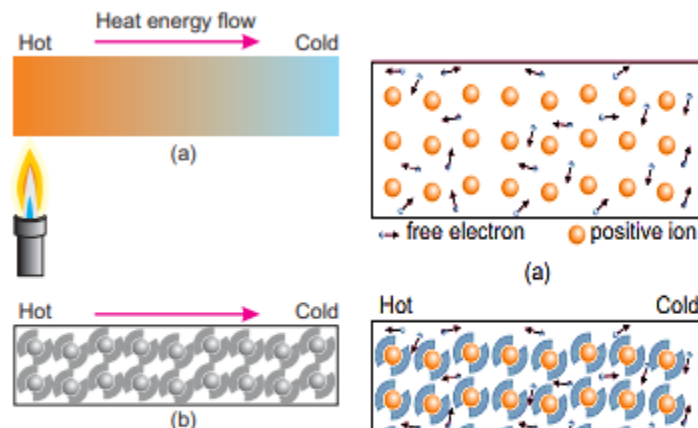
Figure 9.2: In solids, heat is transferred from one part to other parts from atoms to atoms or molecules to molecules due to collisions.

Figure 9.3: Conduction of heat in metals.

The mode of transfer of heat by vibrating atoms and free electrons in solids from hot to cold parts of a body is called conduction of heat.

All metals are good conductors of heat. The substances through which heat does not conduct easily are called bad conductors or insulators. Wood, cork, cotton, wool, glass, rubber, etc. are bad conductors or insulators.

9.1.1THERMAL CONDUCTIVITY

Conduction of heat occurs at different rates in different materials. In metals, heat flows rapidly as compared to insulators such as wood or rubber. Consider a solid block as shown in Figure 9.4. One of its two opposite faces each of cross-sectional area A is heated to a temperature T_1. Heat Q flows along its length L to 1 opposite face at temperature T_2 in t seconds.

> **The amount of heat that flows in unit time is called the rate of flow of heat.**

Thus   Rate of flow of heat $= \dfrac{Q}{t}$ .... (9.1)

It is observed that the rate at which heat flows through a solid object depends upon various factors.

9.1.1.1   CROSS-SECTIONAL AREA OF THE SOLID

Larger cross-sectional area A of a solid contains a larger number of molecules and free electrons on each layer parallel to its cross-sectional area and hence greater will be the rate of flow of heat through the solid. Thus

Rate of flow of heat $\dfrac{Q}{t} \propto A$

9.1.1.2   LENGTH OF THE SOLID

The larger the length between the hot and cold ends of the solid, the more time it will take to conduct heat to the colder end, and the smaller the rate of flow of heat. Thus

Rate of flow of heat $\dfrac{Q}{t} \propto \dfrac{1}{L}$

9.1.1.3   TEMPERATURE DIFFERENCE BETWEEN ENDS

The greater the temperature difference $T_1 - T_2$ between the hot and cold faces of the solid, the greater will be the rate of flow of heat. Thus

Rate of flow of heat $\dfrac{Q}{t} \propto (T_1 - T_2)$

Combining the above factors, we get

$$\dfrac{Q}{t} \propto \dfrac{A(T_1 - T_2)}{L}$$

Rate of flow of heat $\dfrac{Q}{t} = \dfrac{kA(T_1 - T_2)}{L}$ ... ... (9.2)

Here k is the proportionality constant called the thermal conductivity of the solid. Its value depends on the nature of the substance and is different for different materials. From equation (9.2), we find k as:

$$k = \frac{Q}{t} \times \frac{L}{A(T_1 - T_2)} \quad \dots \quad \dots \quad \dots \quad (9.3)$$

Thus, the thermal conductivity of a substance can be defined as:

> **The rate of flow of heat across the opposite faces of a metre cube of a substance maintained at a temperature difference of one kelvin is called the thermal conductivity of that substance.**

9.1.2 USE OF CONDUCTORS AND NON-CONDUCTORS:

In houses, good thermal insulation means lower consumption of fuel. For this, the following measures may be taken to save energy.

• Hot water tanks are insulated by plastic or foam lagging.

• Wall cavities are filled with plastic foam or wool.

• The ceiling of rooms is covered by insulating materials (false ceiling).

• Double-glazed windowpanes are used. These windowpanes have air between glass sheets that provides good insulation.

Good conductors are used when quick transfer of heat is required through a body. Thus cookers, cooking plates, boilers, radiators condensers refrigerators, etc. are made of metals such as aluminum or copper. Similarly, metal boxes are used for making ice, ice cream, etc.

Insulators or bad conductors are used in home utensils such as handles of saucepans, hot plates, spoons, etc. They are made up of wood or plastic. Air is one of the bad conductors or best insulators. That is why cavity walls i.e. two walls separated by an air space and double-glazed windows keep the houses warm in winter and cool in summer. Materials that trap air i.e. wool, felt, fur, feathers, polystyrenes, and fibre glass are also bad conductors. Some of these materials are used for leggings to insulate water pipes, hot water cylinders, ovens, refrigerators, walls, and roofs of houses. Woolen cloth is used to make warm winter clothes.

9.1.2 CONVECTION

Liquids and gases are poor conductors of heat. However, heat is transferred through fluids (liquids or gases) easily by another method called convection. Why is a balloon inflated with hot air as shown in the figure? 9.6 rises? A liquid or a gas becomes lighter

(less dense) as it expands on heating. Hot liquid or gas rises above the heated area. The cooler liquid or gas from the surroundings fills the place which in turn is heated up. In this way, all the fluid is heated up. Therefore, the transfer of heat through fluids takes place by the actual movement of heated molecules from hot to cold parts of the fluid.

Transfer of heat by actual movement of molecules from hot place to a cold place is known as convection.

EXPERIMENT 9.1

Take a beaker and fill two-thirds of it with water. Heat the beaker by keeping a burner below it. Drop two or three crystals of potassium permanganate in the water. It will be seen that colored streaks of water formed by the crystals move upwards above the flame and then move upward from sideways as shown in Fig 9.4, These colored streaks show up the path of the currents in the liquids.

Figure 9.4: Crystals of potassium permanganate are used to show the movement of water on heating.

9.1.3 RADIATION

Our Sun is the major source of heat energy. But how does this heat energy reach the Earth? It reaches us neither by conduction nor by convection because the space between the Sun and the Earth's atmosphere is empty. There is a third mode called radiation by which heat travels from one place to another. It is through radiation that heat reaches us from the Sun.

Radiation is the mode of transfer of heat from one place to another in the form of waves called electromagnetic waves.

How does this heat reach us directly from a fireplace? Figure 9.14 shows a fireplace used for room heating. Heat does not reach us by conduction through air from a fireplace because air is a poor conductor of heat. Heat does not reach us by convection because the air getting heat from the fireplace does not move in all directions. Hot air moves

upward from the fireplace. The heat from the fireplace reaches us directly by a different process in the form of waves called radiation. A sheet of paper or cardboard kept in the path of radiation stops these waves from reaching us.
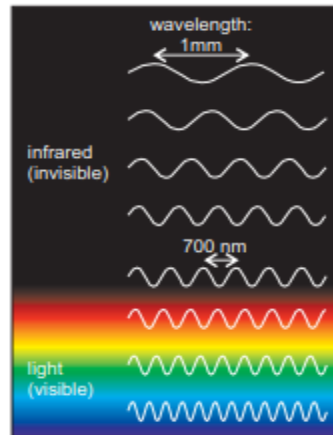
Figure 9.5: Thermal radiations and visible light spectrum.

Figure 9.6: Heat from the fireplace reaches us by radiation.

Radiations are emitted by all bodies. The rate at which radiations are emitted depends upon various factors such as

• Colour and texture of the surface

• Surface temperature

• Surface area

Why does a cup of hot tea become cold after some time? Why does a glass of chilled water become hot after some time?

All the objects lying inside a room including the walls, roof, and floor of the room are radiating heat. However, they are also absorbing heat at the same time. When the temperature of an object is higher than its surroundings then it is radiating more heat than it is absorbing.

As a result, its temperature goes on decreasing till it becomes equal to its surroundings. At this stage, the body is giving out the amount of heat equal to the amount of heat it is absorbing. When the temperature of an object is lower than its surroundings, then it is radiating less heat than it is absorbing.

Fig. 9.7

Radiations from
Leslie's cube.
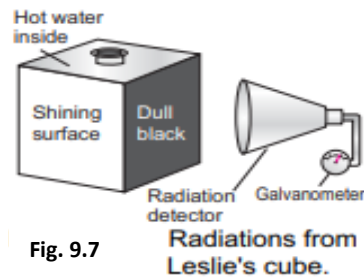
As a result, its temperature goes on increasing till it becomes equal to its surroundings. The rate at which various surfaces emit heat depends upon the nature of the surface. Various surfaces can be compared using Leslie's cube.

Emission and Absorption of Radiation

A Leslie cube is a metal box having faces of different natures as shown in Figure 9.7. The four faces of Leslie's cube may be as follows:

• A shining silvered surface

• A dull black surface

• A white surface

•A coloured surface

Fig.9.8   A comparison of
absorption of radiation.
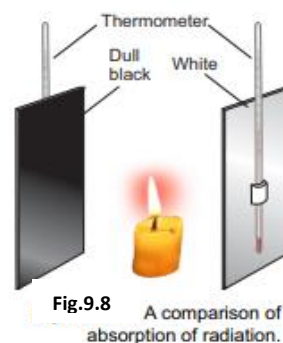
Hot water is filled in Leslie's cube and is placed with one of its faces towards a radiation detector. It is found that a black dull surface is a good emitter of heat. The rate at which various surfaces absorb heat also depends upon the nature of those surfaces. For example, take two surfaces, one is dull black, and the other is a silver polished surface as shown in Figure 9.16 with a candle in the middle of the surface. It is found that A dull black surface is a good absorber of heat as its temperature rises rapidly. A polished surface is a poor absorber of heat as its temperature rises very slowly. The observations made from the setup shown in Figure 9.16 are shown in the table given below,

Surfaces	Emitter	Absorber	Reflector
dull black surface	best	best	worst
coloured surface	good	good	bad
White surface	bad	bad	good
shining silvered surface	worst	worst	best

It is also found that the transfer of heat by radiation is also affected by the surface area of the body emitting or absorbing heat. The larger the area, the greater will be the transfer of heat. It is due to this reason that large numbers of slots are made in radiators to increase their surface area.

9.1.4 GREENHOUSE EFFECT

How can the temperature in a greenhouse be maintained? Light from the Sun contains thermal radiation (infrared) of long wavelengths as well as light and ultraviolet radiation of short wavelengths. Glass and transparent polythene sheets allow radiation of short wavelengths to pass through easily but not long wavelengths of thermal radiation. Thus, a greenhouse becomes a heat trap. Radiations from the Sun pass easily through the glass and warm up the objects in a greenhouse. These objects and plants as shown in figure 9.9, give out radiation of much longer wavelengths. Glass, and transparent polythene sheets.

Fig. 9.9: A Green House

do not allow them to escape easily and are reflected in the greenhouse. This maintains the inside temperature of the greenhouse. The greenhouse effect promises better growth of some plants. Carbon dioxide and water also behave similarly to radiations as glass or

polythene. Earth's atmosphere contains carbon dioxide and water vapors. It causes the greenhouse effect as shown in Figure 9.10 and thus maintains the temperature of the Earth. In recent years, the percentage of carbon dioxide has increased considerably. This has caused an increase in the average temperature of the Earth by trapping more heat due to the greenhouse effect. This phenomenon is known as global warming. This has serious implications for the global climate.
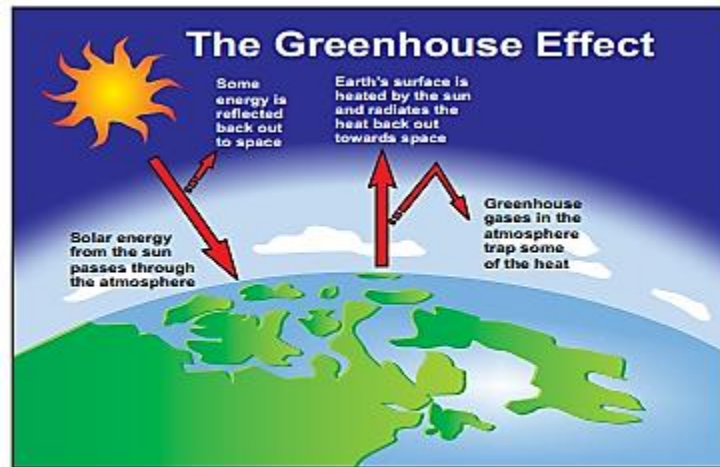
Figure 9.10: Greenhouse effect in global warming

APPLICATION AND CONSEQUENCES OF RADIATION

Different objects absorb different amounts of heat radiation falling upon them reflecting the remaining part. The amount of heat absorbed by a body depends upon the color and nature of its surface. A black and rough surface absorbs more heat than a white or polished surface. Good absorbers are also good radiators of heat. Thus, a black-coloured body gets hot quickly absorbing heat reaching it during a sunny day, and cools down quickly by giving out its heat to its surroundings. The bottoms of cooking pots are made black to increase the absorption of heat from fire. Like light rays, heat radiation also obeys laws of reflection. The amount of heat reflected from an object depends upon its color and the nature of the surface. White surfaces reflect more than colored or black surfaces. Similarly, polished surfaces are better reflectors than rough surfaces and the reflection of heat radiation is greater from polished surfaces. Hence, we wear white or light-coloured clothes in summer which reflect most of the heat radiation reaching us during the hot day. We polish the interior of the cooking and hot pots to reflect most of the heat radiation within them.

EXERCISE

1. Why metals are good conductors of heat?
2. Explain why:
a) Metal feels colder to touch than wood kept in a cold place.
b) The land breeze blows from the land towards the sea.
c) Is a double-walled glass vessel used in the thermos flask?
d) deserts soon get hot during the day and soon get cold after sunset.
3. Why does the conduction of heat not take place in gases?
4. What measures do you suggest conserving energy in houses?
5. Why does the transfer of heat in fluids take place by convection?
6. What is meant by convection current?
7. Suggest a simple activity to show the convection of heat in gases not given in the book.
8. How does heat reach us from the Sun?
9. How various surfaces can be compared by a Leslie cube?
10. What is the greenhouse effect?
11. Explain the impact of the greenhouse effect on global warming.

Chapter # 10

Gas Laws

Students will be able to understand:

- Boyle's law
- Charles laws
- Applications of gas laws
- Effects of changes in atmospheric pressure on physiology of the human body.

10.1 STATES OF MATTER

Matter exists in four states i.e., solid, liquid, gas, and plasma. The simplest form of matter is the gaseous state and most of matter around us is in the solid state. Liquids are less common than solids, gases, and plasmas. The reason is that the liquid state of any substance can exist only within a relatively narrow range of temperature and pressure. Let us look at the general properties of gases, liquids, and solids. Kinetic molecular theory of gases can help us understand their properties.

10.1.1 Properties of Gases

1. Gases don't have a definite volume and occupy all the available space. The volume of gas is the volume of the container.

2. They don't have a definite shape and take the shape of the container just like liquids.
3. Due to the low densities of gases, as compared to those of liquids and solids, the gases bubble through liquids and tend to rise.

4. Gases can diffuse and effuse. This property is negligible in solids but operates in liquids as well.

5. Gases can be compressed by applying pressure because there are large empty spaces between their molecules.

6. Gases can expand on heating or by increasing the available volume. Liquids and solids, on the other hand, do not show an appreciable increase in volume when they are heated.

7. When sudden expansion of gases occurs cooling takes place, it is called Joule Thomson effect.

8. Molecules of gases are in a constant state of random motion They can exert a certain pressure on the walls of the container and this pressure is due to the number of collisions.

9. The intermolecular forces in gases are very weak.

10.1.2 Properties of Liquids:

1. Liquids don't have a definite shape but have a definite volume. Unlike solids, they adopt the shape of the container.

2. Molecules of liquids are in a constant state of motion. The evaporation and diffusion of liquid molecules is due to this motion.

3. The densities of liquids are much greater than those of gases but are close to those of solids.

4. The spaces among the molecules of liquids are negligible just like solids.

5. The intermolecular attractive forces in liquids are intermediate between gases and solids. The melting and boiling points of gases, liquids and solids depend upon the strength of such forces.

6. Molecules of liquids possess kinetic energy due to their motion. Liquids can be converted into solids on cooling i.e., by decreasing their kinetic energy. Molecules of liquids collide among themselves and exchange energy but those of solids cannot do so.

10.1.3 Properties of Solids:

1. The particles present in solid substances are very close to each other and they are tightly packed. Due to this reason solids are non-compressible, and they cannot diffuse into each other.

2. There are strong attractive forces in solids which hold the particles together firmly and for this reason, solids have definite shape and volume.

3. The solid particles possess only vibrational motion.

10.2 GAS LAWS:

It is a matter of common observation that when external conditions of temperature and pressure are changed, the volume of a given quantity of all gases is affected. This effect is nearly the same irrespective of the nature of the gas. So, gases show a uniform behavior towards external conditions. The gas laws describe this uniform behavior of gases. The relationships between volume of a given amount of gas and the prevailing conditions of temperature and pressure are called the gas laws. Different scientists, like Boyle, Charles, Graham and Dalton have given their laws relating to the properties of gases.

The gas laws are a group of laws that govern the behavior of gases by providing relationships between the following:

i. The volume occupied by the gas.
ii. The pressure exerted by a gas on the walls of its container.
iii. The absolute temperature of the gas.
iv. The amount of gaseous substance (or) the number of moles of gas.

The gas laws were developed towards the end of the 18th century by numerous scientists (after whom the individual laws are named). The five gas laws are listed below:

Boyle's Law: It provides a relationship between the pressure and the volume of a gas.

Charles's Law: It provides a relationship between the volume occupied by a gas and the absolute temperature.

Gay-Lussac's Law: It provides a relationship between the pressure exerted by a gas on the walls of its container and the absolute temperature associated with the gas.

Avogadro's Law: It provides a relationship between the volume occupied by a gas and the amount of gaseous substance.

The Combined Gas Law (or the Ideal Gas Law): It can be obtained by combining the four laws listed above.

10.2.1 Boyle's Law:

Boyle's law gives the relationship between the pressure of a gas and the volume of the gas at a constant temperature. The volume of a gas is inversely proportional to the pressure of a gas at a constant temperature.
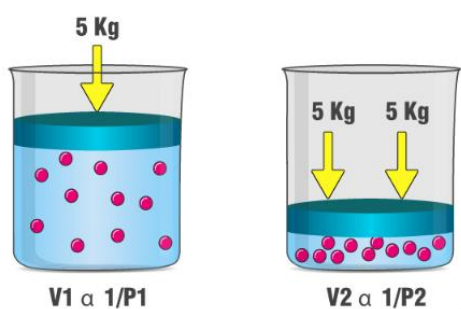
Boyle's law equation is written as:

$$V \propto 1/P$$

Or

$$P \propto 1/V$$

Or

$$PV = k$$

Where V is the volume of the gas, P is the pressure of the gas, and K_1 is the constant. Boyle's Law can be used to determine the current pressure or volume of gas and can also be represented as,

$$P_1V_1 = P_2V_2$$

Experimental Verification of Boyle's Law:

The following diagram indicates that at constant temperature say at 250C, the volume of a given quantity of a gas is reduced in proportion to the increase in pressure. Let us take a gas in a cylinder having a moveable piston.
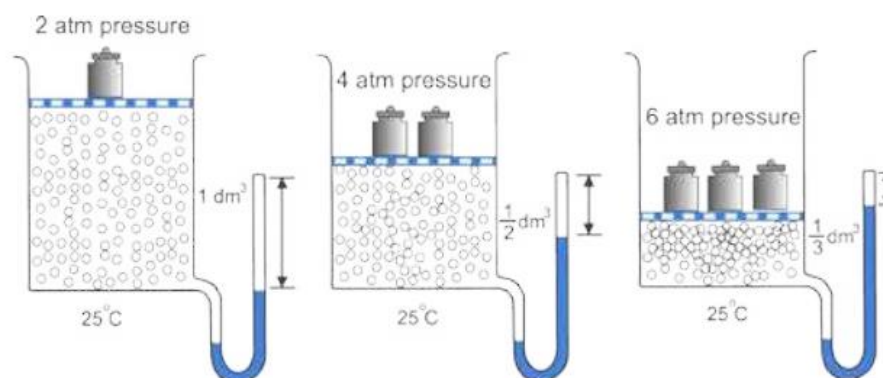
The cylinder is also attached with a manometer to read the pressure of the gas directly. Let the initial volume of gas is 1 dm^3 and its pressure is 2 atmospheres when the piston has one weight on it. When the piston is pressed twice with the help of two equal weights, the pressure becomes four atmospheres. Similarly, when the piston is loaded with a mass three times greater, then the pressure becomes six atmospheres.

The initial volume of the gas at two atmospheres is 1 dm^3 it is reduced to 1/2 dm^3 and then 1/3 dm^3 with an increase of weights, respectively.

$$P_1 V_1 = \text{2 atm x 1 d } m^3 = 2 \text{ d } m^3 \text{ atm} = k$$

$$P_2 V_2 = \text{4 atm x 1/2 d } m^3 = 2 \text{ d } m^3 \text{ atm} = k$$

$$P_3 V_3 = \text{6 atm x 1/3 d } m^3 = 2 \text{ d } m^3 \text{ atm} = k$$

Hence Boyle's law is verified. The value of k will remain the same for the same quantity of gas at the same temperature.

10.2.2 **Charle's Law:**

Charle's law states that at constant pressure, the volume of a gas is directly proportional to the temperature (in Kelvin) in a closed system. Basically, this law describes the relationship between the temperature and volume of the gas.
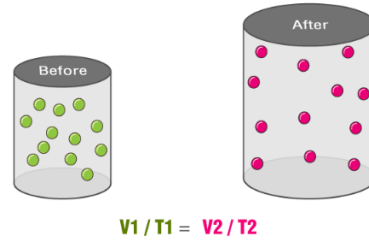
$V_1 / T_1 = V_2 / T_2$

Mathematically, Charle's law can be expressed as,

$$V \propto T$$

Where,

V = volume of gas

T = temperature of the gas in Kelvin.

Another form of this equation can be written as,

$$V_1 / T_1 = V_2 / T_2$$

Experimental Verification of Charles 's Law:

Let us consider a certain amount of gas enclosed in a cylinder fitted with a movable piston. The volume of the gas is V1 and its temperature is T1. When the gas in the cylinder is heated both volume and the temperature of the gas increase.
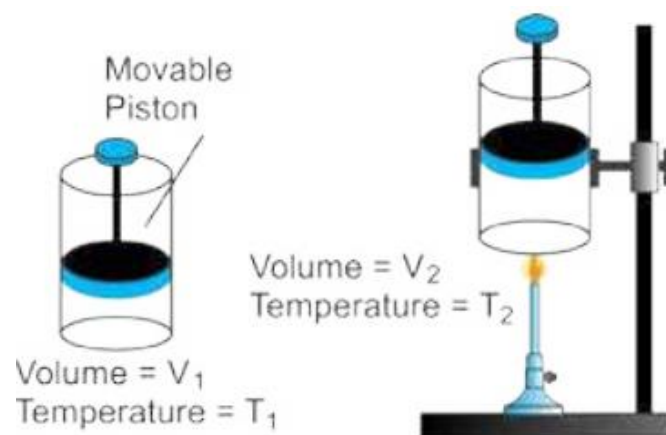
The new values of volume and temperature are V_2 and T_2 respectively. Fig. shows that,

$$V_1 / T_1 = V_2 / T_2$$

Hence Charles's law is verified.

10.2.3 DALTON'S LAW OF PARTIAL PRESSURES:

John Dalton studied the mixtures of gases and gave his law of partial pressures. According to this law, the total pressure exerted by a mixture of non-reacting gases is equal to the sum of their individual partial pressures. Let the gases are designated as 1,2,3, and their partial pressures are p_1, p_2, p_3. The total pressure (P) of the mixture of gases is given by

$$P_t = p_1 + p_2 + p_3$$

The partial pressure of a gas in a mixture of gases is the pressure that it would exert on the walls of the container if it were present all alone in that same volume under the same temperature. Let us have four cylinders of the same volume, i.e., 10 dm3 each and three gases H_2, CH_4, and O_2 are separately enclosed in the first three of them at the same temperature. Let their partial pressures be 400 torr, 500 torr and 100 torr respectively.

All these gases are transferred to a fourth cylinder of capacity 10 dm3 at the same temperature. According to Dalton's law

$$P_t = p_{H_2} + p_{CH_4} + p_{O_2} = (400 + 500 + 100) \text{ torr}$$
$$P_t = 1000 \text{ torr}$$

These three non-reacting gases are behaving independently under normal conditions. The rapidly moving molecules of each gas in a mixture have equal opportunities to collide with the walls of the container. Hence, each gas exerts a pressure independent of the pressure of other gases. The total pressure is the result of the total number of collisions per unit area in each time. Molecules of each gas move independently, so the general gas equation (PV = nRT) can be applied to the individual gases in the gaseous mixture.

$$p_{H_2}V = n_{H_2}RT \qquad p_{H_2} = n_{H_2}\frac{RT}{V} \qquad p_{H_2} \, \alpha \, n_{H_2}$$

$$p_{CH_4}V = n_{CH4}RT \qquad p_{CH_4} = n_{CH_4}\frac{RT}{V} \qquad p_{CH_2} \, \alpha \, n_{CH_4}$$

$$p_{O_2}V = n_{O_2}RT \qquad p_{O_2} = n_{O_2}\frac{RT}{V} \qquad p_{O_2} \, \alpha \, n_{O_2}$$

$$\frac{RT}{V} \text{ is a constant factor for each gas.}$$

All these gases have their partial pressures. Since volumes and temperatures are the same, so their number of moles will be different and will be directly proportional to their partial pressures. Adding these three equations

$$P_t = p_{H_2} + p_{CH_4} + p_{O_2}$$

$$P_t = (n_{H_2} + n_{CH_4} + n_{O_2}) \frac{RT}{V}$$

$$P_t = n_t \frac{RT}{V} \qquad \text{where } n_t = n_{H_2} + n_{CH_4} + n_{O_2}$$

$$P_t V = n_t RT$$

According to equation, the total pressure of the mixture of gases depends upon the total number of moles of the gases.

Calculation of Partial Pressure of a Gas The partial pressure of any gas in a mixture of gases can be calculated, provided one knows the mass of that gas or its number of moles along with the total pressure and the total number of moles present in the mixture. To have a relationship, let us suppose that we have a mixture of gas A and gas B. This mixture is enclosed in a container having volume (V). The total pressure is one atm. The number of moles of the gases A and B are n_A and n_B respectively. If they are maintained at temperature T, then,

$$P_t V = n_t RT \qquad .... \text{(equation for the mixture of gases)}$$
$$p_A V = n_A RT \qquad .... \text{(equation for gas A)}$$
$$p_B V = n_B RT \qquad .... \text{(equation for gas B)}$$

Divide the First two equations.

$$\frac{p_A V}{P_t V} = \frac{n_A RT}{n_t RT}$$

$$\frac{p_A}{P_t} = \frac{n_A}{n_t}$$

$$p_A = \frac{n_A}{n_t} P_t$$

$$p_A = x_A P_t$$

$$p_B = x_B P_t$$

Similarly Partial pressure of a gas is the mole fraction of that gas multiplied by the total pressure of the mixture. Remember that mole fraction of anyone of the gases in the mixture is less than unity. Moreover, the sum of mole fractions is always equal to unity.

Applications of Dalton's Law of Partial Pressures:

The following are the four important applications of Dalton's Law of partial pressures. 1. Some gases are collected over water in the laboratory. The gas during collection gathers water vapours and becomes moist. The pressure exerted by this moist gas is, therefore, the sum of the partial pressures of the dry gas and that of water vapours.

The partial pressure exerted by the water vapours is called aqueous tension.

$$P_{moist} = P_{dry} + P_{w.vap}$$
$$P_{moist} = P_{dry} + \text{aqueous tension}$$
$$P_{dry} = P_{moist} - \text{aqueous tension}$$

While solving the numerical the aqueous tension is subtracted from the total pressure (P moist).

2. Dalton's law finds its applications during the process of respiration. The process of respiration depends upon the difference in partial pressures. When animals inhale air oxygen moves into the lungs as the partial pressure of oxygen in the air is 159 torr, while the partial pressure of oxygen in the lungs 116 torr. CO_2 produced during respiration moves out in the opposite direction, as its partial pressure is higher in the lungs than that in the air.

3. At higher altitudes, the pilots feel uncomfortable breathing because the partial pressure of oxygen in the un-pressurized cabin is low, as compared to 159 torr, where one feels comfortable breathing.

4. Deep Sea divers take oxygen mixed with an inert gas say He and adjust the partial pressure of oxygen according to the requirement. In sea after every 100 feet depth, the diver experiences approximately 3 atm pressure, so normal air cannot be breathed in depth of sea. Moreover, the pressure of N_2 increases in depth of sea and it diffuses in the blood.

Effects of changes in atmospheric pressure on physiology of the human body:

Barometric pressure, also called atmospheric pressure or air pressure, is the force or weight of the air surrounding us. Barometric pressure is measured by an instrument called the barometer.

One of the most common types of barometers is a mercury barometer in which the height of a column of mercury that exactly balances the weight of the column of atmosphere over the barometer represents the barometric pressure at that point. At standard sea level, the barometric pressure equals 760 mm (29.92 inch) of mercury. A rise in barometric pressure is generally considered an improvement in the weather, while a fall in barometric pressure may mean worsening weather.

A fall in barometric pressure can affect health in various ways.

Mountain sickness or altitude sickness: It refers to a group of general symptoms occurring on climbing or walking to a higher altitude or elevation too quickly. At heights above 1,500-3,000 m (5,000-10,000 feet), the pressure is low enough to produce altitude sickness. This happens especially when the person ascends too rapidly, not allowing their bodies to adapt or acclimatize to the fall in pressure and oxygen levels with increasing height. The symptoms may include:

  i. Headache
 ii. Nausea
iii. Feeling uneasy or sick
 iv. Shortness of breath
  v. Tiredness or fatigue
 vi. Dizziness

If left untreated, the symptoms may progress to severe breathlessness, cough, vomiting, confusion, and unconsciousness. These symptoms occur because the fall in barometric pressure allows the tissues in the lungs and brain to swell (visualize the change in the size of a sponge when you squeeze and release it) in an attempt to get more oxygen. The dilated blood vessels in the brain may cause headaches and swelling of the brain. The swelling puts pressure on the brain, squeezing it against the skull.

Altitude sickness may rarely advance to a more severe form of the illness called high altitude cerebral edema (HACE). HACE occurs when brain swelling becomes severe, manifesting as severe headache, confusion, lethargy, lack of coordination, irritability, vomiting, seizures, coma, and eventually death if untreated. Severe altitude sickness may cause the capillaries (tiny blood vessels) in the lungs to become swollen and leaky. This leads to an accumulation of fluid in the air sacs of the lungs. This condition is known as high altitude pulmonary edema (HAPE). HAPE severely decreases the exchange of oxygen in the lungs, which may cause breathing difficulties and even death.

Worsening of arthritis: It refers to the inflammation of joints. This may be due to the changes in the joint fluid as the pressure changes. Arthritis manifests as painful, swollen joints that may be associated with joint stiffness. Elderly people often describe

the pain associated with a fall in barometric pressure as a storm coming in their knees (storms are associated with a sudden fall in barometric pressure). Some people may be more sensitive to weather changes experiencing more stiffness, pain, and swelling with a barometric pressure decline. Scientists suggest that a fall in air pressure allows the tissues (including muscles and tendons) to swell or expand. This exerts pressure on the joints resulting in increased pain and stiffness. A fall in air pressure may exert a greater effect if it is accompanied by a fall in temperature as well. A lower temperature makes the joint fluids thicker, eventually worsening the symptoms.

Headaches: Some people report worsening of headaches including those caused by sinusitis (sinus inflammation) and migraines when the barometric pressure falls. The skull has several air pockets called sinuses that keep the skull light. When air pressure drops, there is a difference in the pressure of the outside air and the air in your sinuses. This can cause a headache with a small change in pressure. This is the same reason why your ears "pop" when the airplane ascends causing a difference in the air pressure on either side of the eardrums. When the barometric pressure falls beyond a certain limit, there may be associated swelling of blood vessels and tissues in and around the brain causing worsening of headache. The pain is generally felt in one or both temples along with other symptoms such as nausea, vomiting, intolerance to light, and numbness in the face and neck.

EXERCISE

1. What is Boyle's law of gases? Give its experimental verification.
2. What are isotherms? What happens to the positions of isotherms when they are plotted at high temperature for a particular gas.
3. Why do we get a straight line when pressures exerted on a gas are plotted against inverse of volumes? This straight line changes its position in the graph by varying the temperature. Justify it.
4. How will you explain that the value of the constant k in the equation PV = k depends upon
   (i) the temperature of a gas
   (ii) the quantity of a gas
5. What is Charles's law? Which scale of temperature is used to verify that V/T = k (pressure and number of moles are constant)?
6. A sample of carbon monoxide gas occupies 150.0 mL at 25.0°C. It is then cooled at constant pressure until it occupies 100.0 mL . What is the new' temperature?
   (Ans: 198.8K or -74.4 °C).
7. Dalton's law of partial pressures is only obeyed by those gases which don't have attractive forces among their molecules. Explain it.
8. Derive an equation to find out the partial pressure of a gas knowing the individual moles of component gases and the total pressure of the mixture.
9. Explain that the process of respiration obeys Dalton's law of partial pressures.
10. Helium gas in a 100 cm^3 container at a pressure of 500 torr is transferred to a container with a volume of 250 cm^3 .
11. What will be the new pressure a. if no change in temperature occurs.
    (Ans: 2 0 0 torr)
12. If its temperature changes from 20 "C to 15°C? (Ans: 196.56 torr)
13. A sample of krypton with a volume of 6.25 dm^3 , a pressure of 765 torr and a temperature of 20 °C is expanded to a volume of 9.55 dm^3 and a pressure of 375 torr. What will be its inal temperature in °C?
    (Ans: T = -53.6°c)
14. Working at a vacuum line, a chemist isolated a gas in a weighing bulb with a volume of 255 cm^3 , at a temperature of 25 °C and under a pressure in the bulb of 10.0 torr. The gas weighed 12.1 mg. What is the molecular mass of this gas?
    (Ans: 87.93g mol^{-1})

Chapter # 11

Sound

Students will be able to understand:

- Sound
- Characteristics of sound.
- How sound is produced?
- Characteristics of sound
- Transmission and reflection of sound
- Echoes
- Ultrasound.

We know that vibrations of objects in any medium produce waves. For example, the vibrator of ripple tank produces water waves. The medium in this case is liquid, but it can also be a gas or a solid. Now we will discuss another type of waves that we can hear i.e., sound waves.

11.1 SOUND WAVES:

Like other waves, the sound is also produced by vibrating bodies. Due to vibrations of bodies, the air around them also vibrates and the air vibrations produce a sensation of sound in our ears. For example, in a guitar, the sound is produced due to the vibrations of its strings (Fig. 11.1). Our voice results from the vibrations of our vocal cords. The human heart beats and vibrations of other organs like lungs also produce sound waves. Doctors use stethoscopes to hear this sound.
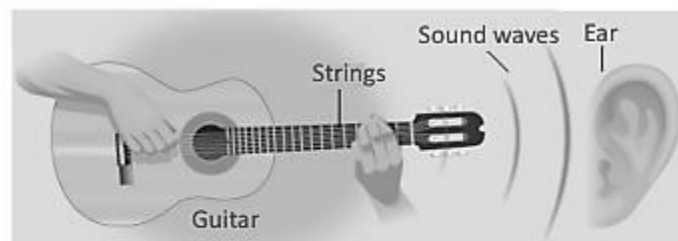
Fig. 11.1: Vibrations of guitar strings produce sound waves

11.1.1 SOUND IS PRODUCED BY A VIBRATING BODY
Experiment:

In school laboratories, we use a device called tuning fork to produce a particular sound. If we strike the tuning fork against rubber hammer, the tuning fork will begin to vibrate (Fig. 11.2). We can hear the sound produced by tuning fork by bringing it near our ear.

We can also feel the vibrations by slightly touching one of the prongs of the vibrating tuning fork with a plastic ball suspended from a thread (Fig. 11.3).

Touch the ball gently with the prong of a vibrating tuning fork. The tuning fork will push the ball because of its vibrations.

Now if we dip the vibrating tuning fork into a glass of water, we will see a splash (Fig. 11.4). What does make the water splash? From this activity, we can conclude that sound is produced by vibrating bodies. Sound Requires Material Medium for its Propagation.
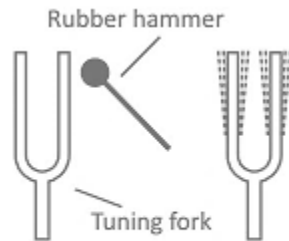
Fig. 11.2: Strike a rubber hammer on a tuning fork
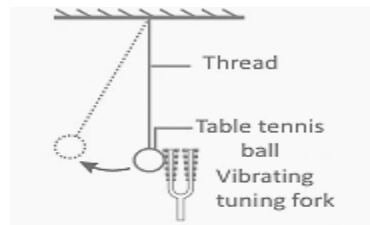
Fig. 11.3: The production of sound waves from a vibrating tuning fork

An electric bell is suspended in the bell jar with the help of two wires connected to a power supply. By setting ON the power supply, the electric bell will begin to ring. We can hear the sound of the bell. Now start pumping out air from the jar using a vacuum pump. The sound of the bell starts becoming more and more feeble and eventually dies out, although the bell is still ringing. When we put the air back into the jar, we can hear the sound of the bell again. From this activity, we conclude that sound waves can only travel/propagate in the presence of air (medium).
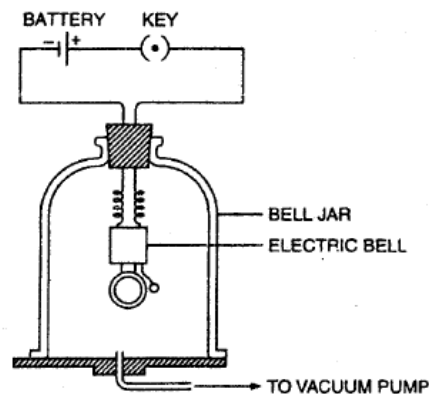
Fig. 11.4

11.2 Longitudinal Nature of Sound Waves

Propagation of sound waves produced by vibrating tuning fork can be understood by a vibrating tuning fork as shown in Fig.11.5. Before the vibration of tuning fork, density of air molecules on the right side is uniform (Fig.11.5-a). When the right prong of tuning fork moves from mean position O to B (Fig.11.5-b), it exerts some pressure on the adjacent layer of air molecules and produces a compression.
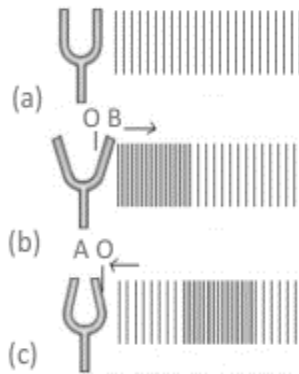
Fig.11.5: Vibrations of tuning fork after striking with a rubber.

This compressed air layer in turn compresses the layer next to it and so on. A moment later, the prong begins to move from B towards A (Fig.11.5-c). Now the pressure in the adjacent layer decreases and a rarefaction is produced. This rarefaction is transferred to the air layer next to it and so on. As the tuning fork moves back and forth rapidly, a series of compressions and rarefactions are created in the air. In this way, sound waves propagate through the air. As in the direction of propagation of sound wave Fig.11.5, is along the direction of oscillating air molecules. This shows the longitudinal nature of sound waves. The distance between two consecutive compressions or rarefactions is the wavelength of the sound wave.

11.3 CHARACTERISTICS OF SOUND

Sounds of different objects can be distinguished on the basis of different characteristics as described below:

11.3.1   Loudness:

Loudness is the characteristic of sound by which loud and faint sounds can be distinguished. When we talk to our friends, our voice is low, but when we address a public gathering our voice is loud. Loudness of a sound depends upon a**mplitude of the vibrating body**.
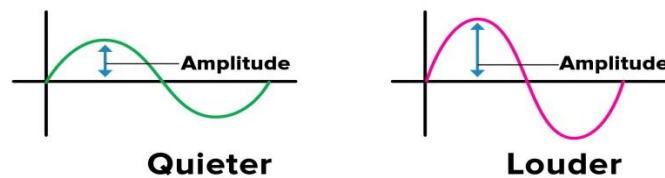
Fig, 11.6

The loudness of the sound varies directly with the amplitude of the vibrating body (Fig.11.6). The sound produced by a sitar will be loud if we pluck its wires more violently. Similarly, when we beat a drum forcefully, the amplitude of its membrane increases, and we hear a loud sound.

Area of the vibrating body:

The loudness of sound also depends upon the area of the vibrating body. n number of factors. Some of them are discussed below:

For example, sound produced by a large drum is louder than that by small one because of its large vibrating area. If we strike a tuning fork on a rubber pad, a feeble sound will be heard. But if the vibrating tuning fork is placed vertically on the surface of a bench, we will hear a louder sound. From this, we can conclude that the loudness increases with the area of the vibrating body and vice versa.
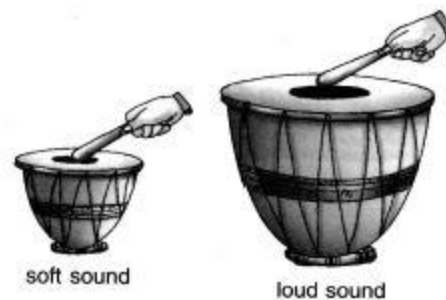
Fig.11.7

11.3.2 Distance from the vibrating body:
Loudness of sound also depends upon the distance of the vibrating body from the listener. It is caused by the decrease in amplitude due to an increase in distance. Loudness also depends upon the physical condition of the ears of the listener. A sound appears louder to a person with sensitive ears than to a person with defective ears. However, there is a characteristic of sound which does not depend upon the sensitivity of the ear of the listener, and it is called intensity of sound.

Pitch: Pitch is the characteristic of sound by which we can distinguish between a shrill and a grave sound. It depends upon the frequency. A higher pitch means a higher frequency and vice versa. The frequency of the voice of ladies and children is higher than that of men. Therefore, the voice of ladies and children is shrill and of high pitch. The relationship between frequency and pitch is illustrated in Fig. 11.8.
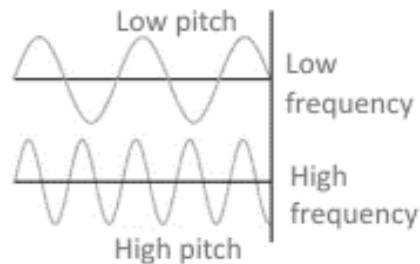
Fig 11.8: Variation of pitch with frequency

Quality: The characteristic of sound by which we can distinguish between two sounds of same loudness and pitch is called quality. While standing outside a room, we can distinguish between the notes of a piano and a flute being played inside the room. This is due to the difference in the quality of these notes. Figure 11.9 shows the waveform of the sound produced by a tuning fork, flute, and clarinet. The loudness and the pitch of these three sounds are the same but their waveforms are different. So, their quality is different, and they can be distinguished from each other. Intensity The sound waves transfer energy from the sounding body to the listener. The intensity of sound depends on the amplitude of sound wave and is defined as: Sound energy passing per second through a unit area held perpendicular to the direction of propagation of sound waves is called intensity of sound. Intensity is a physical quantity and can be measured accurately. The unit of intensity of sound is watt per square metre (W).
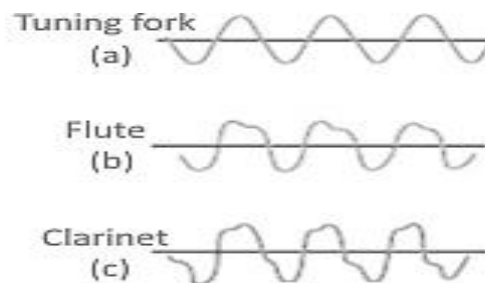
Fig 11.9: Sound waveforms produced by (a) a tuning fork, (b) a flute, and (c) a clarinet, are all at approximately the same frequency. Pressure is plotted vertically, time

Sound Intensity Level:

The human ear responds to intensities ranging from 10^{-12} W^{-2} m to more than 1 W m^{-2} (which is loud enough to be painful). Because the range is so wide, intensities are scaled by factors of ten. The barely audible and the faintest intensity of sound i.e., 10^{-12} Wm^{-2} is taken as reference intensity, called zero bel (a unit named after Alexander Graham Bell).

The loudness of a sound depends not only on the intensity of the sound but also on the physical conditions of the ear. The human ear is more sensitive to some frequencies than the others.

The loudness (L) of a sound is directly proportional to the logarithm of intensity i.e.,

L log I L = K log I .... (11.1)

where K is a constant of proportionality.

Let L_0 be the loudness of the faintest audible sound of intensity and L be the loudness of an unknown sound of intensity.

$$L_o = K \log I_o \qquad \text{.......... (11.2)}$$

Subtracting Eq. (11.2) from Eq. (11.1), we get

$$L - L_o = K (\log I - \log I_o) = K \log \frac{I}{I}$$

This difference, (L- L$_o$), between the loudness L of an unknown o sound and the loudness L is called the intensity level of the o unknown sound. Therefore, the intensity level of an unknown sound is given by:

$$\text{Intensity level} = K \log \frac{I}{I_o} \qquad \text{..........} \qquad (11.3)$$

The value of K depends not only on the units of and but I Io also on the unit of intensity level. If the intensity of any I unknown sound is 10 times greater than the intensity of Io the faintest audible sound i.e., =10 Io and the intensity level I Io of such a sound is taken as unit, called bel, the value of K becomes 1. Therefore, using K =1, Eq. (11.3) becomes

$$\text{Intensity level} = \log \frac{I}{I} \quad \text{(bel)} \qquad \text{..........} \qquad (11.4)$$

bel is a very large unit of intensity level of a sound. Generally, a smaller unit called decibel is used. Decibel is abbreviated as (dB). It must be remembered that 1 bel is equal to 10 dB. If the intensity level is measured in decibels, Eq. (11.4) becomes,

$$\text{Intensity level} = 10 \log \frac{I}{I_\circ} \text{ (dB)} \quad .... \quad (11.5)$$

Using Eq. (11.5), we can construct a scale for measuring the intensity level of sound. Such scale is known as "decibel scale".

IMPORTANT OF ACOUSTICS:

The technique or method used to absorb undesirable sounds by soft and porous surfaces is called acoustic protection.

Reflection of sound is more prominent if the surface is rigid and smooth, and less if the surface is soft and irregular. Soft, porous materials, such as draperies and rugs absorb large amount of sound energy and thus quiet echoes and softening noises. Thus by using such material in noisy places we can reduce the level of noise pollution.

Fig. 11.10 acoustic protection

However, if the surface of classrooms or public halls is too absorbent, the sound level may be low for the audience. Sometimes, when sound reflects from the walls, ceiling, and floor of a room, the reflecting surfaces are too reflective, and the sound becomes garbled. This is due to multiple reflections called reverberations. In the design of lecture halls, auditorium, or theater halls, a balance must be achieved between reverberation and absorption. It is often advantageous to place reflective surfaces behind the stage to direct sound to the audience.

Generally, the ceilings of lecture halls, conference halls and theatre halls are curved so that sound after reflection may reach all the corners of the hall (Fig 11.11). Sometimes curved sound boards are placed behind the stage so that sound after reflection distributed evenly across the hall (Fig. 11.12).
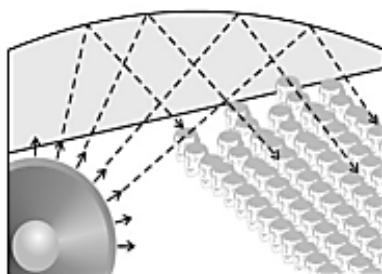
Fig. 11.11: Curved ceiling of a conference hall   Fig. 11.12: Soundboard used in a big hall
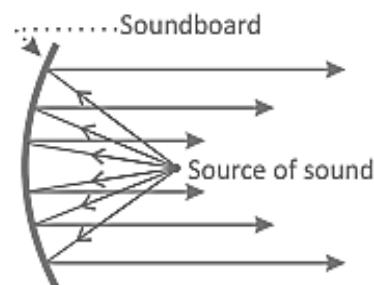
11.6 REFLECTION (ECHO) OF SOUND:

When we clap or shout near a reflecting surface such as a tall building or a mountain, we hear the same sound again a little later. What causes this? This sound that we hear is called an echo and is a result of the reflection of sound from the surface. When sound is incident on the surface of a medium it bounces back into the first medium. This phenomenon is called echo or reflection of sound. The sensation of sound persists in our brain for about 0.1 s. To hear a clear echo, the time interval between our sound and the reflected sound must be at least 0.1 s. If we consider -1 speed of sound to be 340 ms at a normal temperature in air, we will hear the echo after 0.1 s. The total distance covered by the sound from the point of generation to the reflecting -1 surface and back should be at least 340 m s × 0.1 s = 34.0 m. Thus, for hearing distinct echoes, the minimum distance of the obstacle from the source of sound must be half of this distance, i.e., 17 m. Echoes may be heard more than once due to successive or multiple reflections.

11.7 ULTRASOUND:

Sounds of frequency higher than 20, 000 Hz which are inaudible to normal human ear are called ultrasound or ultrasonics.
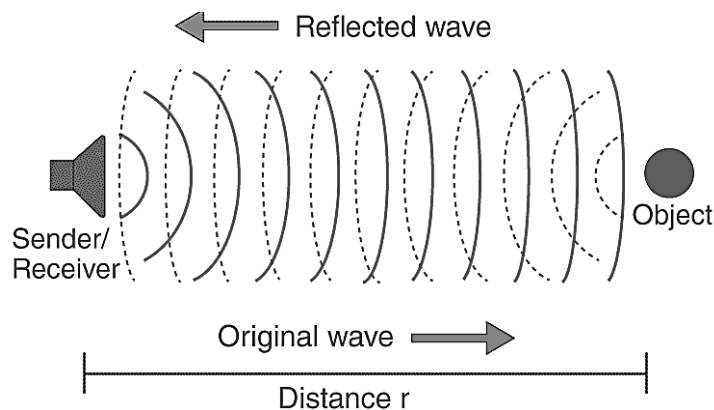
Fig. 11.13: working of ultra sound waves.

Uses of Ultrasound

• Ultrasonic waves carry more energy and higher frequency than audible sound waves. Therefore, according to the wave equation v = f λ, the wavelength of ultrasonic waves is very small and is very useful for detecting very small objects.

• Ultrasonics are utilized in medical and technical fields.

• In medical field, ultrasonic waves are used to diagnose and treat different ailments. For diagnosis of different diseases, ultrasonic waves are made to enter the human body through transmitters. These waves are reflected differently by different organs, tissues, or tumors etc. The reflected waves are then amplified to form an image of the internal

organs of the body on the screen (Fig.11.13). Such an image h e l p s i n detecting the defects in these organs.

• Powerful ultrasound is now being used to remove blood clots formed in the arteries.

• Ultrasound can also be used to get the pictures of thyroid gland for diagnosis purposes.
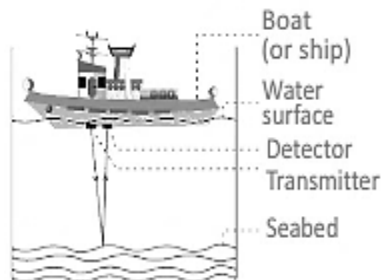
Fig. 11.14: Ultrasonics are used to measure the depth of water by echo method

• Ultrasound is used to locate underwater depths or is used for locating objects lying deep on the ocean floor, etc. The technique is called SONAR, (sound navigation and ranging). The sound waves are sent from a transmitter, and a receiver collects the reflected sound (Fig.11.14). The time-lapse is calculated, knowing the speed of sound in water, the distance of the object from the ocean surface can be estimated.

EXERCISE

1. What is the necessary condition to produce sound?
2. What is the effect of the medium on the speed of sound? In which medium does sound travels faster: air, solid or liquid? Justify your answer.
3. How can you prove the mechanical nature of sound by a simple experiment?
4. What do you understand by the longitudinal wave? Describe the longitudinal nature of sound waves. Sound is a form of wave. List at least three reasons to support the idea that sound is a wave.
5. What is the difference between the loudness and intensity of sound? Derive the relationship between the two. On what factors does the loudness of sound depend?
6. What do you mean by the term intensity level of the sound? Name and define the unit of intensity level of sound.
7. What are the units of loudness? Why do we use a logarithmic scale to describe the range of the sound intensities we hear?
8. Describe the effect of change in amplitude on loudness and the effect of change in frequency on pitch of sound.
9. If the pitch of sound is increased, what are the changes in the following? a. the frequency b. the wavelength c. the wave velocity d. the amplitude of the wave.
10. If we clap or speak in front of a building while standing at a particular distance, we rehear our sound after some time. Can you explain how does this happens?
11. What is the audible frequency range for the human ear? Does this range vary with the age of people?
12. Describe the importance of acoustic protection.
13. What are the uses of ultrasound in medicine?

NUMERICAL PROBLEMS

1. A normal conversation involves sound intensities of about 3.0×10^{-6} W m^{-2} . What is the decibel level for this intensity? What is the intensity of the sound for 100 dB?

   Ans.(64.8 dB, 0.01 W m^{-2})

2. If at Anarkali Bazar Lahore, intensity level of sound is 80 dB, what will be the intensity of sound there?

   Ans. (10^{-4} W m^{-2})

3. At a particular temperature, the speed of sound in air is 330 m/s . If the wavelength of a note is 5 cm, calculate the frequency of the sound wave. Is this frequency in the audible range of the human ear?

   Ans. (6.6 x 10 Hz, Yes)

4. A doctor counts 72 heartbeats in 1 min. Calculate the frequency and period of the heartbeats.

Ans. (1.2 Hz, 0.83 s)

5. A marine survey ship sends a sound wave straight to the seabed. It receives an echo 1.5 s later. The speed of sound in seawater is 1500 m/s . Find the depth of the sea at this position.

Ans. (1125 m)

6. A student clapped his hands near a cliff and heard the echo after 5 s. What is the distance of the cliff from the student if the speed of the sound is taken as 346 m/s ?

Ans. (865 m)

7. A ship sends out ultrasound that returns from the seabed and is detected after 3.42 s. If the speed of ultrasound through seawater is 1531 m/s, what is the distance of the seabed from the ship?

Ans. (2618 m)

8. The highest frequency sound humans can hear is about 20,000 Hz. What is the wavelength of sound in air at this frequency at a temperature of 20°C? What is the wavelength of the lowest sounds we can hear of about 20 Hz? Assume the speed of sound in air at 20 C is 343 m/s.

Chapter # 12

Light

Students will be able to understand the concept about:

- Transmission
- Reflection
- Refraction of light
- Lenses and their types.

Nothing can travel faster than the speed of light in space.it travels straight with the speed of 300,000 km/s. it is denoted by "c". The light passing through material such as air, water, or glass, however, travels relatively slowly. It is because of light scattered and is slowed down by the atoms in the material as it passes through them.

Light is the main focus of this unit. We shall describe different phenomena of light such as reflection, refraction and total internal reflection. We will learn how images are formed by mirrors and lenses and will discuss the working principle of compound microscope and telescope.

12.1 Transmission:

Light does not need a material medium to travel. It can travel through a vacuum too. When light emerging from a source hit an object, it can be transmitted, absorbed, reflected, refracted, or dispersed.

When light falls on transparent objects such as air, water, glass, etc, it passes through them unchanged (a) It is called transmission of light.

Translucent objects such as frosted glass, thin curtains, tracing paper etc allow only some part of light to pass through them. That is why we do not see clearly across them.

When light falls on an opaque object, most of the light is absorbed (fig. b). That is why we cannot see through them. When light is absorbed, heat is generated. Black
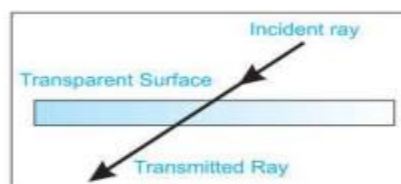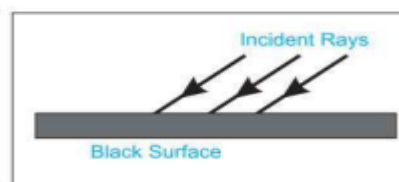
Fig (a)   Fig (b)
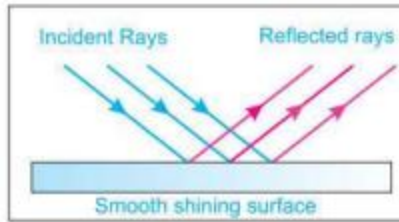
Fig (c)

substances absorb most of the light. What do you see when you look at the mirror? You see your image. It is due to reflection. The bouncing back of light from a smooth shiny Surface is called reflection (c).

12.2 REFLECTION OF LIGHT:

Reflection of light is illustrated in Fig. 12.1. When a ray of light from air along the path AO falls on a plane mirror M, it is reflected along the path OB. The ray AO is called incident ray while the ray OB is called reflected ray. The angle between incident ray AO and normal N, i.e., AON is called the angle < of incidence represented by i. The angle between the normal and the reflected ray OB, i.e., NOB is called angle of < reflection represented by r.
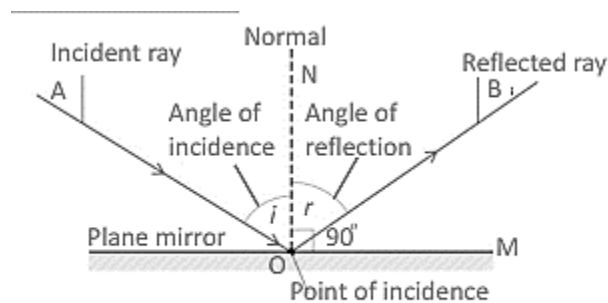
Fig. 12.1: Reflection of light

Now we can define the phenomenon of reflection as:

When light travelling in a certain medium fall on the surface of another medium, a part of it turns back in the same medium.

12.2.1 Laws of Reflection

The incident ray, the normal, and the reflected ray at the point of incidence all lie in the same plane.

The angle of incidence is equal to the angle of reflection i.e., i = r.

12.2.2 Types of Reflection

Nature of reflection depends on smoothness of the surface. For example, a smooth surface of silver reflects rays of light in one direction only. The reflection by these smooth surfaces is called regular reflection (Fig.12.2). Most of the objects in everyday world are not smooth on the microscopic level. The rough surfaces of these objects reflect the rays of light in many directions. Such type of reflection is called irregular reflection (Fig. 12.3).
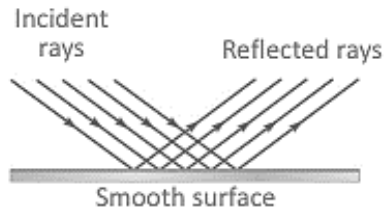
Fig. 12.2: Regular reflection Fig. 12.3: Irregular reflection
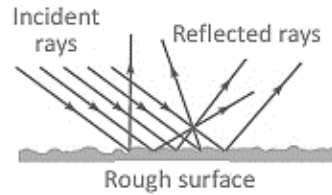
12.2.3 SPHERICAL MIRRORS

A mirror who's polished, reflecting surface is a part of a hollow sphere of glass or plastic is called a spherical mirror. In a spherical mirror, one of the two curved surfaces is coated with a thin layer of silver followed by a coating of red lead oxide paint. Thus, one side of the spherical mirror is opaque, and the other side is a highly polished reflecting surface. Depending upon the nature of reflecting surface, there are two types of spherical mirrors as shown in Fig.12.4.
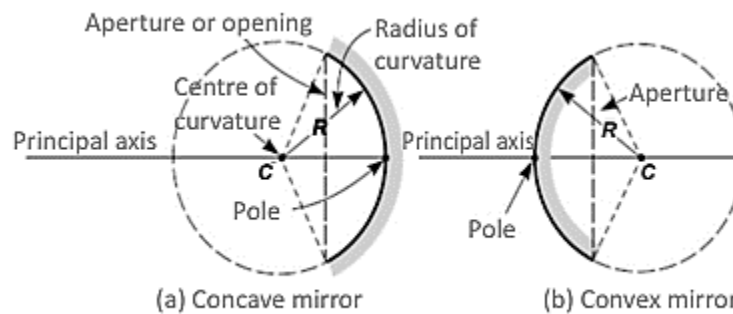
(a) Concave mirror   (b) Convex mirror
Fig. 12.4: Types of spherical mirrors

a) Concave Mirror:

A spherical mirror whose inner curved surface is reflecting is called concave mirror. In concave mirror the size of the image depends on the position of the object. Both virtual and real images can be formed by a concave mirror.

Convex Mirror: A spherical mirror whose outer curved surface is reflecting is called convex mirror. In convex mirror the size of the image is always smaller than the object. Only virtual and erect image is formed by a convex mirror.

Pole: It is the midpoint of the curved surface of spherical mirror. It is also called vertex.

Centre of Curvature (C): A spherical mirror is a part of a sphere. The centre of this sphere is called centre of curvature.

Radius of Curvature (R): It is the radius of the sphere of which the spherical mirror is a part. Principal Axis: It is the line joining centre of curvature and pole of the spherical mirror.

The Principle focus (F): After reflection from a concave mirror, rays of light parallel to the principal axis converge to point F. This point is called "The Principal Focus" of the mirror (Fig.12.5-a). Hence, Concave mirrors are also called converging mirrors. Since rays actually pass through this point, therefore, it is called real focus.

b) Convex mirror

In the case of a convex mirror, rays parallel to the principal axis after reflection appear to come from point F situated behind the mirror. In other words, rays of light appear to diverge from F. This point is called the principal focus of the convex mirror. Convex mirrors are also called diverging mirrors. The principal focus of a convex mirror is virtual because the reflected rays do not pass through it but appear to do so (Fig. 12.5-b).

Focal length (f): It is the distance from the pole to the principal focus measured along the principal axis (Fig12.5). The focal length is related to the radius of curvature by f =R/2. This means that as the radius of curvature is reduced, so too is the focal length of the reflecting surface.

Fig. 12.5
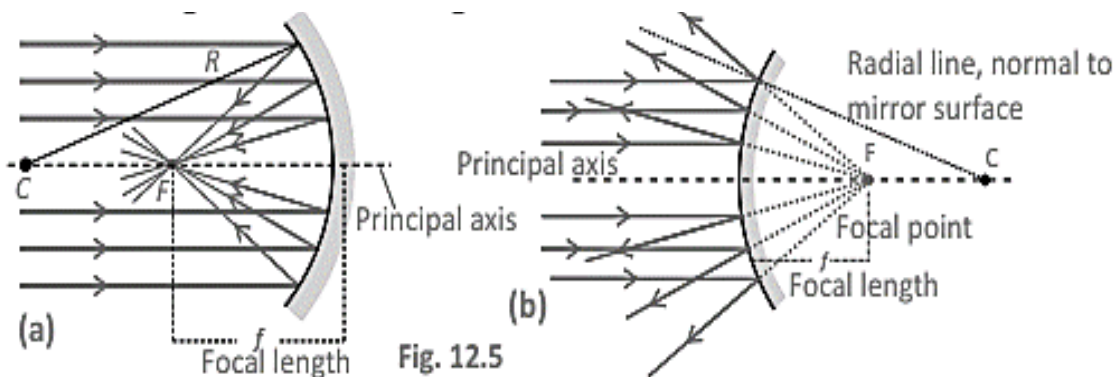
Characteristics of Focus of a Concave and a Convex Mirror

Convex Mirror	Concave Mirror
The Focus lies behind the mirror	The focus is in front of the mirror
The focus is virtual as the rays of light after reflection appear to come from the focus.	The focus is real as the rays of light after reflection converge at the focus.

Reflection of Light by Spherical Mirrors:

Like plane surfaces, spherical surfaces also reflect light following the two laws of reflection as stated for plane surfaces. Fig.12.6 shows how light is reflected by the spherical surfaces of concave and convex mirrors according to the two laws of reflection.

a) Concave mirror   b) Convex mirror
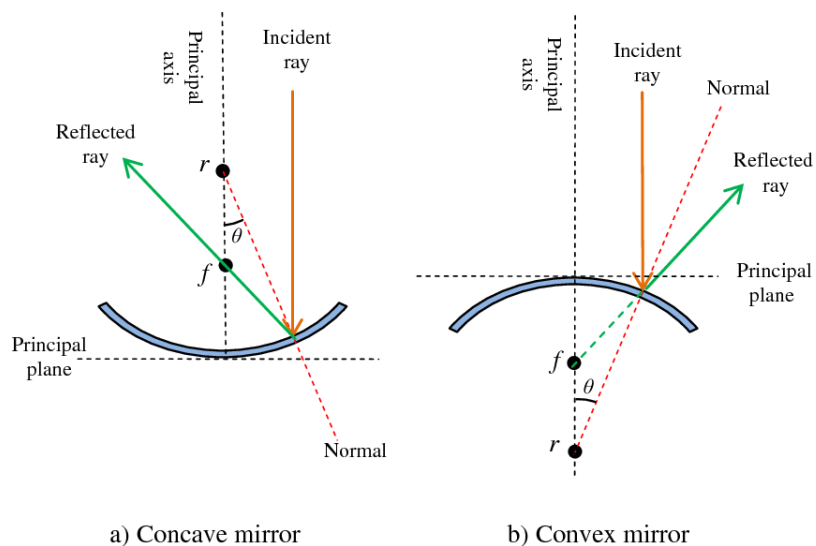
Fig 12.6: Light rays traveling parallel to the axis of a) a concave mirror are reflected so that they all pass approximately through a common focal point f, b) a convex mirror are reflected so that they appear to come from a focal point f, located behind the mirror

IMAGE LOCATION BY SPHERICAL MIRROR FORMULA:

How can we talk about the nature of an image (whether the image is real or imaginary, inverted, or erect) formed in a mirror? How can we talk about the size of the image compared with the size of the object? To answer these questions, one method is a graphical or ray diagram. But we can also answer these questions by using a mathematical formula called the mirror formula defined as:

The mirror formula is the relationship between object distance p, image distance q from the mirror, and focal length f of the mirror.

Thus, we can write the mirror formula as:

$$\frac{1}{f} = \frac{1}{p} + \frac{1}{q} \qquad .... \qquad (12.1)$$

Equation (12.1) is true for both concave and convex mirrors. However, sign conventions should be followed to apply this equation for solving problems related to mirrors.

Sign Conventions

Quantity	When Positive (+)	When Negative (−)
Object distance p	Real object	Virtual object
Image distance q	Real Image	Virtual image
Focal length f	Concave mirror	Convex mirror

12.3 REFRACTION OF LIGHT

If we dip one end of a pencil or some other object into water at an angle to the surface, the submerged part looks bent as shown in Fig.12.7. Its image is displaced because the light coming from the underwater portion of the object changes direction as it leaves the water. This bending of light as it passes from one transparent medium into another is called refraction.
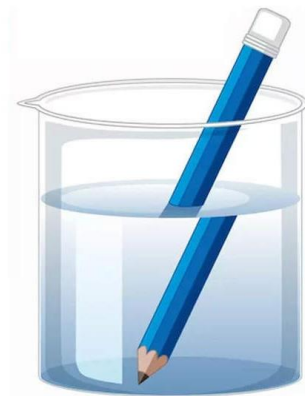
Refraction of light can be explained with the help of Fig.12.8. A ray of light IO traveling from air falls on the surface of a glass block.
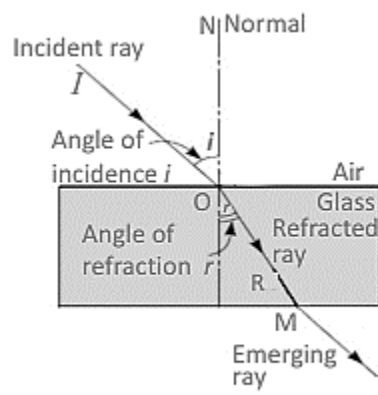
Fig 12.7

Fig. 12.8: Refraction of light by a glass block

At the air-glass interface, the ray of light IO changes direction bends towards the normal, and travels along the path OR inside the glass block. The rays and OR are called the incident ray and IO the refracted ray respectively. The angle 'i' made by the incident ray with the normal is called the angle of incidence. The angle 'r' made by the refracted ray with the normal is called the angle of refraction. When a refracted ray leaves the glass, it bends away from the normal and travels along a path ME.

Thus, the process of bending of light as it passes from air into glass and vice versa is called refraction of light.

12.3.1 LAWS OF REFRACTION

The incident ray, the refracted ray, and the normal at the point of incidence all lie in the same plane.

The ratio of the sine of the angle of incidence 'i' to the sine of the angle of refraction 'r' is always equal to the:

$$\text{constant i.e., } \sin i / \sin r = \text{constant} = n$$

where the ratio sin i / sin r is known as the refractive index of the second medium with respect to the first medium. So, we have

$$\frac{\sin i}{\sin r} = n \quad \text{....... (12.2)}$$

It is called Snell's law.

12.3.2 Speed of light in a medium

Refraction of light is caused by the difference in speed of light in different media. For example, the speed of light in air is approximately 3.0×10^8 ms^{-1} However, when light travels through a medium, such as water or glass, its speed decreases. The speed of light in water is approximately 2.3×10^8 m/s, while in glass, it is approximately 2.0×10^8 m/s.1 To describe the change in the speed of light in a medium, we use the term index of refraction or refractive index.

Refractive Index

The refractive index 'n' of a medium is the ratio of the speed of light 'c' in air to the speed 'v' of light in the medium:

$$\text{Refractive Index} = \frac{\textit{Speed of light in air}}{\textit{Speed of light in medium}}$$

$$\text{or} \quad n = \frac{c}{v} \quad \text{........ (12.3)}$$

12.4 TOTAL INTERNAL REFLECTION

When a ray of light traveling in a denser medium enters a rarer medium, it bends away from the normal (Fig.12.9-a). If the angle of incidence 'i' increases, the angle of refraction 'r' also increases. For a particular value of the angle of o incidence, the angle of refraction becomes 90. The angle of incidence, that causes the refracted ray in the rarer medium o to bend through 90 is called critical angle (Fig.12.9-b). When the angle of incidence becomes larger than the critical angle, no refraction occurs. The entire light is reflected into the denser medium (Fig.12.9-c). This is known as the total internal reflection of light.
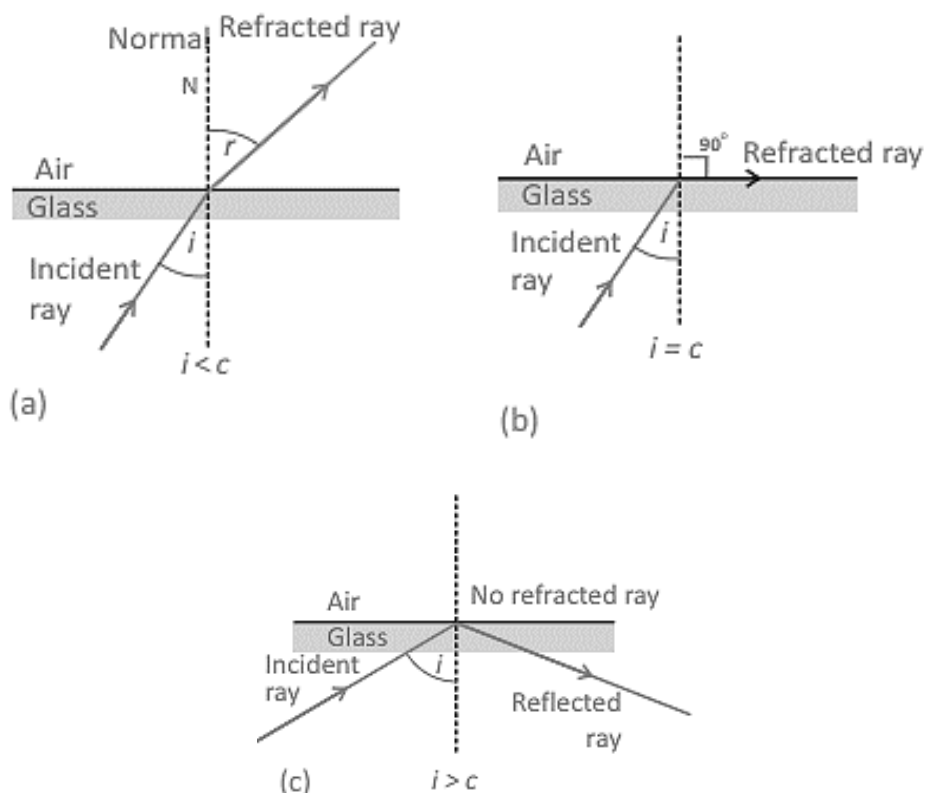
Fig. 12.9: Condition for total internal reflection

12.4.1 APPLICATIONS OF TOTAL INTERNAL REFLECTION:

i.   Totally Internal Reflecting Prism

Many optical instruments use right-angled prisms to reflect a beam of light through 90° or 180° (by total internal reflection) such as cameras, binoculars, periscope, and telescopes. One of the angles of a right-angled prism is 90°. When a ray of light strikes the face of the prism perpendicularly, it enters the prim without deviation and strikes the hypotenuse at an angle of 45° (Fig.12.10). Since the angle of incidence 45° is greater than the critical angle of the glass which is 42°, the light is reflected by the prism through an angle of 90°. Two such prisms are used in periscope (Fig.12.11). In Fig.12.12, the light is totally reflected by the prism at an angle of 180°. Two such prisms are used in binoculars (Fig.12.13).
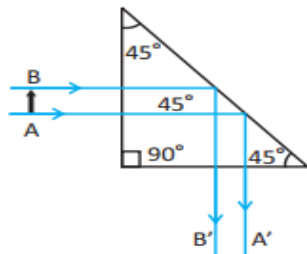
Fig.12.10: Total internal reflection through right angled prism
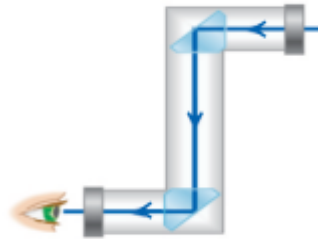
Fig. 12.11: Prism periscope
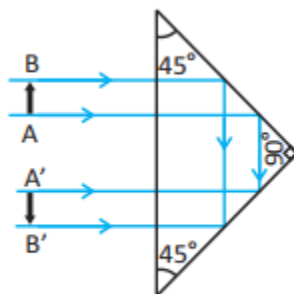
Fig. 12.12
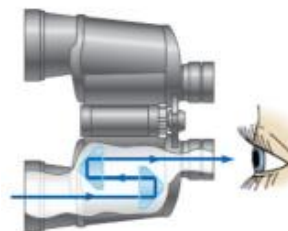
Fig. 12.13: Binoculars

ii.   Optical Fibre

Total internal reflection is used in fibre optics which has a few advantages in the telecommunication field. Fibre optics consists of hair-size threads of glass or plastic through which light can be traveled (Fig. 12.14). The inner part of the fiber optics is called the core that carries the light, and an outer concentric shell is called cladding. The core is made from glass or plastic with a relatively high index of refraction. The cladding is made of glass or plastic but of relatively low refractive index. Light entering from one end of the core strikes the core-cladding boundary at an angle of incidence greater than the critical angle and is reflected into the core (Fig. 12.14). In this way, light travels many

kilometers with a small loss of energy. In Pakistan, optical fiber is being used in telephone and advanced telecommunication systems. Now we can listen to thousands of phone calls without any disturbance.
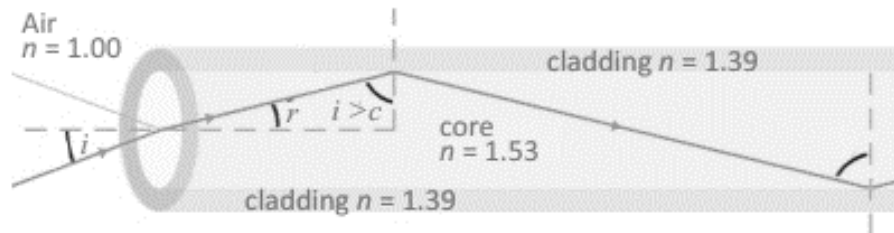
Fig.12.14: Passage of light through optical fibre

iii.   Light Pipe

A light pipe is a bundle of thousands of optical fibers bound together. They are used to illuminate inaccessible places the doctors or engineers. For example, doctors view inside the human body. They can also be used to transmit images from one place to another (Fig. 12.15).
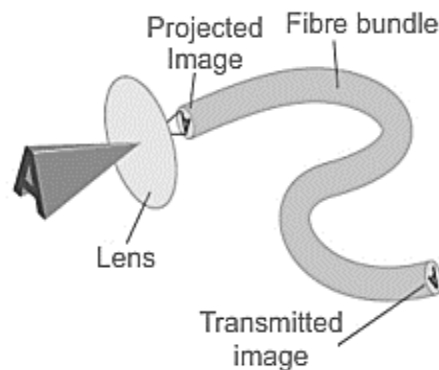
Fig.12.15: A lens and light pipe can be used together to produce a magnified transmitted image of an object

iv.   Endoscope

An endoscope is a medical instrument used for exploratory diagnostics and surgical purposes. An endoscope is used to explore the interior organs of the body. Due to its small size, it can be inserted through the mouth and thus eliminates the invasive surgery. The endoscopes used to examine the stomach, bladder, and throat are called Gastroscope, Cystoscope, and Bronchoscope respectively. An endoscope uses two fiber-optic tubes through a pipe.
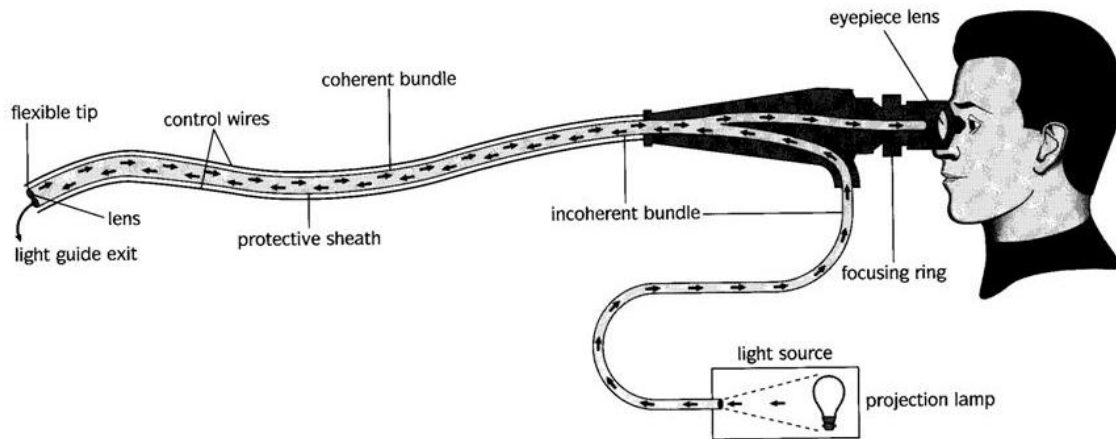
Fig 12.16: Endoscope

A medical procedure using any type of endoscope is called endoscopy. The light shines on the organ of the patient to be examined by entering through one of the fiber tubes of the endoscope. Then light is transmitted back to the physician's viewing lens through the other fibre tube by total internal reflection (Fig.12.16). Flexible endoscopes have a tiny camera attached to the end. A doctor can see the view recorded by the camera on a computer screen.

12.5 LENSES:

A lens is any transparent material having two surfaces, of which at least one is curved. Lenses refract light in such a way that an image of the object is formed. Lenses of many different types are used in optical devices such as cameras, eyeglasses, microscopes, telescopes, and projectors. They also enable millions of people to see clearly and read comfortably.
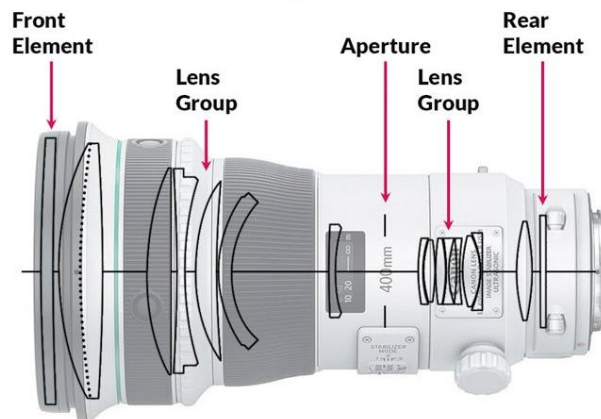
Fig.12.17: The anatomy of a camera lens.

12.5.1 Types of Lenses

There are different types of lenses. The lens which causes incident parallel rays to converge at a point is known as a convex or converging lens. This lens is thick at the center but thin at the edges (Fig.12.18). Another type of lens causes the parallel rays of light to diverge from a point. This is called a concave or diverging lens. This lens is thin at the center and thick at the edges (Fig.12.19).

Double convex   Plano-convex   Concavo-convex

Fig.12.18: Convex lenses

Double concave   Plano-concave   Convexo-concave

Fig.12.19: Concave lenses

Fig. 12.20: Convex lens
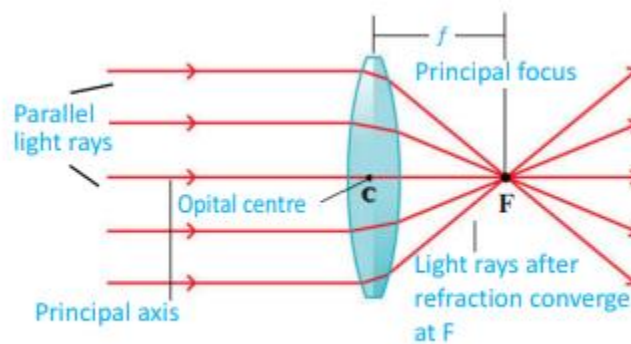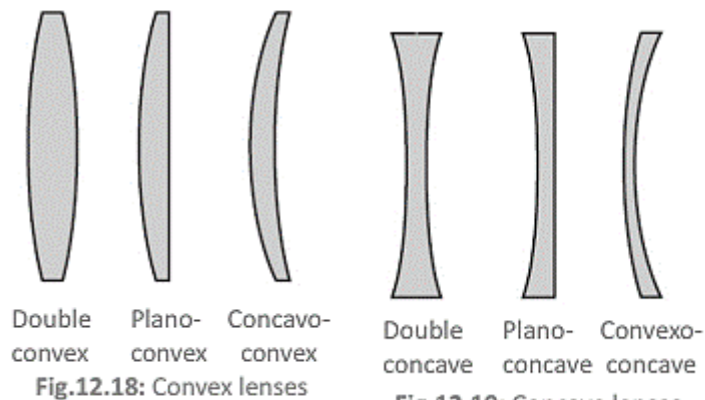
12.5.2 Lens Terminology Principal Axis:

Each of the two surfaces of a spherical lens is a section of a sphere. The line passing through the two centres of curvatures of the lens is called the principal axis (Fig. 12.20). Optical Centre, C: A point on the principal axis at the centre of the lens is called the optical centre (Fig. 12.20).

Principal Focus, F:

The light rays traveling parallel to the principal axis of a convex lens after refraction meet at a point on the principal axis, called principal focus or focal point F. Hence, a convex lens is also called converging lens. For a concave lens, the parallel rays appear to come from a point behind the lens called principal focus F (Fig. 12.21). Hence concave lens is also called diverging lens.

Focal Length, f: This is the distance between the optical centre and the principal focus (Fig. 12.21).
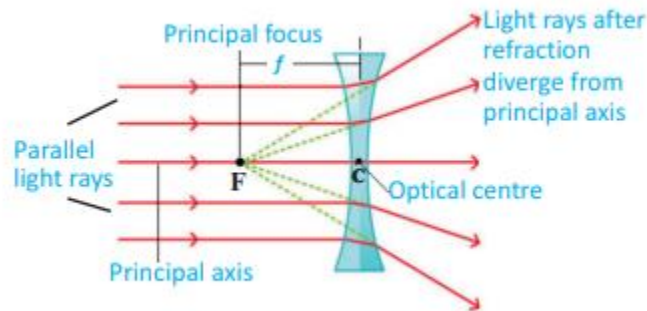
Fig. 12.21: Concave lens

12.5.3 Power of a Lens

Power of a lens is defined as the reciprocal of its focal length in metres. Thus

Power of a lens = P = 1 / focal length in metres

The SI unit of power of a lens is "Dioptre", denoted by a -1 symbol D. If f is expressed in metres so that 1 D = 1 m. Thus, 1 Dioptre is the power of a lens whose focal length is 1 metre. Because the focal length of a convex lens is positive, therefore, its power is also positive. Whereas the power of a concave lens is negative, for it has negative focal length.

12.5.4 IMAGE FORMATION BY LENSES:
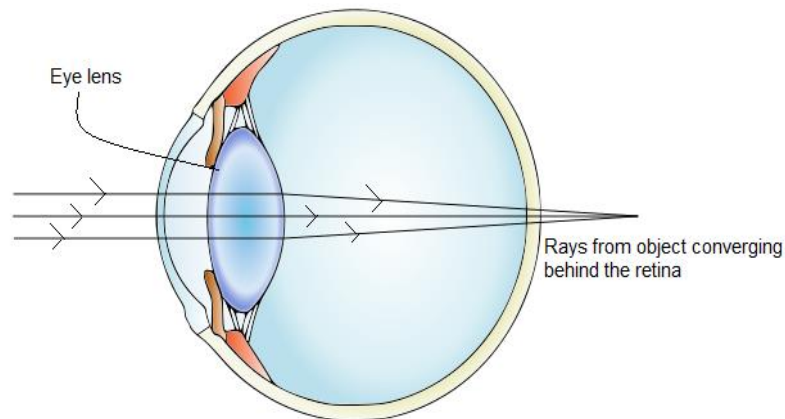
Fig.12.22

In mirrors, images are formed through reflection, but lenses form images through refraction. This is explained with the help of ray diagrams as follows: Image formation in convex lens can be explained with the help of three principal rays shown in Fig.12.23

1. The ray parallel to the principal axis passes through the focal point after refraction by the lens.

2. The ray passing through the optical centre passes straight through the lens and remains undeviated.

3. The ray passing through the focal point becomes parallel to the principal axis after refraction by the lens.
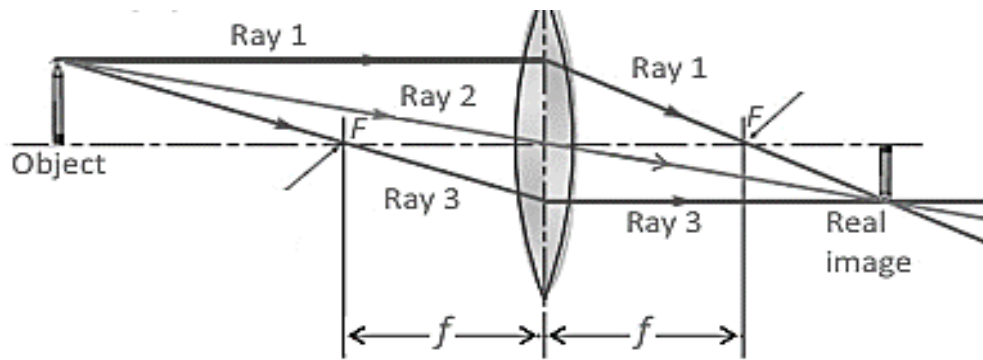
Fig. 12.23: Convex Lens

The ray diagram for concave lens is shown in Fig.12.24.
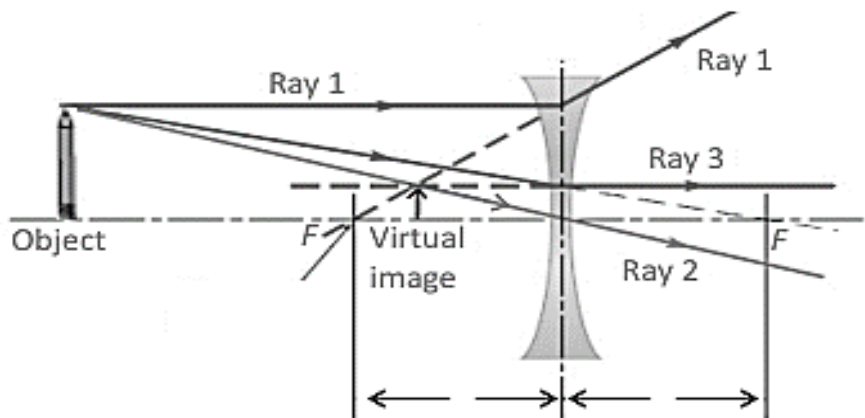
Fig. 12.24: Concave Lens

Image Formation in Convex Lens In class VIII, we have learned image formation by lenses. Let us briefly revise the image formation by convex lens (Fig.12.25).
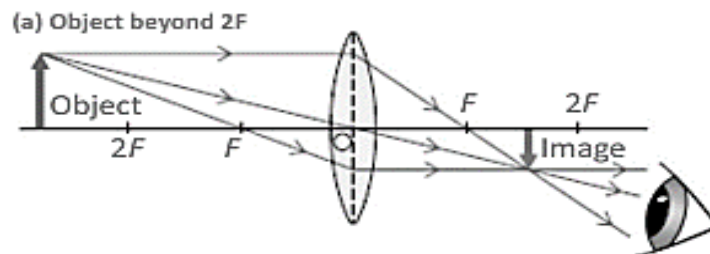
(a) Object beyond 2F

The image is between F and 2F, real, inverted, smaller than the object.
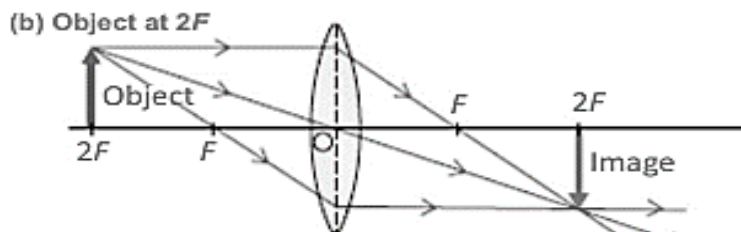
(b) Object at 2F

The image is at 2F, real, inverted, the same size as the object.

(c) Object between *F* and *2F*

Object

2F   F   F   2F
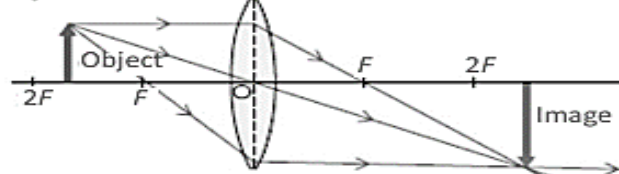
Image

The image is beyond 2F, real, inverted, larger than the object.

(d) Object at *F*
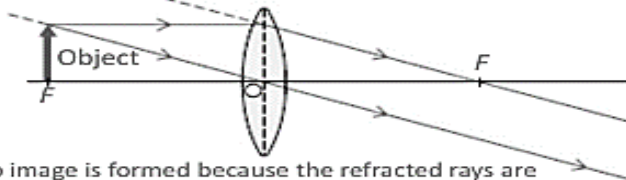
Object   F

No image is formed because the refracted rays are
parallel and never meet.

(e) Object between lens and *F*

Image

Object   F

F
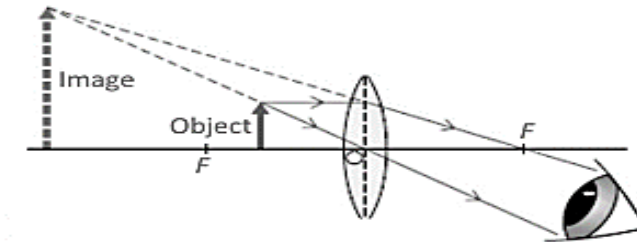
The image is behind the object, virtual, erect, larger than the object.
Fig. 12.25

12.5.5 IMAGE LOCATION BY LENS EQUATION

In Fig.12.26, let an object OP is placed in front of a convex lens at a distance p. A ray PR parallel to the principal axis after refraction passes through focus F. Another ray PC meets the first ray at point P' after passing through the optical centre C. If this process is repeated for the other points of the object, a real and inverted image O'P' is formed at a distance q from the lens.

the lens.   p   → Thin lens

P   R

Image
F   O'

O
Object   F'   C
f   p'
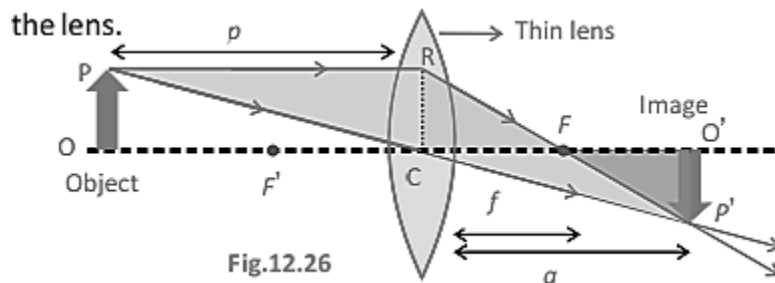
Fig.12.26
q

What is the size of image formed in a lens for a particular distance of object from the lens? What is the nature of the image, i.e., whether image is real or imaginary, erect or

inverted? Lens formula is a tool that we use to answer all such questions. We define lens as,formula

The relation between the object and image distance from the lens in terms of the focal length of the lens is called lens formula.

$$\frac{1}{f} = \frac{1}{p} + \frac{1}{q} \qquad \text{......... (12.4)}$$

Equation (12.4) is valid for both concave and convex lenses. However, sign conventions should be followed while using this equation to solve problems related to lenses.

Sign Conventions for Lenses Focal Length:

• f is positive for a converging lens

• f is negative for a diverging lens.

Object Distance:

• p is positive if the object is towards the left side of the lens. It is called a real object.

 • p is negative if the object is on the right side of the lens. It is called a virtual object.

Image Distance:

• q is positive for a real image made on the right side of the lens by a real object.

 • q is negative for a virtual image made on the left side of the lens by real object.

12.6 APPLICATIONS OF LENSES

Now we discuss applications of lenses in some optical devices such as cameras, slide projectors and photograph enlargers.

CAMERA:

A simple camera consists of a light-proof box with a converging lens in front and a light-sensitive plate or film at the back. The lens focuses images to be photographed onto the film. In a simple lens camera, the distance between lens and film is fixed which is equal to the focal length of the lens. In the camera, the object is placed beyond 2F. A real, inverted and diminished image is formed in this way as shown in Fig.12.27.
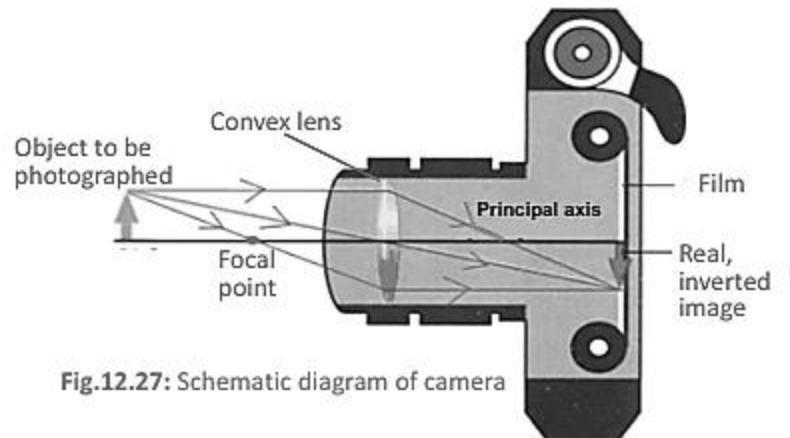
Fig.12.27: Schematic diagram of camera

SLIDE PROJECTOR:

Fig.12.28 shows how a slide or movie projector works. The light source is placed at the centre of curvature of a converging or concave mirror. The concave mirror is used to reflect light in parallel rays. The condenser is made up of 2 converging lenses that refract the light so all parts of the slide are illuminated with parallel rays.
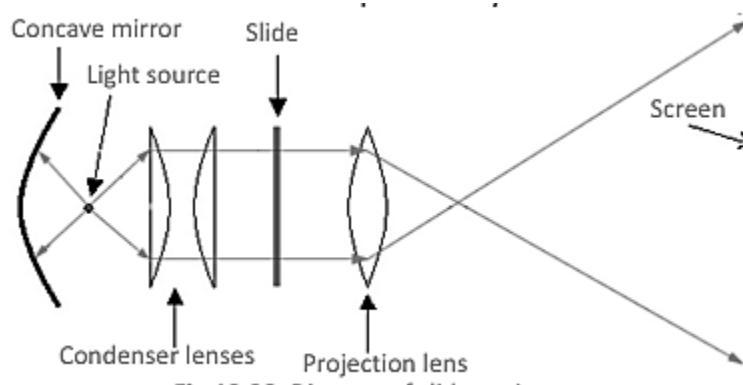
Fig.12.28: Diagram of slide projector

The projection or converging lens provides a real, large and inverted image. It must be real to be projected on a screen. The slide (object) must be placed between F and 2F of projection lens to produce a real, large, and inverted image. Because the image is inverted, the slide must be placed upside down and laterally inverted so we can see the image properly.

EXEXERCISE

1. What do you understand by the reflection of light? Draw a diagram to illustrate reflection at a plane surface.
2. Describe the following terms used in reflection: (i) normal (ii) angle of incidence (iii) angle of reflection.
3. State laws of reflection. Describe how they can be verified graphically.
4. Define refraction of light. Describe the passage of light through parallel-sided transparent material.
5. Define the following terms used in refraction: (i) angle of incidence (ii) angle of refraction.
6. What is meant by the refractive index of a material? How would you determine the refractive index of a rectangular glass slab?
7. State the laws of refraction of light and show how they may be verified using rectangular glass slabs and pins.
8. What is meant by the term total internal reflection?
9. State the conditions for total internal reflection.
10. What is the critical angle? Derive a relationship between the critical angle and the refractive index of a substance.
11. What are optical fibers? Describe how total internal reflection is used in light propagating through optical fibers.
12. Define the following terms applied to a lens: (i) principal axis (ii) optical center (iii) focal length.
13. What is meant by the principal focus of a (a) convex lens and (b) concave lens? Illustrate your answer with ray diagrams.
14. Describe how light is refracted through convex lens.

NUMERICAL PROBLEMS

1. An object 10.0 cm in front of a convex mirror forms an image 5.0 cm behind the mirror. What is the focal length of the mirror? Ans. (-10 cm)
2. An object 30 cm tall is located 10.5 cm from a concave mirror with focal length 16 cm.
   (a) Where is the image located?
   (b) How high is it? Ans. [(a) 30.54 cm (b) 87.26 cm]
3. An object and its image in a concave mirror are of the same height, yet inverted, when the object is 20 cm from the mirror. What is the focal length of the mirror? Ans. (10 cm)
4. Find the focal length of a mirror that forms an image 5.66 cm behind the mirror of an object placed at 34.4 cm in front of the mirror. Is the mirror concave or convex? Ans. (-6.77 cm, Convex mirror)

5. An image of a statue appears to be 11.5 cm behind a concave mirror with focal length 13.5 cm. Find the distance from the statue to the mirror. Ans. (77.62 cm)

6. An image is produced by a concave mirror of focal length 8.7 cm. The object is 13.2 cm tall and at a distance 19.3 cm from the mirror.
    (a) Find the location and height of the image.
    (b) Find the height of the image produced by the mirror if the object is twice as far from the mirror. Ans. [(a) 10.83 cm (b) 5.42 cm] 15.84 cm,

7. Nabeela uses a concave mirror when applying makeup. The mirror has a radius of curvature of 38 cm.
    (a) What is the focal length of the mirror?
    (b) Nabeela is located 50 cm from the mirror. Where will her image appear?
    (c) Will the image be upright or inverted? Ans. [(a) 19 cm, (b) 30.64 cm, (c) upright]

8. An object 4 cm high is placed at a distance of 12 cm from a convex lens of focal length 8 cm. Calculate the position and size of the image. Also state the nature of the image. Ans. (24 cm, 8 cm, image is real, inverted and magnified)

9. An object 10 cm high is placed at a distance of 20 cm from a concave lens of focal length 15 cm. Calculate the position and size of the image. Also, state the nature of the image. Ans. (-8.57 cm, 4.28 cm, image is virtual, erect and diminished)

10. A convex lens of focal length 6 cm is to be used to form a virtual image three times the size of the object. Where must the lens be placed? Ans. (4 cm)

11. A ray of light from air is incident on a liquid surface at an angle of incidence 35 . Calculate the angle of refraction if the refractive index of the liquid is 1.25. Also calculate the critical angle between the liquid air inter-face. Ans. (27.31°, 53.13°)

12. The power of a convex lens is 5 D. At what distance the object should be placed from the lens so that its real and 2 times larger image is formed. Ans. (30 cm).

Chapter # 13

Charge

Students will be able to understand:

- Coulomb's law
- Capacitor and capacitance
- Capacitor in series and in parallel.

The study of charges at rest is called electrostatics or static electricity.

If we run a plastic comb through our hair and then bring it near small pieces of paper, the comb attracts them (Fig.13.1). Similarly, amber when rubbed with silk, attracts the small pieces of paper. This property of attraction or repulsion between substances is due to the electric charges they acquire during rubbing.

Fig.13.1: Comb rubbed with hair attracts small pieces of paper

We can produce electric charge by rubbing a neutral body with another neutral body. The following activities show that we can produce two types of electric charges through the process of rubbing.

Always remember these things that:

- Charge is a basic property of a material body due to which it attracts or repels another object.
- Friction produces two different types of charge on different materials (such as glass and plastic).
- Like charges always repel each other.
- Unlike charges always attract each other.
- Repulsion is the sure test of charge on a body.

13.1 COULOMB'S LAW

We know that a force of attraction or repulsion acts between two charged bodies. How is this force affected when the magnitude of the charge on the two bodies or the distance between them is changed?

To find the answers to these questions, a French scientist Charles Coulomb (1736–1806) in 1785 experimentally established the fundamental law of electric force between two stationary Coulomb's Law:

The force of attraction or repulsion between two-point charges is directly proportional to the product of the magnitude of charges and inversely proportional to the square of the distance between them.

Therefore,
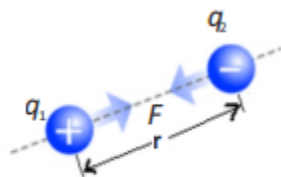
Fig.13.2 (a) Attraction between opposite charges
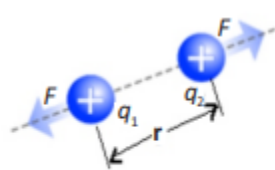
Fig. 13.2(b) Repulsion between similar charges

$$F \sqrt{\frac{1}{r^2}}$$

.... (13.1)

.... (13.2)

Combining Eqs. (13.1) and (13.2), we get

$$F = k \frac{q_1 q_2}{r^2}$$

.... (13.3)

Eq. (13.3) is known as Coulomb's law.

Where F is the force between the two charges and is called the Coulomb force, q_1 and q_2 are the of two magnitudes charges and 'r' is the distance between the two charges (Fig.13.2). k is the constant of proportionality.

The value of k depends upon the medium between the two charges. If the medium between the two charges is air, then the value of k in SI units will be $9 \times 10^9 Nm^2 C^{-2}$.

Coulomb's law is true only for point charges whose sizes are very small as compared to the distance between them.

13.2 Capacitor and capacitance

To store the charge, a device which is called capacitor is used. It consists of two thin metal plates, parallel to each other separated by a very small distance (Fig. 13.3). The medium between the two plates is air or a sheet of insulators. This medium is called dielectric.
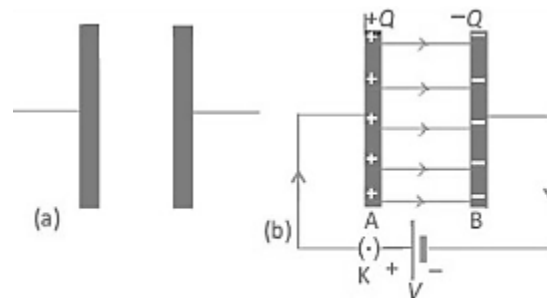
Fig.13.3 (a) Parallel plate capacitor **Fig. 13.13 (b)** plates of capacitor connected with battery

If a capacitor is connected to a battery of V volts, then the battery transfers a charge +Q from plate B to plate A, so that -Q charge appears on plate A and +Q charge appears on plate B.

The charges on each plate attract each other and thus remain bound within the plates. In this way, the charge is stored in a capacitor for a long time. Also, the charge Q stored on plates is directly proportional to the potential difference V across the plates i.e.,

$$Q \quad V$$
$$Q = CV \qquad ....(13.8)$$

where C is the constant of proportionality, called the capacitance of the capacitor and is defined as the ability of the capacitor to store charge. It is given by the ratio of charge and the electric potential as:

$$C = \frac{Q}{V}$$

SI unit of capacitance is farad (F), defined as: If one coulomb of charge given to the plates of a capacitor produces a potential difference of one volt between the plates of the capacitor then its capacitance would be one farad. farad is a large unit, usually, we use a smaller unit such as micro farad (μF), nano farad (nF) and pico farad (pF) etc.

13.2.1 Combinations of Capacitors

Capacitors are manufactured with different standard capacitances, and by combining them in series or in parallel, we can get any desired value of the capacitance.

13.2.1.1 Capacitors in Parallel

In this combination, the left plate of each capacitor is connected to the positive terminal of the battery by a conducting wire. In the same way, the right plate of each capacitor is connected to the negative terminal of the battery (Fig. 13.14).

This type of combination has the following characteristics:

Each capacitor connected to a battery of voltage V has the same potential difference V across it. i.e.,

$$V_1 = V_2 = V_3 = V$$
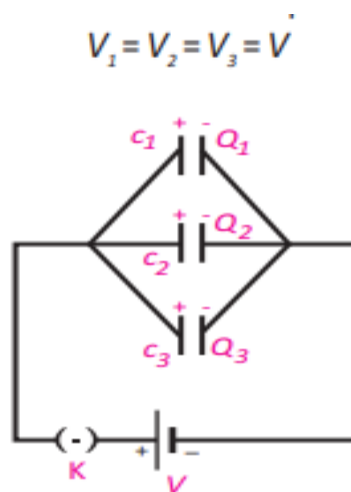
Fig.13.4: Capacitors in parallel combination

The charge developed across the plates of each capacitor will be different due to the different value of capacitances.

The total charge Q supplied by the battery is divided among the various capacitors. Hence,

$$Q = Q_1 + Q_2 + Q_3$$

$$Q = C_1V + C_2V + C_3V$$

or $$\frac{Q}{V} = C_1 + C_2 + C_3$$

Thus, we can replace the parallel combination of capacitors with one equivalent capacitor having capacitance C , such that

$$C_{eq} = C_1 + C_2 + C_3$$

In the case of 'n' capacitors connected in parallel, the equivalent capacitance is given by

$$C_{eq} = C_1 + C_2 + C_3 ....C_n .....   \qquad (13.9)$$

The equivalent capacitance of a parallel combination of capacitors is greater than any of the individual capacitances.

13.2.1.2 Capacitors in Series:

In this combination, the capacitors are connected side by side i.e., the right plate of one capacitor is connected to the left plate of the next capacitor (Fig. 13.15). This type of combination has the following characteristics:
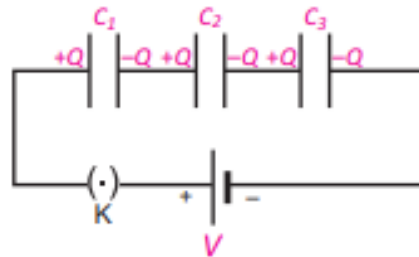
Fig.13.5: capacitors in series combination.

Each capacitor has the same charge across it. If the battery supplies + Q charge to the left plate of the capacitor C1, due to induction – Q charge is induced on its right plate and +Q charge on the left plate of the capacitor C i.e.,

$$Q_1 = Q_2 = Q_3 = Q$$

The potential difference across each capacitor is different due to different values of capacitances.

The voltage of the battery has been divided among the various capacitors. Hence

$$V = V_1 + V_2 + V_3$$
$$= \frac{Q}{C_1} + \frac{Q}{C_2} + \frac{Q}{C_3}$$
$$= Q\left[\frac{1}{C_1} + \frac{1}{C_2} + \frac{1}{C_3}\right]$$
$$\frac{V}{Q} = \left[\frac{1}{C_1} + \frac{1}{C_2} + \frac{1}{C_3}\right]$$

Thus, we can replace series combination of capacitors with one equivalent capacitor having capacitance C i.e.,

$$\frac{1}{C_{eq}} = \frac{1}{C_1} + \frac{1}{C_2} + \frac{1}{C_3}$$

In the case of 'n' capacitors connected in series, we have

$$\frac{1}{C_{eq}} = \frac{1}{C_1} + \frac{1}{C_2} + \frac{1}{C_3} + \cdots + \frac{1}{C_n} \qquad \ldots(13.10)$$

13.2.2 Uses of Capacitors

Capacitors have a wide range of applications in different electrical and electronic circuits. For example, they are used for tuning transmitters, receivers, and transistor radios. They are also used for table fans, ceiling fans, exhaust fans, fan motors in air conditioners, coolers, motors washing machines, air conditioners, and many other appliances for their smooth working. Capacitors are also used in electronic circuits of computers etc. Capacitors can be used to differentiate between high-frequency and low-frequency signals which makes them useful in electronic circuits. For example, capacitors are used in the resonant circuits that tune radios to frequencies. Such circuits are called filter circuits. One type of capacitor may not be suitable for all applications. Ceramic capacitors are generally superior to other types and therefore can be used in vast ranges of applications.

EXERCISE

1. Explain Coulomb's law of electrostatics and write its mathematical form.
2. How can you show by simple experiments that there are two types of electric charges? What do you mean by the capacitance of a capacitor? Define units of capacitance.
3. Derive the formula for the equivalent capacitance for a series combination of several capacitors.
4. How much negative charge has been removed from a positively charged −11 electroscope, if it has a charge of 7.5×10^{-11} C?
5. Enlist some uses of capacitors.
6. The charge of how many negatively charged particles would be equal to 100 μC. Assume charge on one negative particle is 1.6×10^{-19} C? Ans. (6.25×10^{14})
7. Two-point charges $q_1 = 10$ μC and $q_2 = 5$ μC are placed at a distance of 150 cm. What will be the Coulomb's force between them? Also, find the direction of the force. Ans. (0.2 N, the direction of repulsion)
8. The force of repulsion between two identical positive charges is 0.8 N, when the charges are 0.1 m apart. Find the value of each charge. Ans. $(9.4 \times 10^{-7}$C$)$
9. Two charges repel each other with a force of 0.1 N when they are 5 cm apart. Find the forces between the same charges when they are 2 cm apart. Ans. (0.62 N)
10. A capacitor holds 0.06 coulombs of charge when fully charged by a 9 volt battery. Calculate capacitance of the capacitor. Ans. $(6.67 \times 10^{-3}$F$)$
11. A capacitor holds 0.03 coulombs of charge when fully charged by a 6 volt battery. How much voltage would be required for it to hold 2 coulombs of charge?

Ans.(400 V)

Chapter # 14

Alternate Current

Students will be able to understand:

- Basic definition of A.C.
- RMS value
- Peak value
- Sine wave

14.1 Introduction:

An **alternating current** can be defined as a current that changes its magnitude and polarity at regular intervals of time. It can also be defined as an electrical current that repeatedly changes or reverses its direction opposite to that of Direct Current or DC, which always flows in a single direction as shown below.
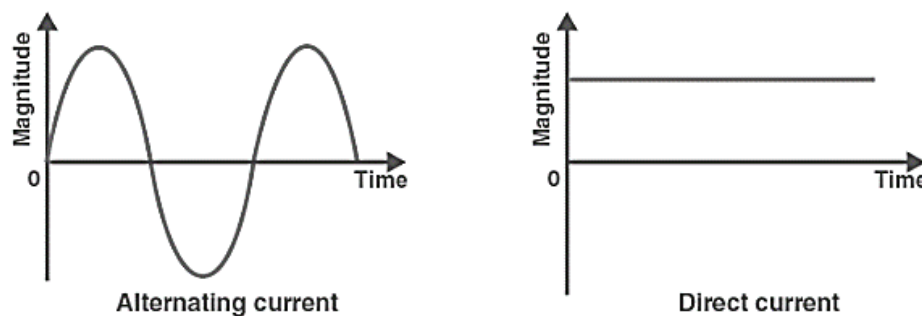
Fig 14.1

From the graph, we can see that the charged particles in AC tend to start moving from zero. It increases to a maximum and then decreases back to zero completing one positive cycle. The particles then reverse their direction and reach the maximum in the opposite direction after which AC again returns to the original value completing a negative cycle. The same cycle is repeated and again.

Alternating currents are also accompanied usually by alternating voltages. Besides, alternating current is also easily transformed from a higher voltage level to a lower voltage level.

14.1.1 Alternating Current Production

Alternating current can be produced or generated by using devices that are known as alternators. However, alternating current can also be produced by different methods

where many circuits are used. One of the most common or simple ways of generating AC is by using a basic single coil AC generator, which consists of two-pole magnets and a single loop of wire having a rectangular shape.

In this setup, the AC generator follows Faraday's principle of electromagnetic induction where it converts mechanical energy into electrical energy.

Meanwhile, AC is supplied to pieces of equipment using 3 wires. They are as follows.

Power is transmitted by the hot wire.

The neutral wire which is connected to the earth provides a return path for the current in the hot wire.

The third wire that is also connected to the earth is linked to the metallic parts of the equipment to mainly eliminate electric shock hazards.

14.1.2 Application of Alternating Current:

AC is the form of current that is mostly used in different appliances.

Some of the examples of alternating current include audio signal, radio signal, etc.

i.   An alternating current has a wide advantage over DC as AC is able to transmit power over large distances without great loss of energy.
ii. AC is used mostly in homes and offices mainly because the generating and transporting of AC across long distances is a lot easier. Meanwhile, AC can be converted to and from high voltages easily using transformers.
iii. AC is also capable of powering electric motors that further convert electrical energy into mechanical energy. Due to this, AC also finds its use in many large appliances like refrigerators, dishwashers, and many other appliances.

14.1.3 Alternating Current Waveform

Before we learn more about this topic, let us quickly understand a few key terms.

i.   The time interval between a definite value of two successive cycles is the **period**.
ii. The number of cycles or number of periods per second is **frequency**.
iii. The maximum value in both directions is the **amplitude**.
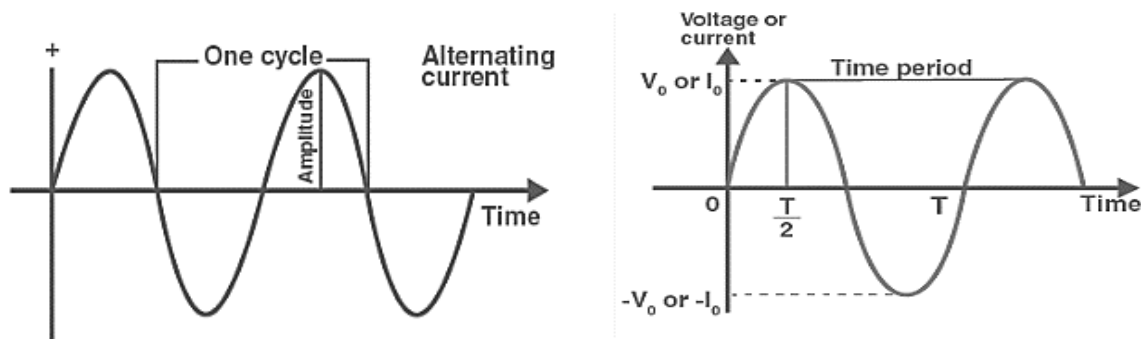
Fig. 14.2

The normal waveform of AC in most of the circuits is sinusoidal in nature in which the positive half period corresponds with the positive direction of the current and vice-versa. In addition, a triangular or square wave can also be used to represent the alternating current waveform.

Audio amplifiers that deal with analogue voice or music signals produce irregular AC waves. Some electronic oscillators produce square or saw tooth waves.

14.1.4 Average Value of AC

The average value is usually defined as the average of the instantaneous values of alternating current over a complete cycle. The positive half cycle of asymmetrical waves such as a sinusoidal voltage or current waveform will be equal to the negative half cycle. This implies that the average value after the completion of a full cycle is equal to zero.

Since both cycles do some work, the average value is obtained by avoiding the signs. Therefore, the average value of alternating quantities of sinusoidal waves can be considered by taking the positive cycle only.

14.2 RMS Value of AC Wave:

RMS value is defined as the square root of means of squares of instantaneous values. It can also be described as the amount of AC power that generates the same heating effect as an equivalent DC power.

The RMS value of a set of values (or a continuous-time waveform) is the square root of the arithmetic mean of the squares of the values, or the square of the function that defines the continuous waveform. In physics, the RMS current value can also be defined as the "value of the direct current that dissipates the same power in a resistor."

If we connect an ordinary D.C. ammeter to measure alternating current, it would measure its value as averaged over a cycle. It can be seen in Fig. 14.2 that the average

value of current and voltage over a cycle is zero, but the power delivered during a cycle is not zero because power is I^2R and the values of I^2 are positive even for negative values of *I*. Thus the average value of I^2 is not zero and is called the mean square current. The alternating current or voltage is actually measured by square root of its mean square value known as root mean square (rms) value.

Let us compute the average value of V^2 over a cycle. Fig.14.3 shows an alternating voltage and the way its V^2 values vary. Note that the values of V^2 are positive on the negative half cycle also. As the graph of V^2 is symmetrical about the line $\frac{1}{2} V_o^2$, so for this figure the mean or the average value of V^2 is $\frac{1}{2} V_o^2$. The root mean square value of V is obtained by taking the square root of $\frac{V_o 2}{2}$.

$$V_{rms} = \sqrt{\frac{V_o^2}{2}} = \frac{V_o}{\sqrt{2}} = 0.7 V_o \cdots\cdots$$

Fig 14.3

Similarly $\qquad I_{rms} = \frac{I_o}{\sqrt{2}} = 0.7 I_o$
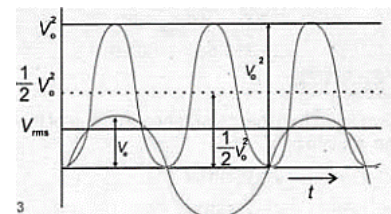
Most of the alternating current and voltage meters are calibrated to read rms values. When we speak of A.C. meter reading, we usually mean rms values unless stated otherwise.

14.3 Peak value of sine wave:

Peak value of sine wave means the maximum value found on a sinusoidal waveform compared to the zero level.

Other commonly used amplitude values are root mean square and peak-to-peak.

Peak of sine is 1.414 times larger than root mean square (RMS) on a sine wave.

Peak of sine is 1/2 of its peak-to-peak value.

means the maximum value found on a sinusoidal waveform compared to the zero level.

Other commonly used amplitude values are root mean square and peak-to-peak.

Peak of sine is 1.414 times larger than root mean square (RMS) on a sine wave.

Peak of sine is 1/2 of its peak-to-peak value.
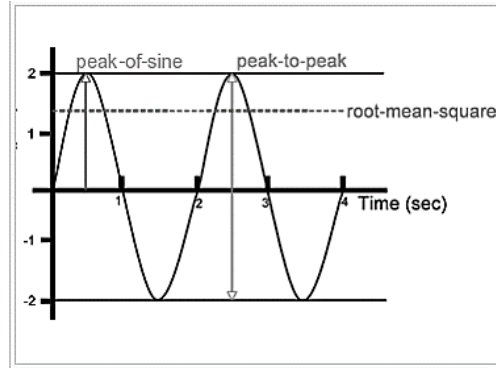
Alternating Current vs Direct Current

Alternating Current	Direct Current
It is safer to transfer over longer distances and helps to maintain electric power.	Cannot be transferred over very long distances. It has a tendency to lose electric power.
AC can change its direction during electric flow.	It does not change its direction during flow and remains constant.
The frequency of AC depends from country to country.	The frequency of DC is zero.
An alternator is used to generate AC.	DC current is generated by generators, photovoltaic cells and batteries.
Electrons keep changing their directions. It can be either backward or forward.	Electrons move only in the forward direction.
Load on AC can be inductive, capacitive or resistive.	Load on DC is always resistive.

EXERCISE

1. Define the period of alternating current.
2. Define the frequency of the alternating current.
3. What is the RMS value of an AC?
4. What are the advantages of AC over DC current?
5. Define the period of alternating current.
6. Differentiate between DC and AC?

Chapter # 15

Magnets and Magnetism

Students will be able to understand:

- Properties
- Magnetic field
- Magnetic lines of force
- Electromagnet
- Magnetic effect of electric current
- Motor and generator effect of current
- Magnetic and electric induction, Transformer.

Electromagnetism is the study of magnetic effects of current. The use of electromagnetism in different fields of science and technology is very wide. Motors and electric meters are based on the effect of magnetism produced by the electric current in wires. Generators produce electric current due to the movement of wires near very large magnets.

15.1 MAGNETIC EFFECTS OF A STEADY CURRENT:

Ampere discovered that when a current passes through a conductor, it produces magnetic field around it. To demonstrate this, we take a straight conductor wire and pass it vertically through a cardboard. Now connect (Fig.15.1-a) the two ends of the conductor wire with the terminals of the battery so that current flows through the circuit in the clockwise direction. The lines of force of the magnetic field produced around the wire would be in the form of concentric circles. If we place a compass needle at different points in the region of magnetic field, it will align along the direction of magnetic field. Also if we sprinkle some iron filings on the cardboard around the wire, they will align themselves in concentric circles in the clockwise direction.
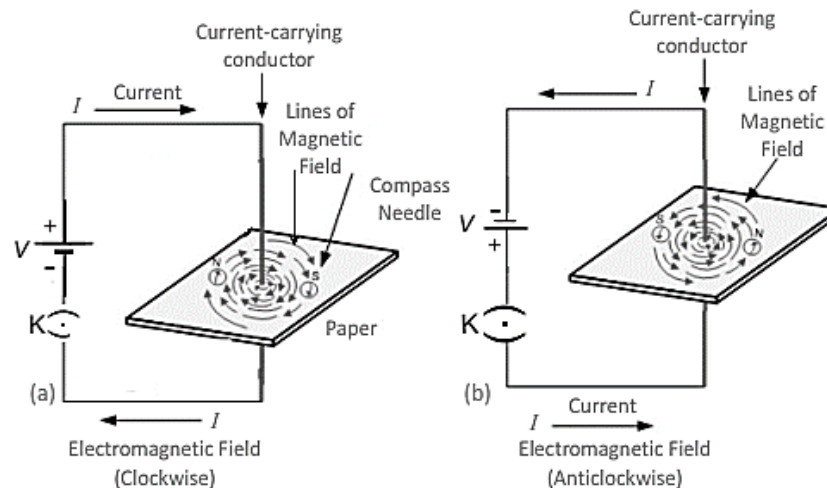
Fig. 15.1

If we reverse the direction of the current by reversing the terminals of the battery, the compass needle also reverses its direction. Now the magnetic field lines will align in the anticlockwise direction (Fig.15.1-b). The magnetic field produced is stronger near the current-carrying conductor and weaker farther away from it.

15.2 Direction of magnetic field:

The direction of the magnetic field is governed by the direction of the current flowing through the conductor. A simple method of finding the direction of magnetic field around the conductor is the Right-Hand Grip Rule.

Grasp a wire with your right hand such that your thumb is pointed in the direction of current. Then curling fingers of your hand will point in the direction of the magnetic field.
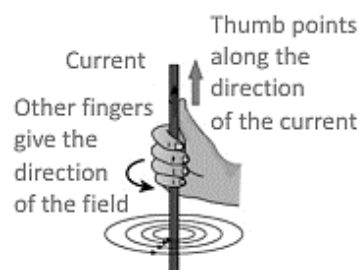
Fig.15.2: Right hand grip rule

15.3 Magnetic field of a solenoid

A coil of wire consisting of many loops is called a solenoid (Fig.15.3). The field from each loop in a solenoid adds to the fields of the other loops and creates greater total field strength. Electric current in the solenoid of wire produces magnetic field which is similar to the magnetic field of a permanent bar magnet. When this current-carrying solenoid is brought close to a suspended bar magnet, one end of the solenoid repels the north pole of the bar magnet. Thus, the current carrying has a north and a south pole and solenoid behaves like a magnet.
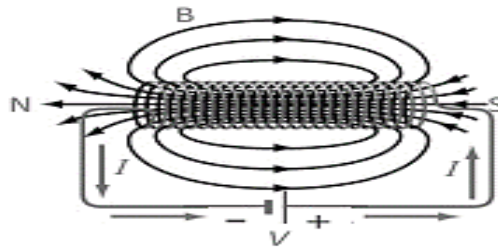
Fig15.3: Magnetic field due to solenoid.

The type of temporary magnet, which is created when current flows through a coil, is called an electromagnet.

The direction of the field produced by a coil due to the flow of conventional current can be found with the help of right hand grip rule (Fig.15.4) stated as
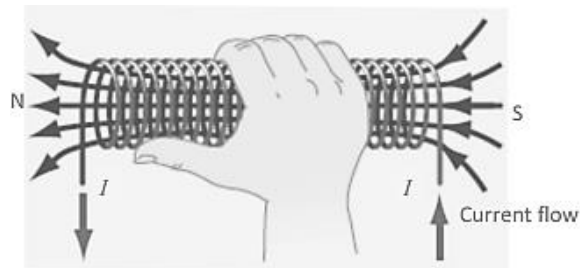
Fig 15.4: Right hand grip rule for a coil.

If we grip the coil with our right hand by curling our fingers in the direction of the conventional current, our thumb will indicate the north pole of the coil.

15.4 D.C. MOTOR:

We can see from Fig. 15.5 that the simple coil placed in a magnet cannot rotate more than 90°. The forces push the PQ side of the coil up and the RS side of the loop down until the loop reaches the vertical position. In this situation, plane of the loop is

perpendicular to the magnetic field and the net force on the coil is zero. So the loop will not continue to turn because the forces are still up and down and hence balanced.
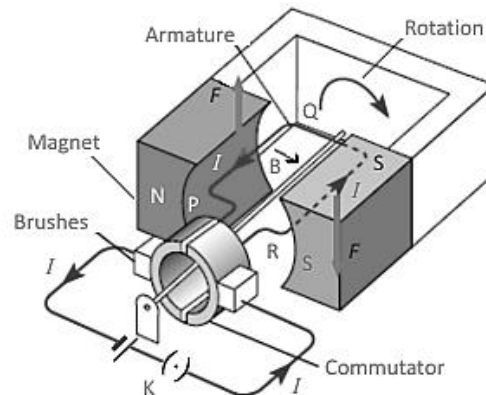
Fig: 15.5: working principle of D.C motor

How can we make this coil to rotate continuously? It can be done by reversing the direction of the current just as the coil reaches its vertical position. This reversal of current will allow the coil to rotate continuously. To reverse direction of current, the connection to coil is made through an arrangement of brushes and a ring that is split into two halves, called a split ring commutator (Fig. 15.9). Brushes, which are usually pieces of graphite, contact the commutator and allow current to flow into the loop. As the loop rotates, so does the commutator. The split ring is arranged so that each half of the commutator changes brushes just as the coil reaches the vertical position. Changing brushes reverse the current in the loop. As a result, the direction of the force on each side of the coil is reversed and it continues to rotate. This process repeats at each half-turn, causing coil to rotate in the magnetic field continuously. The result is an electric motor, which is a device that converts electric energy into rotational kinetic energy.

In a practical electric motor, the coil, called the armature, is made of many loops mounted on a shaft or axle. The magnetic field is produced either by permanent magnets or by an electromagnet, called a field coil. The torque on the armature, and, as a result, the speed of the motor, is controlled by varying the current through the motor.

The total force acting on the armature can be increased by

- Increasing the number of turns of the coil

- Increasing the current in the coil

- Increasing the strength of the magnetic field

- Increasing the area of the coil

15.5 ELECTROMAGNETIC INDUCTION:

Hans Christian Oersted and Ampere discovered that an electric current through a conductor produces a magnetic field around it. Michael Faraday thought that the reverse must also be true; that a magnetic field must produce an electric current. Faraday found that he could induce electric current by moving a wire through a magnetic field. In the same year, Joseph Henry also showed that a changing magnetic field could produce electric current. Now we shall discuss Faraday's experiments for the production of e.m.f. in magnetic field.
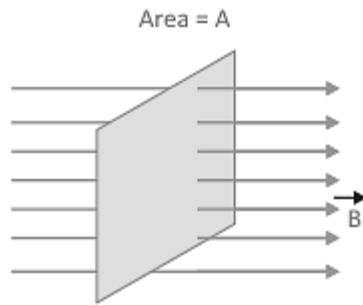
Fig. 15.6 : Maximum strength of magnetic field.   **Fig. 15.7 : Minimum strength of magnetic field.**
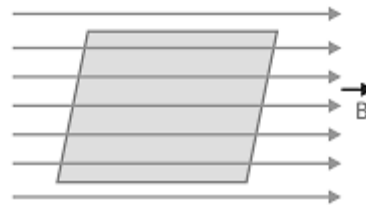
The strength of magnetic field is defined as the number of magnetic lines of force passing through any surface. The number of lines of force is maximum when the surface is held perpendicular to the magnetic lines of force (Fig.15.6). It will be minimum when surface is held parallel to the magnetic lines of force (Fig.15.7). If we place a coil in the magnetic field of a bar magnet, some of the magnetic lines of force will pass through it. If the coil is far away from the magnet, only a few lines of force will pass through the coil (Fig.15.8 -a). However, if the coil is close to the magnet, a large number of lines of force will pass through it (Fig.15.8-b).
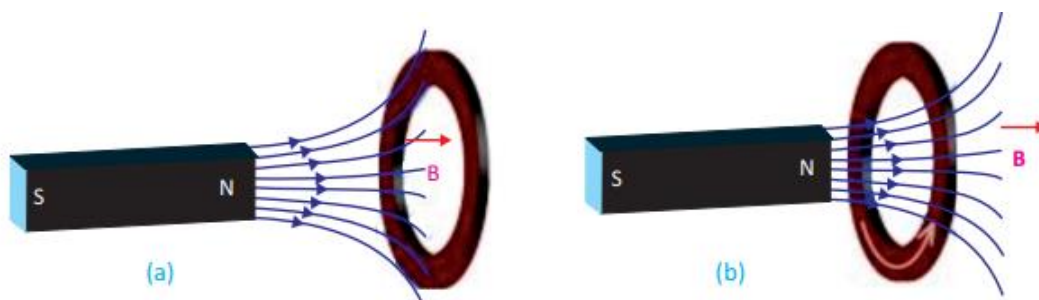
Fig.15.8 : Variation of magnetic field lines of force through a coil placed at different distances from the magnet.

This means, we can change the number of magnetic lines of force through a coil by moving it in the magnetic field. This change in the number of magnetic field lines will induce an e.m.f. in the coil. This is the basic principle of the production of electricity

Activity 15.1:

Take a rectangular loop of wire and connect its two ends with a galvanometer. Now hold the wire stationary or move it parallel to the magnetic field of a strong U-shaped magnet. Galvanometer shows no deflection and hence there is no current. Now move the wire downward through the field, current is induced in one direction as shown by the deflection of the galvanometer (Fig. 15.9-a). Now move the wire upward through the field, current is induced in the opposite direction (Fig. 15.9-b).
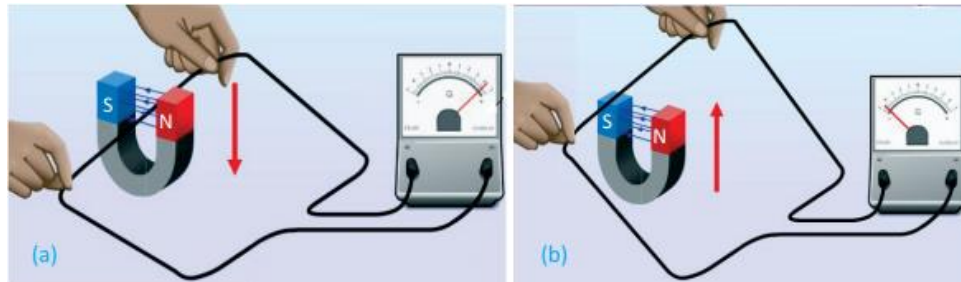
Fig. 15.9: Demonstration of electromagnetic induction by the movement of a wire loop in the magnet field

It implies that an electric current is generated in a wire only when the wire cuts magnetic field lines. This induced current is generated by the induced e.m.f. in the circuit. Faraday found that to generate current, either the conductor must move through a magnetic field or a magnetic field must change across the conductor. Thus,

The process of generating an induced current in a circuit by changing the number of magnetic lines of force passing through it is called electromagnetic induction.

Activity 15.2:

Fig. 15.10 shows one of Faraday's experiments in which current is induced by moving a magnet into the solenoid or out of the solenoid. When the magnet is stationary, no current is induced. When the magnet is moved towards the solenoid, the needle of galvanometer deflects towards right, indicating that a current is being induced in the solenoid (Fig.15.10-a). When the magnet is pulled away from the solenoid, the galvanometer deflects towards left, indicating that the induced current in the solenoid is in the opposite direction (Fig.15.10-b).
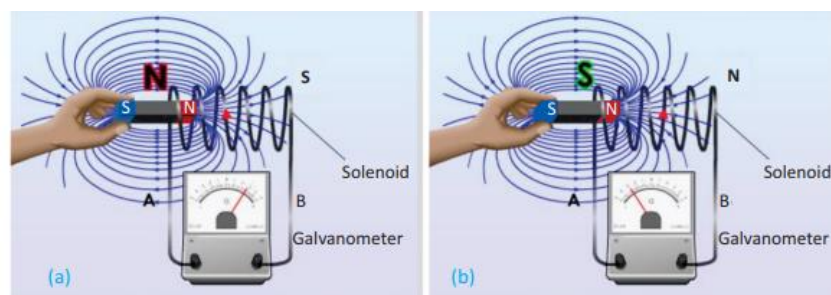
Fig. 15.10: Phenomenon of electromagnetic induction by the movement of a magnet through solenoid. (a) Magnet moves towards the stationary solenoid (b) Magnet moves away from the stationary solenoid

From the above experiments, we conclude that an e.m.f. is induced in the coil when there is a relative motion between the coil and the magnet.

This phenomenon in which an e.m.f. is induced due to the relative motion between the coil and the magnet is called electromagnetic induction. The value of induced e.m.f. in a circuit is directly proportional to the rate of change of number of magnetic lines of force through it.

This is called Faraday's law of electromagnetic induction.

Factors Affecting Induced e.m.f

The magnitude of induced e.m.f. in a circuit depends on the following factors:

1. Speed of relative motion of the coil and the magnet

2. Number of turns of the coil

Direction of induced e.m.f. – Lenz's Law

Lenz devised a rule to find out the direction of a current induced in a circuit. It is explained from the following activity:

Activity 15.:

If we bring a north pole of a bar magnet near a solenoid, an e.m.f. will be induced in the solenoid by electromagnetic induction (Fig. 15.11-a). The direction of the induced current in the solenoid by the induced e.m.f. will be such that it will repel the north pole of the magnet. This is only possible if the right end of the solenoid becomes a north pole. Hence, according to right hand grip rule, the direction of the induced current in the solenoid will be clockwise. Similarly, when we move the north pole of the magnet away from the solenoid, the direction of the induced current will be anticlockwise (Fig.15.11-b). In this case, left end of solenoid becomes south pole.

The direction of an induced current in a circuit is always such that it opposes the cause that produces it.
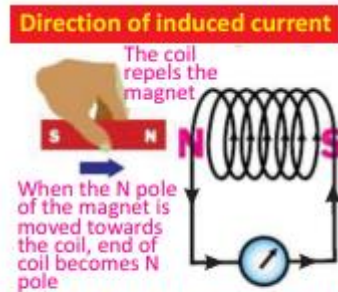
Fig.15.11 (a) Direction of induced current when magnet is moved towards the coil.
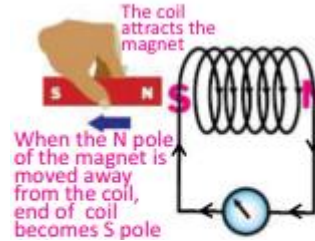
Fig.15.11 (b) Direction of induced current when magnet is moved away from the coil.

If we apply the law of conservation of energy to electromagnetic induction, we realize that the electrical energy induced in a conductor comes from the kinetic energy of the moving magnet. We do some work on the magnet to bring it close to the solenoid. This work consequently appears as electrical energy in the conductor. Thus, mechanical energy of our hand used to push the magnet towards or away from the coil results into electrical energy. Hence, Lenz's law is a manifestation of the law of conservation of energy.

15.6 A.C. GENERATOR:

If a coil is rotated in a magnetic field, a current will be induced in the coil. The strength of this induced current depends upon the number of magnetic lines of force passing through the coil. The number of lines of magnetic force passing through the coil will be maximum when the plane of the coil is perpendicular to the lines of magnetic force. The number of lines of magnetic force will be zero when plane of the coil is parallel to the lines of force.

Fig.15.12 : A.C Generator
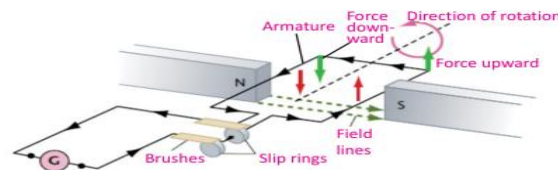
Thus, when a coil rotates in a magnetic field, the induced current in it continuously changes from maximum to minimum value and from minimum to maximum value and so on. This is the basic principle on which an A.C generator works (Fig. 15.12).

The armature is arranged so that it can rotate freely in the magnetic field. As the armature turns, the wire loops cut through the magnetic field lines and induced e.m.f. will be produced. The e.m.f. developed by the generator depends on the length of the wire rotating in the field. Increasing the number of loops in the armature, increases the wire length, thereby increasing the induced e.m.f .Current from a generator

When a generator is connected in a closed circuit, the induced e.m.f. generates an electric current. As the loop rotates, the strength and the direction of the current changes as shown in Fig. 15.13.
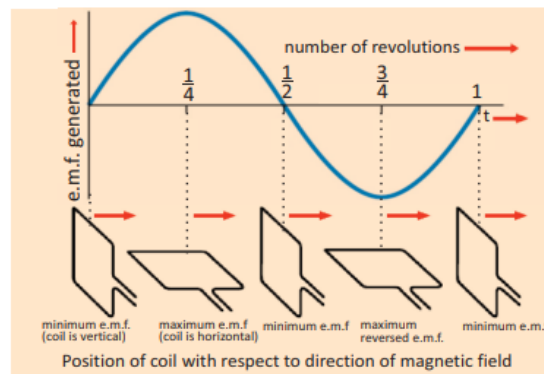
Fig 15.13 : e.m.f Vs time for AC generator

When the plane of will is perpendicular to field, the number of lines of magnetic force passing the trough it is maximum. But the change in the number of line through the coil is minimum. So e.m.f. induced is minimum.

The current is minimum when the plane of the loop is perpendicular to the magnetic field; that is, when the loop is in the vertical position. As the loop rotates from the vertical to the horizontal position, it cuts through larger magnetic field lines per unit of time, thus the e.m.f and the current increase. When the loop is in the horizontal position, the plane of the loop becomes parallel to the field, so the e.m.f and the current reach its maximum value. As the loop continues to turn, the segment that was moving up begins to move down and reverses the direction of the e.m.f and the current in the loop. This change in direction takes place each time the loop turns through 180°. Thus, the e.m.f and the current change smoothly from zero to some maximum values and back to zero during each half-turn of the loop.

15.7 TRANSFORMER:

The transformer is a practical application of mutual induction. Transformers are used to increase or decrease AC voltage . The usage of transformers is common because they change voltages with relatively little loss of energy. In fact, many of the devices in our homes, such as game systems, printers, and stereos use transformers for their work.

15.7.1 Working of a transformer
A transformer has two coils, electrically insulated from each other, but wound around the same iron core. One coil is called the primary coil. The other coil is called the secondary coil. Number of turns on the primary and the secondary coils are represented by N and N respectively. P S When the primary coil is connected to a source of AC voltage, the

changing current creates a changing magnetic field, which is carried through the core to the secondary coil. In the secondary coil, the changing field induces an alternating e.m.f.
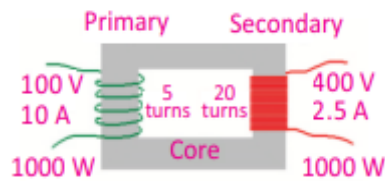
Fig. 15.14 (a) step up transformer
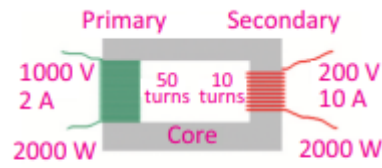
Fig. 15.14 (b) step down transformer

The e.m.f. induced in the secondary coil, called the secondary voltage V , is proportional to the primary voltage V . The S P secondary voltage also depends on the ratio of the number of turns on the secondary coil to the number of turns on the primary coil, as shown by the following expression:

$$\frac{V_s}{V_p} = \frac{N_s}{N_p}$$

If the secondary voltage is larger than the primary voltage, the transformer is called a step-up transformer (Fig.15. 14-a). If the secondary voltage is smaller than the primary voltage, the transformer is called a step-down transformer (Fig.15. 14-b). In an ideal transformer, the electric power delivered to the secondary circuit is equal to the power supplied to the primary circuit. An ideal transformer dissipates no power itself, and for such a transformer, we can write:

$$P_p = P_s$$
$$V_p I_p = V_s I_s$$

EXERCISE

1. State and explain the rule by which the direction of the lines of force of the magnetic field around a current-carrying conductor can be determined.
2. You are given an unmarked magnetized steel bar and bar magnet; its north and south ends are marked N and S respectively. State how would you determine the p o l a r i t y at each end of the unmarked bar?
3. When a straight current-carrying conductor is placed in a magnetic field, it experiences a force. State the rule by which the direction of this force can be found out.
4. What is an electric motor? Explain the working principle of D.C motor.
5. Draw a labelled diagram to illustrate the structure and working of A.C generator.
6. What is a transformer? Explain the working of a transformer in connection with mutual induction.
7. What is the difference between a generator and a motor?
8. Can a transformer operate on direct current?
9. A transformer is needed to convert a mains 240 V supply into a 12 V supply. If there are 2000 turns on the primary coil, then find the number of turns on the secondary coil. Ans. (100)
10. A step-up transformer has a turn ratios of 1 : 100. An alternating supply of 20 V is connected across the primary coil. What is the secondary voltage? Ans. (2000 V)
11. A transformer, designed to convert the voltage from 240 V a.c mains to 12 V, has 4000 turns on the primary coil. How many turns should be on the secondary coil? If the transformer were 100% efficient, what current would flow through the primary coil when the current in the secondary coil was 0.4 A? Ans. (200, 0.02A)

Chapter # 16

Electricity

Students will be able to understand:

- Definition of current
- free electrons
- conductor and insulators,
- P.D.
- Resistance,
- Resistance laws
- Ohm's law
- Circuit
- series circuit
- parallel circuit
- Power and energy.

Charges in motion constitute electric current. This chapter will introduce you to current electricity and related phenomena such as conventional current, Ohm's law, electric power, Joule's heating effect, hazards of electricity and safety measures. We will also learn how current, or voltage is measured in a circuit by electrical devices.

16.1 ELECTRIC CURRENT:

The rate of flow of electric charge through any cross-sectional area is called current.

If the charge Q is passing through any area in time t, then current I flow through it will be given by

$$Current = \frac{Charge}{Time}$$

$$\text{or} \qquad I = \frac{Q}{t} \quad \ldots\ldots\ldots$$

SI unit of current is ampere (A).

If a charge of one coulomb passes through a cross-sectional area in one second, then current is one ampere. Smaller Units of current are milli ampere (mA), micro ampere (µA), which are defined below as:

$$1 \text{ mA} = 10^{-3} \text{ A}$$
$$1 \text{ μA} = 10^{-6} \text{ A}$$

Battery is one of the sources of current. The electrochemical reaction inside a battery separates positive and negative electric charges. This separation of charges sets up potential difference between the terminals of the battery. When we connect a conducting wire across the terminals of the battery, the charges can move from one terminal to the other due to the potential difference. The chemical energy of the battery changes to electrical potential energy. The electrical potential energy decreases as the charges move around the circuit.
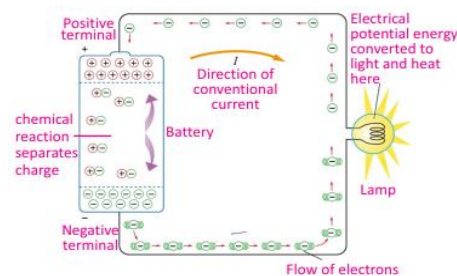
Fig. 16.1

This electrical potential energy can be converted to other useful forms of energy (heat, light, sound etc.). It is only the energy which changes form but the number of charge carriers and the charge on each carrier always remains the same (i.e., charge carriers are not used up). Instead of electrical potential energy we use the term electric potential which is potential energy per unit charge.

16.1.1 Conventional Current:

Conventional current is defined as:

Current flowing from positive to negative terminal of a battery due to the flow of positive charges is called conventional current Conventional current produces the same effect as the current flowing from negative terminal to the positive terminal due to the flow of negative charges.
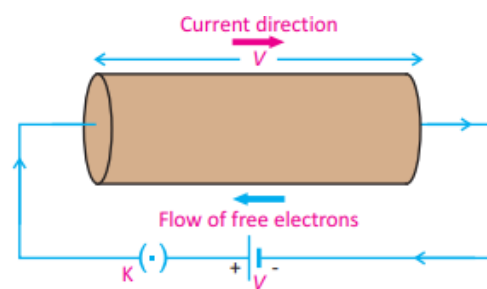
Fig 16.2 Current flows in a conductor when it is connected to a battery

16.1.2 The Measurement of Current

How can we come to know that current has been established in the conductor? For this purpose, we use different electrical instruments which detect the current in the circuit. Galvanometer and ammeter are some common examples of current measuring instruments.

Galvanometer is a very sensitive instrument and can detect small current in a circuit (Fig.16.3). A current of few milliamperes is sufficient to cause full scale deflection in it. While making the connections polarity of the terminals of the galvanometer should be taken into consideration. Generally, the terminal of the galvanometer with red color shows the positive polarity while that of with black colour shows the negative polarity. An ideal galvanometer should have very small resistance to pass the maximum current in the circuit. After suitable modification, the galvanometer can be converted into an ammeter (Fig. 16.4). A large current of the range such as 1 A or 10 A can be measured through the ammeter. Like galvanometer, ammeter is also connected in series, so the current flowing in the circuit also passes through the ammeter (Fig.16.5).

Fig. 16.3 A Galvanometer   Fig. 16.4: An Ammeter
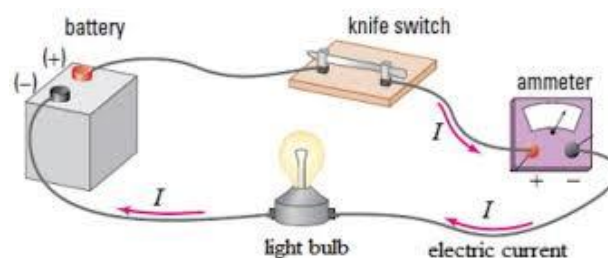
Fig. 16.5 Schematic diagram showing the measurement of current.

16.2 POTENTIAL DIFFERENCE:

When one end A of a conductor is connected to the positive terminal and its other end B is connected to the negative terminal of the battery, then the potential at A becomes higher than the potential at B (Fig.16.6).
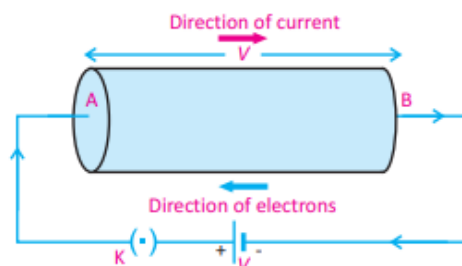
Fig 16.6

This causes a potential difference between the two ends of the conductor. The flow of current continues as long as there is a potential difference. The agency which provides the potential difference for the steady flow of current in the copper wire is the battery. As the current flows from higher potential to the lower potential through the conductor, the electrical energy (due to current) is converted into other forms (heat and light etc.).

When current flows through the conductor, it experiences a resistance in the conductor by collisions with atoms of the conductor. The energy supplied by the battery is utilized in overcoming this resistance and is dissipated as heat and other forms of energy. The dissipation of this energy is accounted for by the potential difference across the two ends of the light bulb.

Thus, Potential difference across the two ends of a conductor causes the dissipation of electrical energy into other forms of energy as charges flow through the circuit. SI unit of potential difference is volt. A potential difference of 1 V across a bulb means that each coulomb of charge or 1 ampere of current that passes through the bulb consumes 1 joule of energy. When a bulb is lit, the energy is taken from the current and is transformed into light and heat energy.

16.3 OHM'S LAW:

Activity 16.1:

Take a nichrome wire of about 50 cm length and apply a potential difference of 1.5 V from a battery (Fig.16.7 a). Measure the current flowing through the wire using an ammeter connected to it in series. Also measure the potential difference across the wire using a voltmeter connected across it. Obtain a set of readings for I and V, by increasing the number of cells. Plot a graph between I and V. This will be a straight line (Fig.16.7-b).
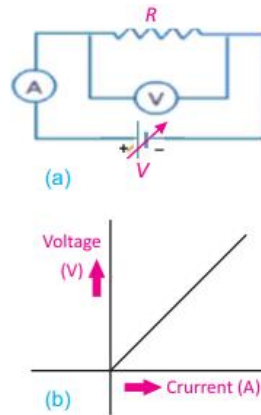
Fig. 16.7

If V is the potential difference across the two ends of any conductor, then current I will flow through it. The value of the current changes with the changes in potential difference and is explained by Ohm's law, stated as:

The amount of current passing through a conductor is directly proportional to the potential difference applied across its ends, provided the temperature and the physical state of the conductor does not change.

$$I \propto V \quad \text{or} \quad V \propto I \quad ... 16.1$$

$$V = IR \quad\quad ... 16.2$$

where R is the constant of proportionality and is the resistance of the conductors. Its SI unit is ohm, denoted by a symbol Ω. If a graph is plotted between the current I and the potential difference V, a straight line will be obtained. Resistance: The property of a substance which offers opposition to the flow of current through it is called its resistance.

This opposition comes from the collisions of moving electrons with atoms of the substance.

16.3.1 Unit of Resistance:
Ohm :

The SI unit of resistance R is ohm. If we put V = 1 V, and I = 1 A, the value of R will be 1. Thus when a potential difference of one volt is applied across the ends of a conductor and one ampere of current passes through it, then its resistance will be one ohm.

16.3.2 FACTORS AFFECTING RESISTANCE:
A short pipe offers less resistance to water flow than a long pipe. Also, the pipe with larger cross sectional area offers less resistance than the pipe having smaller cross sectional area. Same is the case for the resistance of wires that carry current. The resistance of a wire depends both on the cross-sectional area and length of the wire and

on the nature of the material of the wire. Thick wires have less resistance than thin wires. Longer wires have more resistance than short wires. Copper wire has less resistance than steel wire of the same size. Electrical resistance also depends on temperature.

At a certain temperature and for a particular substance

The resistance R of the wire is directly proportional to the length of the wire i.e.

$$R \propto L \qquad \text{.... 16.3}$$

It means, if we double the length of the wire, its resistance will also be doubled, and if its length is halved, its resistance would become one half.

The resistance R of the wire is inversely proportional to the area of cross section A of the wire i.e.,

$$R \propto \frac{1}{A} \qquad \text{.... 16.4}$$

It means that a thick wire would have smaller resistance than a thin wire. After combining the two equations, we get

$$R \propto \frac{L}{A}$$

$$R = \rho \frac{L}{A} \qquad \text{.... 16.5}$$

where 'ρ' is the constant of proportionality, known as specific resistance. Its value depends upon the nature of conductor i.e., copper, iron, tin, and silver would each have a different value of 'ρ'. If we put L = 1 m, and A = 1 m^2 in Eq. (16.5), then R = ρ, i.e., the resistance of one metre cube of a substance is equal to its specific resistance. The unit of 'ρ' is ohm-metre (Ω m).

16.4 CONDUCTORS:

Why do we always use metal wires for conduction of electricity? Because they are good conductors of electricity and offer less resistance to the flow of current. But how can they conduct electricity with much ease? Metals like silver and copper have an excess of free electrons which are not held strongly with atoms of metals. These free electrons move randomly in all directions inside metals. When we apply an external electric field, these electrons can easily move in a specific direction. This movement of free electrons in a particular direction under the influence of an external field causes the flow of current in metal wires. The resistance of conductors increases with increase in temperature. This is due to increase in the number of collisions of electrons with themselves and with the atoms of the metals.

16.5 INSULATORS:

All materials contain electrons. The electrons in insulators, like rubber, however, are not free to move. They are tightly bound inside atoms. Hence, current cannot flow through an insulator because there are no free electrons for the flow of current. Insulators have very large value of resistance. Insulators can be easily charged by friction and the induced charge remains static on their surface. Other examples of insulators are glass, wood, plastic, fur, silk, etc.

16.6 COMBINATION OF RESISTORS:

Resistors can be connected in two ways.

   i.   Series combination
   ii.   Parallel combination.
   i.   **Series Combination:**

In series combination, resistors are connected end to end (Fig. 16.8) and electric current has a single path through the circuit. This means that the current passing through each resistor is the same.
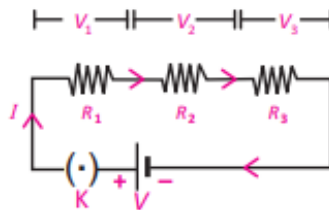
Fig.16.8 : Three resistors in series combination.

Equivalent Resistance of Series Circuit The total voltage in a series circuit divide among the individual resistors so the sum of the voltage across the resistance of each individual resistor is equal to the total voltage supplied by the source. Thus, we can write as

$$V = V_1 + V_2 + V_3 \quad \text{.... 16.6}$$

where V is the voltage across the battery, and V , V , V are the 1 2 3 voltages across resistors R1 , R2 and R3 respectively. If I is the current passing through each resistor, then from Ohm's law

$$V = IR_1 + IR_2 + IR_3$$
$$V = I(R_1 + R_2 + R_3) \quad \text{.... 16.7}$$

We can replace the combination of resistors with a single resistor called the equivalent resistance R such that the e same current passes through the circuit. From Ohm's law

$$V = IR_e$$

Thus, Eq. (16.7) becomes,

$$IR_e = I(R_1 + R_2 + R_3)$$
$$R_e = R_1 + R_2 + R_3 \quad \text{.... 16.8}$$

Thus, the equivalent resistance of a series combination is equal to the sum of the individual resistances of the combination. If resistances R , R , R …….., R are connected in series, then 1 2 3 n , the equivalent resistance of the combination will be given by

$$R_e = R_1 + R_2 + R_3 + \ldots\ldots + R_n$$

ii. **Parallel Combination:**

In parallel combination one end of each resistor is connected with the positive terminal of the battery while the other end of each resistor is connected with the negative terminal of the battery (Fig.14.13). Therefore, the voltage is same across each resistor which is equal to the voltage of the battery i.e.,
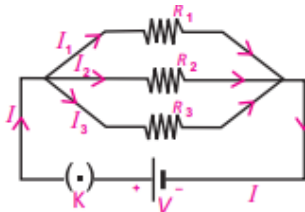
Fig.16.9: Three resistors in parallel combination.

$$V = V_1 = V_2 = V_3$$

Equivalent Resistance of Parallel Circuit

In parallel circuit, the total current is equal the sum of the currents in various resistances i.e.

$$I = I_1 + I_2 + I_3 \quad \text{.... 16.9}$$

Since the voltage across each resistance is V, so by Ohm's law,

$$I_1 = \frac{V}{R_1}, \, I_2 = \frac{V}{R_2} \text{ and } I_3 = \frac{V}{R_3}$$

Thus, Eq.16.9 becomes,

$$I = \frac{V}{R_1} + \frac{V}{R_2} + \frac{V}{R_3}$$

$$I = V\left(\frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3}\right) \qquad \text{.... 16.10}$$

We can replace the combination of resistors with a single resistor called the equivalent resistance R such that the e same current passes through the circuit. From Ohm's law I = V/R Thus, Eq. 16.10 becomes

$$\frac{V}{R_e} = V\left[\frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3}\right]$$

$$\frac{1}{R_e} = \frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} \qquad \text{.... 16.11}$$

Thus, the reciprocal of equivalent resistance of a parallel combination is sum of the reciprocals of the individual resistances, which is less than the smallest resistance of the combination. If resistances R_1 , R_2 , R_3 ...., R_n are connected in parallel, then the equivalent resistance of the combination will be given by

$$\frac{1}{R_e} = \frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} + \ldots\ldots + \frac{1}{R_n}$$

Parallel circuits have two big advantages over series circuits.

1. Each device in the circuit receives the full battery voltage.

2. Each device in the circuit may be turned off independently without stopping the current flowing to the other devices in the circuit. This principle is used in household wiring.

16.7 ELECTRICAL ENERGY AND JOULE'S LAW:

Turbine runs generator to produce electrical energy when water falls on it from higher gravitational potential to lower gravitational potential. Similarly, when charge moves from a higher electric potential to a lower potential, it delivers electric current. Thus, the process during which charges continuously move from a higher potential to a lower potential, becomes a continuous source of electrical energy. Consider two points with a potential difference of V volts. If one coulomb of charge passes between these points; the amount of energy delivered by the charge would be joule. V Hence, when Q coulomb of charge flows between these two points, then we will get QV joules of energy. If we represent this energy by W, then,

Electrical energy supplied by Q charge W = QV joules Now current, when charges Q flow in time t, is defined as:

$$I = \frac{Q}{t}$$
$$\text{or} \quad Q = It$$

So, the energy supplied by Q charge in t seconds = W = V x I x t This electrical energy can be converted into heat and other forms in the circuit.

From Ohm's law, we have $V = IR$
So the energy supplied by Q charge is $W = I^2Rt = \dfrac{V^2t}{R}$
This equation is called Joule's law, stated as:

The amount of heat generated in a resistance due to flow of charges is equal to the product of square of current I, resistance R and the time duration t.

This energy can be utilized for different useful purposes. For example, bulb converts this energy into light and heat, heater and iron into heat, and fans into mechanical energy. Usually, this energy appears as heat in the resistance. This is the reason that we get heat when current passes through a heater.

16.8 ELECTRIC POWER:

The amount of energy supplied by current in unit time is known as electric power.

Hence power P can be determined by the formula Electric power P = electrical energy/time = W/t where W is the electrical energy given by

$$W = QV$$
Therefore, above equation becomes

Electric power $\quad P = \dfrac{QV}{t} = IV = I^2R$

When current I is passing through a resistor R, the electric 2 power that generates heat in the resistance is given by I R. The unit of electric power is watt which is equal to one joule per second (1 Js). It is represented by the symbol W. Electric -1 bulbs commonly used in houses consume 25 W, 40 W, 60 W, 75 W and 100 W of electric power.

Kilowatt-Hour

Electric energy is commonly consumed in very large quantities for the measurement of which joule is a very small unit. Hence, a very large unit of electric energy is needed which is called the kilowatt-hour. It is defined as

The amount of energy delivered by a power of one kilowatt in one hour is called kilowatt-hour.

$$\text{One kilowatt-hour 1 kWh} = 1000\,\text{W} \times 1\,\text{hour}$$
$$= 1000\,\text{W} \times (3600\,\text{s})$$
$$= 36 \times 10^5\,\text{J} = 3.6\,\text{MJ}$$

The energy in kilowatt-hour can be obtained by the following formula: The amount of energy in kilowatt-hour,

$$= \frac{watt \; x \; time \; of \; use \; in \; hours}{1000}$$

The electric meter installed in our houses measures the consumption of electric energy in the units of kilowatt-hour according to which we pay our electricity bills. If the cost of one kilowatt-hour i.e., one unit is known, we can calculate the amount of electricity bill by the following formula:

Cost of electricity = number of units consumed × cost of one unit

$$= \frac{watt \; x \; time \; of \; use \; in \; hours}{1000} \; x \; cost \; of \; one \; unit$$

EXERCISE

1. Can current flow in a circuit without potential difference?
2. Two points on an object are of different electric potential. Does charge necessarily flow between them?
3. In order to measure current in a circuit why ammeter is always connected in series?
4. In order to measure voltage in a circuit voltmeter is always connected in parallel.
5. Discuss. Why in conductors charge is transferred by free electrons rather than by positive charges?
6. How many watt-hours are there in 1000 joules?
7. From your experience of watching cars on the roads at night, are automobile headlamps connected in series or in parallel?
8. A current of 3 mA is flowing through a wire for 1 minute. What is the charge flowing through the wire?   Ans. (180×10^{-3} C)
9. At 100,000 Ω, how much current flows through your body if you touch the terminals of a 12 V battery? If your skin is wet, so that your resistance is only 1000 Ω, how much current would you receive from the same battery?
    Ans. (1.2×10^{-4} A, 1.2×10^{-2} A)
10. The resistance of a conductor wire is 10 MΩ. If a potential difference of 100 volts is applied across its ends, then find the value of current passing through it in mA.
    Ans. (0.01 mA)
11. By applying a potential difference of 10 V across a conductor, a current of 1.5 A passes through it. How much energy would be obtained from the current in 2 minutes?   Ans.(1800 J)
12. Two resistances of 2 kΩ and 8 kΩ are joined in series, if a 10 V battery is connected across the ends of this combination, find the following quantities:
    (a) The equivalent resistance of the series combination.
    (b) Current passing through each of the resistances.
    (c) The potential difference across each resistance.
    Ans. [(a) 10 kΩ (b) 1 mA (c) 2 V, 8 V]
13. Two resistances of 6 k and 12 kΩ are connected in parallel. A 6 V battery is Ω connected across its ends, find the values of the following quantities:
    (a) Equivalent resistance of the parallel combination.
    (b) Current passing through each of the resistances.
    (c) Potential difference across each of the resistance.
    Ans. [(a) 4 kΩ, (b) 1 mA,0.5 mA (c) 6 V]
14. An electric bulb is marked with 220 V, 100 W. Find the resistance of the filament of the bulb. If the bulb is used 5 hours daily, find the energy in kilowatt-hour consumed by the bulb in one month (30 days).   Ans. (484 , 15 Ω kWh)

Chapter # 17

Electromagnetic Radiation

Students will be able to understand:

- Spectrum
- ionization excitation
- Inverse Square law
- Frequency
- wave length

The **electromagnetic spectrum,** in simple terms, is defined as the range of all types of electromagnetic radiation. We shall learn about the concept in detail and understand all its underlying aspects in this lesson.

the electromagnetic spectrum is a range of frequencies, wavelengths, and photon energies covering frequencies from below 1 hertz to above 10^{25} Hz, corresponding to wavelengths which are a few kilometers to a fraction of the size of an atomic nucleus in the spectrum of electromagnetic waves. Generally, in a vacuum, electromagnetic waves tend to travel at speeds which are similar to that of light. However, they do so at a wide range of wavelengths, frequencies, and photon energies.

The electromagnetic spectrum consists of a span of all electromagnetic radiation which further contains many sub ranges, which are commonly referred to as portions. These can be further classified as infrared radiation, visible light or ultraviolet radiation.

17.1 Electromagnetic Waves in the Electromagnetic Spectrum:

The entire range (electromagnetic spectrum) is given by radio waves, microwaves, infrared radiation, visible light, ultra-violet radiation, X-rays, gamma rays and cosmic rays in the increasing order of frequency and decreasing order of wavelength. The type of radiation and their frequency and wavelength ranges are as follows:

Type of Radiation	Frequency Range (Hz)	Wavelength Range
Gamma-rays	$10^{20} - 10^{24}$	$< 10^{-12}$ m
X-rays	$10^{17} - 10^{20}$	1 nm – 1 pm
Ultraviolet	$10^{15} - 10^{17}$	400 nm – 1 nm
Visible	$4 \times 10^{14} - 7.5 \times 10^{14}$	750 nm – 400 nm
Near-infrared	$1 \times 10^{14} - 4 \times 10^{14}$	2.5 µm – 750 nm
Infrared	$10^{13} - 10^{14}$	25 µm – 2.5 µm
Microwaves	$3 \times 10^{11} - 10^{13}$	1 mm – 25 µm
Radio waves	$< 3 \times 10^{11}$	> 1 mm

Fig.17.1 Frequency and wavelength ranges for different types of radiation.

The electromagnetic spectrum can be depicted as follows:
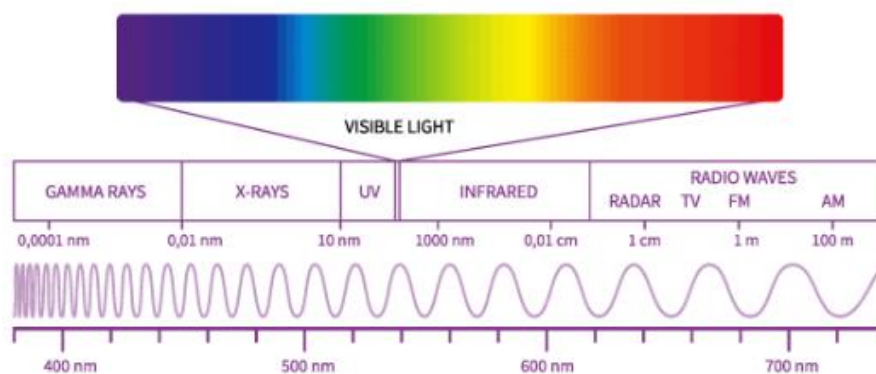
17.2 Uses of electromagnetic waves in our daily life:

Radio: A radio basically captures radio waves that are transmitted by radio stations. Radio waves can also be emitted by gases and stars in space. Radio waves are mainly used for TV/mobile communication.

Microwave: This type of radiation is found in microwaves and helps in cooking at home/office. It is also used by astronomers to determine and understand the structure of nearby galaxies and stars.

Infrared: It is used widely in night vision goggles. These devices can read and capture the infrared light emitted by our skin and objects with heat. In space, infrared light helps to map interstellar dust.

X-ray: X-rays can be used in many instances. For example, a doctor can use an X-ray machine to take an image of our bones or teeth. Airport security personnel use it to see through and check bags. X-rays are also given out by hot gases in the universe.

Gamma-ray: It has a wide application in the medical field. Gamma-ray imaging is used to see inside our bodies. Interestingly, the universe is the biggest gamma-ray generator of all.

Ultraviolet: The Sun is the main source of ultraviolet radiation. It causes skin tanning and burns. Hot materials that are in space also emit UV radiation.

Visible: Visible light can be detected by our eyes. Light bulbs, stars, etc., emit visible light.

Spectroscopy: Spectroscopy is used to study the way different electromagnetic waves interact with matter.

Significance of the Electromagnetic Spectrum

The electromagnetic waves in these different bands have different characteristics depending upon how they are produced, how they interact with matter and their practical applications. Maxwell's equations predicted the existence of an infinite number of frequencies of electromagnetic waves, all traveling with the speed of light. This is the first indication of the existence of the entire electromagnetic spectrum.

Nonetheless, the main significance of the electromagnetic spectrum is that it can be used to classify electromagnetic waves and arrange them according to their different frequencies or wavelengths.

17.3 Practical Applications of Electromagnetic Waves:

● Radio waves and microwaves discovered by Hertz paved the way for wireless television, radio and mobile communication.

● Visible light portion of the electromagnetic spectrum is the reason for all visual aids in daily life. This is the portion of the electromagnetic spectrum that helps us to see all objects, including colours.

● The X-rays discovered by Roentgen proved to be useful in medicine for detecting many ailments or deformities in bones.

● High ultraviolet radiation has energies to ionise the atoms causing chemical reactions.

● Gamma rays discovered by Paul Villard are useful for ionization purposes and nuclear medicine.

In physics, many laws define objects, light, Earth, and many more. These laws are formed for some fundamental quantity. Snell's law, laws of reflection, laws of refraction are some of the laws which are formed on the nature of light. We know that light travels in a straight line and can undergo various phenomena like diffraction, polarization, interference, reflection, and refraction too.

17.4 Excitation:

If all the electrons are in the lowest energy level, we say the atom is in its ground state.If the electrons in an atom absorb energy and move to higher energy states, we say the atom is excited.

Excitation only occurs if the electron absorbs exactly the right amount of energy to move to higher energy levels.

Excitation will occur in the following circumstances:

i. An electron absorbs a photon with exactly the right amount of energy to move it between two levels (so a photon will not be absorbed if the energy it provides to the electron is not enough for it to move between two levels)
ii. An electron absorbs exactly the right amount of energy to move between two levels after colliding with a free electron (that contains the same energy or greater energy than that required for the electron to change energy levels).
iii. As energy is conserved, the energy gained by the electron in the atom equals the energy lost by the free electron
iv. The free electron's kinetic energy after the collision is the same as before the collision minus the energy transferred to the excited electron in the atom.

17.5 Ionization:

If an electron in an atom absorbs enough energy to leave the atom completely, the atom is said to be ionized.

By gaining or losing an electron, it becomes a charged particle (i.e. an ion)

The ionization energy is the minimum energy needed to remove the electron from the atom completely,
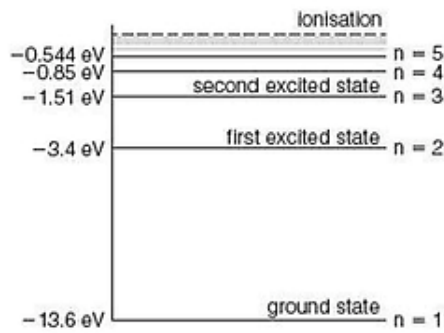
This shows a number of possible excited states for a hydrogen atom.

17.6 Inverse square law:

This law explains the strength of light with respect to the distance of the source.

Inverse square law states that:
"The intensity of the radiation is inversely proportional to the square of the distance".

OR

"The intensity of the light to an observer from a source is inversely proportional to the square of the distance from the observer to the source".

As per this law, light loses its brightness or luminosity as it moves away from the source. For example: when you switch on the light in one corner of the room and when you move away from the source, the light appears dim or less bright due to an increase in the distance (away from the source).

The below figure explains the source of light and the source distance relationship.
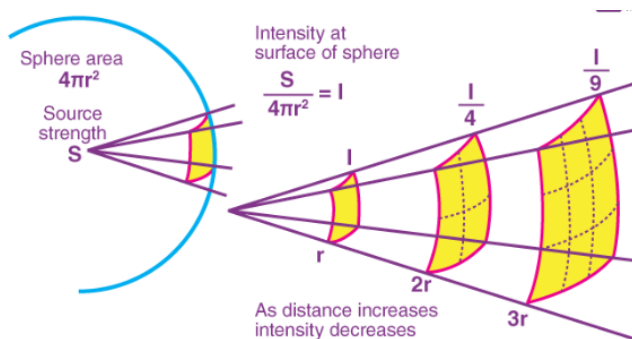
Consider the distance as 2, and when the distance is squared, it results in 4. The inverse of 4 would be 1/4.

So when we compare, we can note that the resultant is a quarter the original power.

Hence light intensity is $1/d^2$

If the distance between an object and the light source is doubled, a given area experiences one-fourth of the light from the source.

When the distance from the source is tripled, the given area experiences one-ninth part of the light.

The Formula of Inverse-Square Law

The formula of inverse-square law is given as,

$$I \alpha \frac{1}{d^2}$$

Where,

d is the distance

I is radiation intensity

Consider light sources of intensity I1 and I2 at the distances d1 and d2. The inverse-square law is articulated as:

$$\frac{I_1}{I_2} \alpha \frac{d_2^2}{d_1^2}$$

Where the intensity of light is measured by candela or Lumen, and the distance is measured in meters.

17.6.1 Applications Of Inverse Square Law:

i.   This law is used to calculate the intensity of any given radiation or distance.
ii. Inverse-square law helps to calculate the source to film distances in X-ray techniques.
iii. It also helps to determine the time of C-ray exposure and the intensity of the X-ray tube used in the process.
iv. Using the standard candle approach, when the brightness of the source is known, it helps to calculate the distance from the Earth.
v.   Inverse-square law is used to measure various astronomical distances.

17.7 Frequency

By the definition of frequency, we can understand that if a body is in periodic motion, it has undergone one cycle after passing through a series of events or positions and returning to its original state. Thus, frequency is a parameter that describes the rate of oscillation and vibration.

"Number of waves passing through a point per unit time t."
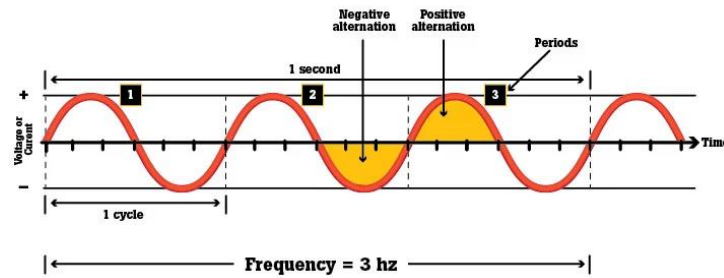
The equation gives the relation between the frequency and the period:

The relation between the frequency and the period is given by the equation:

f=1/T

For a sinusoidal wave represented by the equation:

y (0,t) = -a sin (ωt)

The formula of the frequency with the SI unit is given as:

Formula	$f = \frac{1}{T} = \frac{\omega}{2\pi}$
SI unit	Hertz

17.8 WAVELENGTH

The wavelength of light is defined as **"The distance between the two successive crests or troughs of the light wave".** It is denoted by the Greek letter lambda (λ). Therefore, the distance between either one crest or trough of one wave and the next wave is known as wavelength.

The visible spectrum is nothing but the observable region of the electromagnetic wave which is visible to human eyes. In the electromagnetic spectrum, the visible spectrum ranges from the infrared region to the UV region. The visible light lies in between the infrared and ultraviolet range of wavelengths. The human eye can detect the light spectrum ranging from 400 nanometers (violet) to about 700 nanometers (red). Other electromagnetic radiations are either too small or too large to capture for the human eye and are out of biological limitations.

We can see these waves as the colours of the rainbow where each colour includes a different wavelength. In the visible light, we can also see the sun's outermost layer – the corona.

White Light

Glass Prism
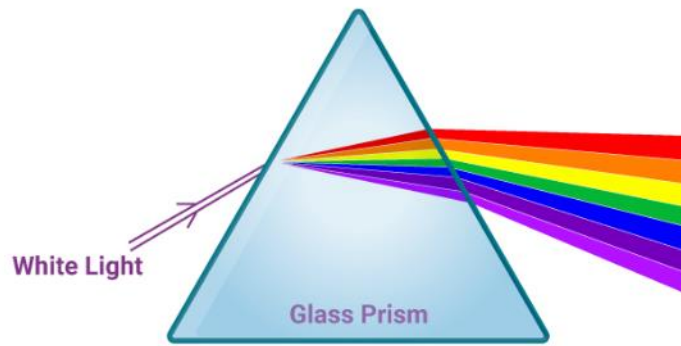
When the visible light travels through a prism, the visible light gets separated into a spectrum of colours. Red colour has the longest wavelength of 700 nm, and violet has the shortest wavelength of 380 nm. These colours arrange themselves according to the wavelength as the spectrum of rainbow colours.

As light has the properties of a wave and a particle, it can be expressed in two equations:

$$v = \lambda f$$
$$E = hf$$

Where,

v is the velocity of the light.

λ is the wavelength of the light.

f is the frequency of the light.

E is the energy of the light wave.

h is the Planck's Constant (6.64 × 10^{-34} joule. Second)

RELATION BETWEEN VELOCITY, FREQUENCY AND WAVELENGTH

Wave is a disturbance in a medium which travels from one place to another and hence has a specific velocity of travelling. This is called the velocity of wave,

which is defined by:

$$Velocity = distance/time$$
$$v = \frac{d}{t}$$

If time taken by the wave in moving from one point to another is equal to its time period T, then the distance covered by the wave will be equal to one wavelength, hence we can write:

$$v = \frac{\lambda}{T}$$

But period T, is reciprocal of the frequency.

$$T = \frac{1}{f}$$

Therefore,

$$v = f\lambda$$

Practical Physics

1. To find the unknown force.

Material Required:

Boxes A, B, and C are stacked on the ground as shown. Boxes B and C both weigh 5N and box A weighs 10N. What is the force of box A on box B?
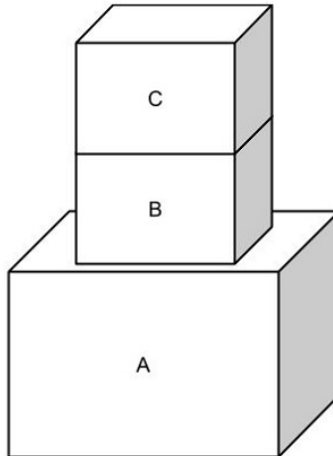
Procedure:

i.   Identify the magnitude and direction of the known forces within the system.
ii.   The weight of both boxes B and C add to a force of 5N+5N=10N which weighs on box A.
iii. Use Newton's Third Law to find the unknown force within a force pair.
iv.   The force of box A on box B is equal and opposite to the 10N force of boxes B and C on box A. Therefore, the force of box A on box B is 10N upward.

Result:

………
………
………

2. To find the center of gravity of an irregular shape.

Material Required:

Given the irregular shape, pencil, and measuring scale.
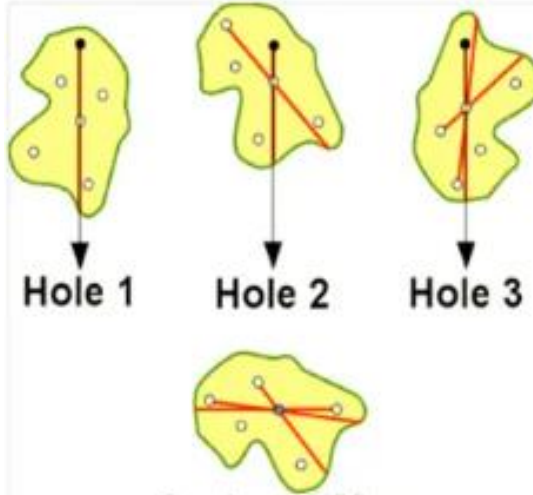
Procedure:

   i.   Hang the items up.
  ii.   Using the same location, suspend a plumb line.
 iii.   Make a mark where the thread is.
 iv.   The center of mass is located along the thread's line.
  v.   Follow the same procedure above with various hanging objects.
 vi.   These lines meet at the center of mass.

Result:

----

3. To verify the law of reflection.

Material Required:

Wooden board, white paper, Measuring scale, pencil, pins Mirror
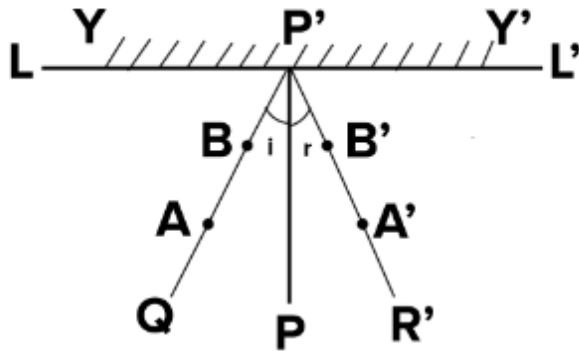
Procedure:

i. First, take a wooden board, then fix a white paper on it.
ii. Draw a straight line in the middle of the paper and mark it as LL'
iii. Then draw a perpendicular to LL' and name it as LL'
iv. Now, place a mirror YY'on the line LL'in a way that the polished side of the mirror is along the line.
v. Fix two pins A and B on the line QP' at least 10cm apart.
vi. Look in the mirror to find the images of pins A and B , two images A' and B' will be shown, fix pins there also.
vii. Join the points A' and B' and a straight line P'R will be formed.
viii. Remove all the pins and mark the points with pen or pencil.
ix. Measure the ∠QP'P = ∠i ∠ PP'R = ∠r. It is found that ∠i = ∠r . Hence, proving the first law of reflection true.
x. Also, the incident ray reflected ray and the normal lie in the plane of the paper, so, they lie in the same plane, Hence, proving the second law of reflection true.

Result:

To find the path of light passing through a prism.

Material Required:

A white sheet, soft board, Thumb pins, 4-6 all pins, Prism, Pencil, Scale, Protractor, Drawing board.
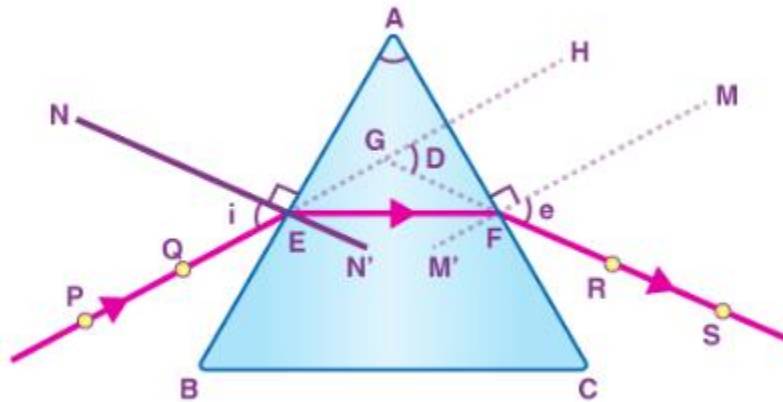
PE - Incident ray
EF - Refracted ray
FS - Emergent ray
A - Angle of the prism
∠i - Angle of incidence
∠r - Angle of refraction
∠e - Angle of emergence
∠D - Angle of deviation

Procedure:

i.   Fix a white sheet on a drawing board using drawing pins.
ii.   Place the triangular prism resting on its triangular base. Using a pencil, draw the outline of the prism.
iii.   Draw NEN normal to the face of the prism AB. Make an angle between 30° and 60° with the normal.
iv.   On the line PE, fix two pins at a distance of 5cm from each other and mark these as P and Q.
v.   Look for the images of the pins at P and Q through the other face of the prism AC.
vi.   Fix two pins at R and S such that they appear as a straight line as that of the P and Q when it is viewed from the AC face of the prism.
vii.   Remove the pins and the prism.
viii. At point F, make the points R and S meet by extending them.
ix.   PQE is the incident ray which is extended till it meets face AC. SRF is the emergent ray which is extended backward to meet at point G.

x.   Now mark the angle of incidence ∠i, angle of refraction ∠r and the angle of emergence ∠e and the angle of deviation ∠D as shown in the experimental setup.

xi.   Repeat the experiment for more angles between 30° and 60°.

xii.   Observations:

xiii.   At surface AB, the light ray enters and bends towards the normal on refraction.

xiv.   At surface AC, the light ray bends away from the normal as it travels from one medium (glass) to the other (air).

xv.   The angle of deviation is observed. Here, the emergent ray bends at an angle from the direction of the incident ray.

Result:

Precautions:

i.   For drawing the boundary of the prism, a sharp pencil should be used.

ii.   Soft board and pointed pins should be used.

iii.   The distance between the pins should be 5cm or more.

iv.   The pins should be fixed vertically and should be encircled when they are removed from the board.

v.   The angle of incidence should be between 30° and 60°.

vi.   The arrows drawn for incident ray, reflected ray and emergent ray should be proper.

vii.   For viewing the collinearity of all four pins and images, the head should be slightly tilted on either side. While doing this it can appear as if all are moving together.

3. Determine the critical angle of glass using a glass prism.

Material Required:

Glass prism, Wooden board, paper, pencil, measuring scale, common pins and drawing pins.
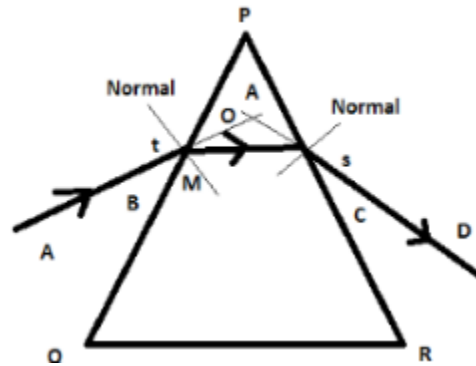
Procedure:

i.   Place the glass prism on the drawing board and draw a line OR on the paper.
ii. Place the paper on the drawing board and fix it with the help of drawing pins.
iii. Draw a normal to the surface of the prism at point M.
iv. Place the prism such that the incident ray falls on the surface of the prism at point and emerges from the other surface at point S.
v.   Rotate the prism until the refracted ray just grazes the surface of the prism at point S.
vi. Measure the angle of incidence i and the angle of refraction r.
vii. Calculate the refractive index of the material using the formula:

$$n = \frac{\sin i}{\sin r}$$

Result:

--.

4. Determine the focal length of a convex lens.

Material Required:

A wooden bench, A convex lens, A lens holder, A screen fixed to a stand, A measuring scale.

Determination of focal length of a thin convex lens
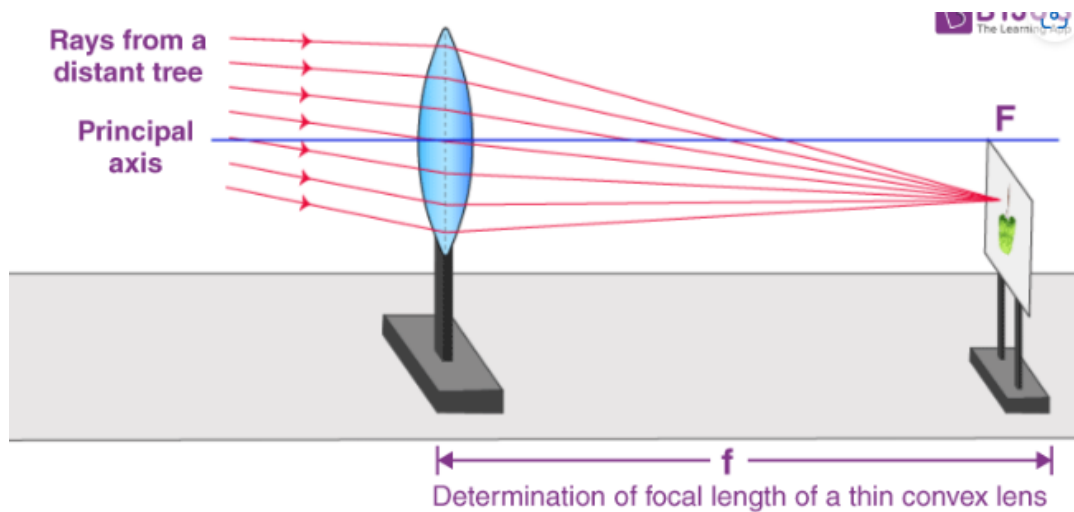
Procedure:

i. Without disturbing the lens and screen, arrange both of them on the wooden bench.
ii. Place the lens on the holder facing a distant object.
iii. Place the holder with the screen on the bench.
iv. The position of the screen should be such that the sharp image of the distant object is obtained on it.
v. The difference between the position of the lens and the screen is equal to the focal length of the given convex lens.
vi. Now shift the focus to towards various other distant object and calculate the focal length of the convex lens.

Observation Table

Sl.no	Position of convex lens (L)	Position of screen (S)	Focal length = (L-S) cm
1	60 cm	50 cm	f_1 = 10cm
2	60 cm	50 cm	f_2 = 10cm
3	60 cm	50 cm	f_3 = 10cm

Calculation:

Following is the mean value of the focal length of the convex lens:

$$\frac{f_1+f_2+f_3}{3} \ cm = 10 \ cm$$

Result:

10 cm is the focal length of the convex lens.

Precautions:

i.   The placement of the convex lens should be vertical.
ii. There should not be any hurdle between the rays of light from the object and the convex lens.
iii. To get a well-illuminated and distinct image of the distinct object, the distant object should be well-illuminated.
iv. The convex lens stand, and the screen should be parallel to the measuring scale.

5.To find the focal point of a lens.

Material Required:

Lens, Paper, Pencil, Wooden Board.

Procedure:

The formula for a parabola is $f = x^2/4a$. To find the focal point of a parabola, follow these steps:

i. Measure the longest diameter (width) of the parabola at its rim.
ii. Divide the diameter by two to determine the radius (x) and square the result (x^2).
iii. Measure the depth of the parabola (a) at its vertex and multiply it by 4 ($4a$).
iv. Divide the answer from Step 2 by the answer to Step 3 ($x^2/4a$). The answer is the distance from the vertex of the parabola to its focal point.

Source: Humboldt State University
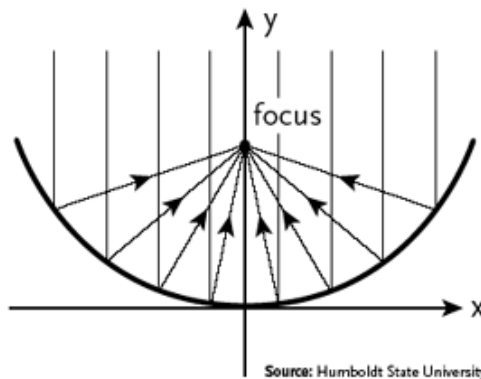
The focal point is the point at which light waves traveling parallel to the axis of the parabola meet after reflecting off its surface.

Result:

6. To find the reflective index of a liquid using a concave mirror.

Material Required:

A concave spherical mirror, water, an optical needle, a clamp stand, one meter scale, plumb line, etc.
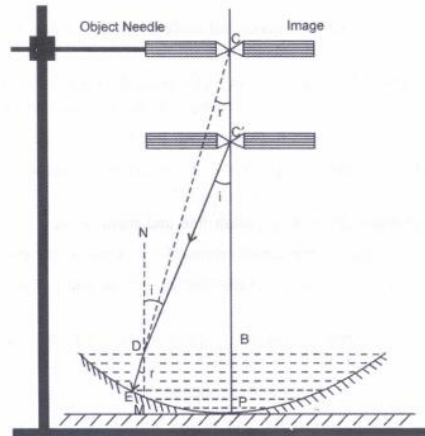
Fig. 1 : Refractive index of liquid

PROCEDURE:

i.   Place the concave mirror on a horizontal surface (plane) so that its principal axis is vertically along.
ii. Hold the optical needle horizontally in a clamp stand so that its tip lies just above the pole 'P' and at a distance equal to 2f (f is focal length) as shown in Figure 1.
iii. Remove the parallax between the needle and its image.
iv. Mark the real and inverted image of the optical needle in the mirror. Note the reading of this image.
v.   Measure the distance (PC) using a plumb line and meter scale.
vi. This measured distance is the actual radius of curvature of the concave mirror.
vii. Now add a small amount of water to the concave mirror which will change the position of the image needle.
viii. Adjust the screw to upper or lower the optical needle and adjust its position from 'C' to 'C'' to remove parallax between the needle and its shifted image is seen.
ix. Measure the distance (PC') which is the apparent radius of curvature of the concave mirror.
x.   Repeatedly experiment four to five times and record the readings.

OBSERVATIONS:

The rough focal length of the concave mirror, f = ----------cm

Table for Refractive index:

No of observations	Actual Radius of curvature (R) cm	Apparent radius of curvature(R') cm	$\mu = \dfrac{R}{R'}$	Mean μ
1				
2				
3				
4				

CALCULATIONS:

Putting the observed values of actual radius R and apparent radius R' in

$$\mu = \frac{R}{R'}$$

and calculate the value of refractive index of the given medium.

RESULT:

The refractive index of water = ------------

PRECAUTIONS:

The mirror and surface of the water should be very clean.

The needle should be clamped properly in horizontal position.

The parallax between the optical needle and its image should be removed properly.

7. Determine the speed of sound in air at room temperature.

Material Required:

Resonance tube, Rubber Pad, Thermometer, Set Squares, Plumb Line, Water in a beaker, Two-timing forks of known frequency.
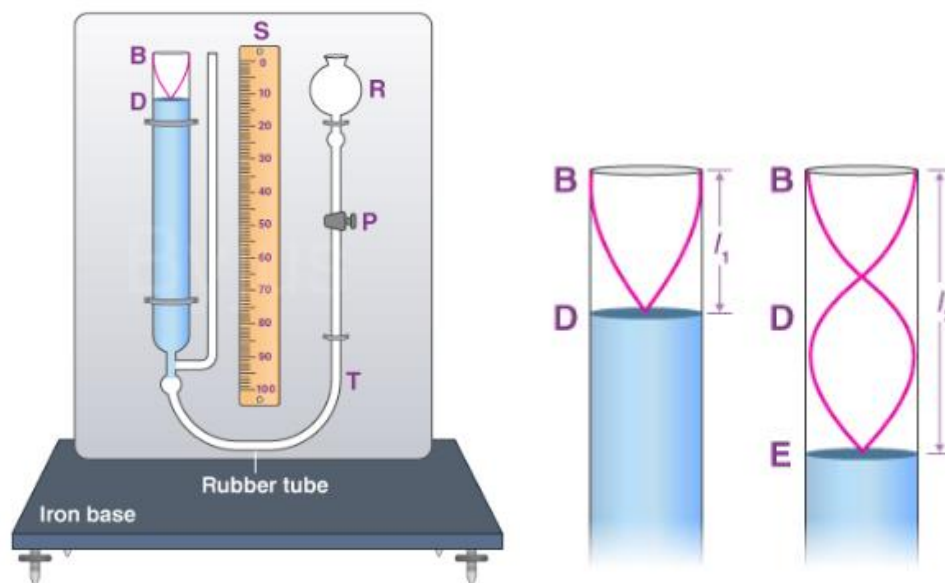
Procedure

i.   By making base horizontal with the help of levelling screws, set the resonance tube vertical.
ii. Fix the reservoir R in the uppermost position.
iii. Loosen the pinch cock P and fill the reservoir and metallic tube completely with water by a beaker.
iv. Tighten the pinch cock, lower the reservoir, and fix it in the lowest position.
v.   Take a tuning fork of higher frequency.

Observations:

The temperature of the air in the air column:

(i) in the beginning _____ °C

(ii) at the end _____°C

The mean temperature is calculated as follows:

$$t = \frac{t_1 + t_2}{2}$$

Frequency of first tuning fork = f_1

Frequency of second tuning fork = f_2

Frequency of tuning fork	Resonance	No. of observation	Position of water level at resonance			Mean Length			
			Water level falling (cm)	Water level rising (cm)	Mean (cm)				
f_1	First	1 2 3					Second	1 2 3	
	f_2	First	1 2 3					Second	1 2 3

Calculation

From the first tuning fork,

$$v_1 = 2f_1(l_2\prime - l_1\prime)$$

From the second tuning fork,

$$v_2 = 2f_2(l_2\prime\prime - l_1\prime\prime)$$

The mean velocity at room temperature is given as follows:

$$v = \frac{v_1 + v_2}{2}$$

Result:

At room temperature, the velocity of sound in air is _____ m/s.

CHEMISTRY

Chapter # 1

Composition of Substance

STUDENTS LEARNING OUTCOMES:

- Distinguish between matter and substance.
- Differentiate between atom, element, and compound.
- Able to identify the types of molecules.
- Differentiate between molecular and empirical formula.
- Identify the representative particles of elements and compounds.

1.1 Matter:

"Matter is anything that takes up space and can be weighed".

In other words, matter has volume and mass. There are many different substances, or types of matter, in the universe.

1.2 Substance:

"A particular kind of matter with uniform properties and composition is called substance". i.e. sugar as well as iron.

1.3 Atom:

Greek philosophers thought that matter is made up of indivisible particles which can't be further subdivided.

Dermocritus called these indivisible particles as atoms (the Greek word atomos means indivisible).

In 1808, an English school teacher John Dalton gave atomic theory. Its main parts are given below:

All matter is composed of atoms.

The atoms of the same elements have similar properties.

The atoms of different elements have different properties.

1.3.1 Modern definition of Atom

Atom is the smallest particle of an element which may or may not exist independently and take part in a chemical reaction. e.g. He, Ne, H, C, Na, Cu, etc.

1.3.2 Sub-atomic Particles

Atom can be further divided into smaller particles called sub-atomic particles like electron, proton, neutron, neutrino, etc. More than 100 sub-atomic particles have been discovered but electron, proton, and neutron are the fundamental particles of an atom.

Berzelius (1779-1848) determined the atomic weight of elements. Moreover, he introduced the symbols for elements.

1.3.3 Evidence of atoms

Atoms are very small in size, and it is not possible to see the atoms by optical microscope.

a) Ordinary Microscope

It can measure the size of an object upto or Above 500nm($1nm=10^{-9}m$) as it uses visible light. A clear *and* accurate image of an object that has a smaller wavelength than visible light can't be obtained.

b) Electron Microscope

It uses a beam of electron instead of visible light because the wavelength of electron is much smaller than that of visible light. Fig.1.1 shows electron microscope photograph of a piece of graphite that has been magnified about 15 million times. The bright bands in the figure are layers of carbon atoms.

Fig. 1.1: Electron Microscope photograph of Graphite

1.3.4 Size and Mass of an Atom

In the twentieth century, X-ray work has shown that:

The diameter of atoms are of order 2×10^{-10}m which is 0.2nm.

Masses of atoms range from 10^{-27} to 10^{-25}kg. Masses of atoms are often expressed in atomic mass unit (amu).

1 amu = 1.66 x10^{-27} kg.

1.4 MOLECULE:

1.4.1 Definition:

The smallest particle of a pure substance which can exist independently is called a molecule. Some of the examples of molecules are H_2, He, Cl_2, H_2O, etc.

1.4.2 Types of Molecules:

Molecules can be divided into various types on the following basis:

i.   On the basis of atomicity
ii.   On the basis of kind (same or different types) of atom
iii. On the basis of size and molecular weight of molecules

i.   On the basis of atomicity:

The number of atoms in a molecule is called its atomicity. On the basis of atomicity,molecules may be:

a) Monoatomic Molecules: These molecules contain only one atom. e.g. He, Ne etc.

b) Diatomic Molecules: These molecules contain two atoms.

e.g. N_2 , O_2 etc.

c) Polyatomic Molecules: These molecules contain more than two atoms e.g., H_2O, NH_3, etc.

ii.   On the basis of kind of atoms:

On the basis of the kind of atoms, molecules may be:

a) Homo-atomic Molecules:

These molecules consist of the same kind of atoms e.g. H_2 , N_2 etc.

b) Hetero-atomic Molecules:

These molecules consist of different kinds of atoms. e.g. HCl , NH_3 etc.

iii.   On the basis of size and molecular weight:

The sizes of the molecules are bigger than atoms. And it depends upon the no. of atoms and their shapes. On the basis of size and weight, molecules may be:

a) Micro-molecules:

Molecules that have small sizes and low molecular weight are known as micromolecules. e.g. O_2, H_2 etc.

b) Macro-molecules:

Molecules that have large sizes and high molecular weights. e.g. Hemoglobin is a macro-molecule. It is found in blood and carries oxygen to all parts of our body. Each molecule of hemoglobin is made up of nearly 10,000 atoms and is 68,000 times heavier than a hydrogen atom.

1.5 Elements, Compounds, and Symbols

In the early ages, only nine elements (carbon, gold, silver, tin, mercury, lead, copper, iron, and sulphur) were known. At that time, it was considered that elements were substances that could not be broken down into simpler units by ordinary chemical processes. By the end of the nineteenth century, sixty-three elements had been discovered.

1.5.1 Naturally occurring elements

Now 118 elements have been discovered, out of which 92 are naturally occurring elements. A modern definition of element is that it is a substance made up of the same type of atoms, having the same atomic number, and cannot be decomposed into simple substances by ordinary chemical means. It means that each element is made up of unique types of atoms that have very specific properties. Elements occur in nature in free or combined form. All the naturally occurring elements found in the world have different percentages in the earth's crust, oceans, and atmosphere.

Table 1.2. shows natural occurrence in percentage by weight of some major elements around us. It shows concentrations of these major elements found in the three main systems of our environment.

Earth's crust		Oceans		Atmosphere	
Oxygen	47%	Oxygen	86%	Nitrogen	78%
Silicon	28%	Hydrogen	11%	Oxygen	21%
Aluminum	7.8%	Chlorine	1.8%	Argon	0.9%

Table 1.1 Natural Occurrences by Weight % of Some Major Elements

1.5.2 State of elements:

Elements may exist in different physical states. i.e. solid , liquid and gas.

Solid : Most of the elements exist as solids e.g., sodium, copper, zinc, gold, etc.

Liquid : There are very few elements that occur in liquid state e.g., mercury and bromine.

Gas : A few elements exist as gases e.g., nitrogen, oxygen, chlorine, and hydrogen.

1.5.3 Classification of elements:

On the basis of their properties, elements are divided into:

  i.   Metals
  ii. non-metals
  iii. metalloids

About 80 percent of the elements are metals.

1.5.4 Symbols:

Elements are represented by symbols, which are abbreviations for the names of elements. A symbol is taken from the name of that element in English, Latin, Greek, or German. If it is one letter, it will be capitalized as H for Hydrogen, N for Nitrogen, and C for Carbon, etc. In the case of two letters symbol, only the first letter is capital e.g., Ca for Calcium, Na for Sodium, and Cl for Chlorine.

1.5.5 Valency:

The unique property of an element is valency. "It is combining capacity of an element with other elements". It depends upon the number of electrons in the outermost shell (valence shell).

In simple covalent compounds, valency is the number of hydrogen atoms that combine with one atom of that element, or the number of bonds formed by one atom of that element e.g., in the following compounds.

The $H-Cl$, $H-O-H$,

$$\begin{array}{c} H \\ | \\ N-H \\ | \\ H \end{array}$$

and

$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

valency of chlorine, oxygen, nitrogen, and carbon is 1, 2, 3, and 4, respectively.

In simple ionic compounds, valency is the number of electrons gained or lost by an atom of an element to complete its octet. Elements having less than four electrons in their valence shell prefer to lose the electrons to complete their octet. For example, atoms of Na, Mg, and Al have 1, 2, and 3 electrons in their valence shells respectively. They lose these electrons to have valency of 1, 2, and 3, respectively. On the other hand, elements

having five or more than five electrons in their valence shells, gain electrons to complete their octet. For example, N, O, and Cl have 5, 6, and 7 electrons in their valence shells respectively. They gain 3, 2, and 1 electron respectively to complete their octet. Hence, they show valency of 3, 2, and 1, respectively. A radical is a group of atoms that have some charge. Valencies of some common elements and radicals are shown in Table 1.3.

Element / Radical	Symbol	Valency	Element / Radical	Symbol	Valency
Sodium	Na	1	Hydrogen	H	1
Potassium	K	1	Chlorine	Cl	1
Silver	Ag	1	Bromine	Br	1
Magnesium	Mg	2	Iodine	I	1
Calcium	Ca	2	Oxygen	O	2
Barium	Ba	2	Sulphur	S	2
Zinc	Zn	2	Nitrogen	N	3
Copper	Cu	1,2	Phosphorus	P	3,5
Mercury	Hg	1,2	Boron	B	3
Iron	Fe	2,3	Arsenic	As	3
Aluminium	Al	3	Carbon	C	4
Chromium	Cr	3	Carbonate	CO_3^{2-}	2
Ammonium	NH_4^+	1	Sulphate	SO_4^{2-}	2
Hydronium	H_3O^+	1	Sulphite	SO_3^{2-}	2
Hydroxide	OH^-	1	Thiosulphate	$S_2O_3^{2-}$	2
Cyanide	CN^-	1	Nitride	N^{3-}	3
Bisulphate	HSO_4^-	1	Phosphate	PO_4^{3-}	3
Bicarbonate	HCO_3^-	1			

Table 1.2 Some Elements and Radicals with their Symbols and Common Valencies

Some elements show more than one valency, i.e., they have variable valency. For example, in ferrous sulphate (Fe_2SO_4) the valency of iron is 2. In ferric sulphate is ($Fe_2(SO_4)_3$), the valency of iron is 3. Generally, the Latin or Greek name for the element 2 4 3 (e.g., Ferrum) is modified to end in 'ous' for the lower valency (e.g. Ferrous) and to end in 'ic' for the higher valency (e.g. Ferric).

1.6 Compound

1.6.1 Definition:

"Compound is a substance made up of two or more elements chemically combined together in a fixed ratio by mass".

As a result of this combination, elements lose their own properties and produce new substances (compounds) that have entirely different properties. Compounds can't be broken down into its constituent elements by simple physical methods.

1.6.2 Examples:

Carbon dioxide is formed when elements of carbon and oxygen combine chemically in a fixed ratio of 12:32 or 3:8 by mass.

Water is a compound formed by a chemical combination between hydrogen and oxygen in a fixed ratio of 1:8 by mass.

1.6.3 Classification:

Compounds can be classified as ionic or covalent compounds.

1.6.3.1 Properties of ionic compounds:

i. Ionic compounds do not exist in independent form.
ii. They form a three-dimensional crystal lattice, in which each ion is surrounded by oppositely charged ions.
iii. These oppositely charged ions attract each other very strongly, as a result, ionic compounds
iv. They have high melting and boiling points.
v. These compounds are represented by formula units e.g., NaCl, KBr, $CuSO_4$.

1.6.3.2 Properties of covalent compounds:

i. The covalent compounds mostly exist in molecular form.
ii. A molecule is a true representative of the covalent compound, and its formula is called molecular formula e.g., H_2O, HCl, H_2SO_4, CH_4 etc.

Table 1.4 Some Common Compounds with their Formulae

Compound	Chemical Formula
Water	H_2O
Sodium chloride (Common salt)	NaCl
Silicon dioxide (Sand)	SiO_2
Sodium hydroxide (Caustic Soda)	NaOH
Sodium carbonate (Washing Soda)	$Na_2CO_3.10H_2O$
Calcium oxide (Quick Lime)	CaO
Calcium carbonate (Limestone)	$CaCO_3$
Sugar	$C_{12}H_{22}O_{11}$
Sulphuric acid	H_2SO_4
Ammonia	NH_3

1.7 Chemical formula:

1.7.1 Definition:

The chemical formula of a compound is the symbolic representation of the composition of a compound.i.e Chemical formula of water is H_2O.

1.7.2 How to write a Chemical Formula

i. Chemical formulae of compounds are written keeping the following steps in consideration.

ii. Symbols of two elements are written side by side, in the order of positive ion first and negative ion later.

iii. The valency of each ion is written on the right top corner of its symbol, e.g. Na^+, Ca^{2+}, Cl^{1-} and O^{2-} .

iv. This valency of each ion is brought to the lower right corner of other ion by the 'cross-exchange' method, e.g.

$$Na^+Cl^-, \qquad Ca^{2+}Cl^- \qquad and \qquad Ca^{2+}O^{2-}$$
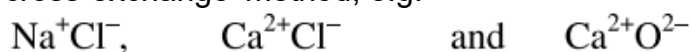
They are written as:

$$NaCl, \qquad CaCl_2 \qquad and \qquad CaO$$
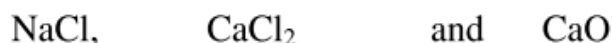
v. If the valencies are same, they are offset and are not written in the chemical formula. But if they are different, they are indicated as such at the same position, e.g., in case of sodium chloride both the valencies are offset and formula is written as NaCl, whereas calcium chloride is represented by formula $CaCl_2$.

vi. If an ion is a combination of two or more atoms which are called radical, bearing a net charge on it, e.g., SO_4^{2-} (sulphate) and PO_4^{3-} (phosphate), then the net charge represents the valency of the radical. The chemical formula of such compounds is written as explained in (iii) and (iv), writing the negative radical within the parenthesis. For example, the chemical formula of aluminum sulphate is written as $Al_2(SO_4)_3$ and that of calcium phosphate as $Ca_2(PO_4)_2$.

1.7.3 Significance of chemical formula:

i. It represents the name of the substance e.g. H_2O for water.

ii. It also represents one mole of the molecules in the balanced chemical equation.

iii. It is in fact one molecule or formula unit of the compound.

iv. It indicates the mass of the compound in amus or grams.

v. It tells the name of the elements as present in the compound.

1.7.4 Types:

Chemical formulae are of two types.

i. Empirical Formula
ii. Molecular Formula

1.7.4.1 Empirical Formula

The simplest type of formula is empirical formula. It is the simplest whole number ratio of atoms present in a compound. The empirical formula of a compound is determined by knowing the percentage composition of a compound.

Examples:

- The covalent compound silica (sand) has simplest ratio of 1:2 of silicon and oxygen respectively. Therefore, its empirical formula is SiO_2 .
- Glucose has simplest ratio 1:2:1 of carbon, hydrogen and oxygen, respectively. Hence, its empirical formula is CH_2O.

The ionic compounds exist in three-dimensional network forms. Each ion is surrounded by oppositely charged ions in such a way to form electrically neutral compound. Therefore, the simplest unit taken as a representative of an ionic compound is called formula unit.

Formula unit

It is defined as the simplest whole number ratio of ions, as present in the ionic compound. For example, formula unit of common salt consists of one Na and one Cl ion and its empirical formula is NaCl. Similarly, formula unit of potassium bromide is KBr, which is also its empirical formula.

1.7.4.2 Molecular Formula

Molecules are formed by the combination of atoms. These molecules are represented by molecular formulae. It shows actual number of atoms of each element present in a molecule of that compound. Molecular formula is derived from empirical formula by the following relationship:

Molecular formula = (Empirical formula)$_n$

Where n is 1,2,3 and so on.

Example

Molecular formula of benzene is C_6H_6 which is derived from the empirical formula CH where the value of n is 6. The molecular formula of a compound may be same or a multiple of the empirical formula. A few compounds having different empirical and molecular formulae are shown in Table 1.5.

Table 1.5 Some Compounds with their Empirical and Molecular Formulae

Compound	Empirical formula	Molecular Formula
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Glucose	CH_2O	$C_6H_{12}O_6$

Some compounds may have same empirical and molecular formula e.g., water (H_2O), Hydrochloric acid (HCl), etc.

Exercise

1. Define empirical formula with an example.
2. Explain why are hydrogen and oxygen considered elements whereas water as a compound.
3. What is the significance of the symbol of an element?
4. Define atomic mass unit. Why is it needed?
5. State the nature and name of the substance formed by combining the following:

   i. Zinc + Copper ii. Water + Sugar

   iii. Aluminium + Sulphur iv. Iron + Chromium + Nickel
6. Differentiate between molecular mass and formula mass, which of the followings have molecular formula?

   H_2O, NaCl, KI, H_2SO_4
7. Differentiate between homoatomic and heteroatomic molecules with examples.
8. Define element and classify the elements with examples

Chapter # 2

Chemical Reactions and Equations

STUDENTS LEARNING OUTCOMES

- Identify the types of Chemical reactions.
- Distinguish between different types of Chemical reactions.
- Daily life examples of Chemical reactions.
- How to balance a given chemical equation
- Balance the given Chemical equation within different chemical environments.

It is our daily observation that water can be changed into ice. It can also be changed into steam. During both these changes, the chemical composition of water (H_2O) and its chemical properties are not changed. It means that liquid water, ice, and steam are three physically different forms of the same substance, i.e., water (H_2O)

When we pass electricity through acidified water (H_2O), it changes into hydrogen (H_2) and oxygen (O_2) which are entirely different substances with different chemical compositions and chemical properties. Such a change in a substance during which entirely new substances with different chemical compositions and properties are formed is called a chemical change. A chemical change is always brought about by a chemical reaction. In this chapter, we will learn further about chemical reactions.

2.1 Chemical Reactions

During Chemical reactions, atoms present in different substances rearrange themselves to form new substances. Burning of coal and natural gas (methane) in air are well-known examples of chemical reactions. Chemically coal is carbon (C). It exists in a solid state and is black. Its burning in the air is a chemical reaction of carbon with oxygen in the air to form carbon dioxide (CO_2). Carbon dioxide is a colourless gas. Its chemical composition and chemical properties are entirely different from those of carbon and oxygen. The rearrangement of atoms that takes place during this chemical reaction can be represented as follows in Figure 2.1
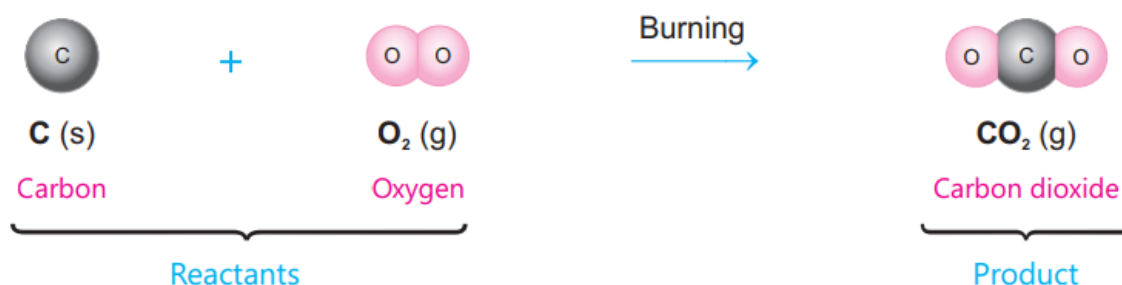
Figure 2.1: Rearrangement of atoms during chemical reaction of carbon with oxygen

Reactants:

Substances which take part in a chemical reaction are called reactants. For example C and O_2 are reactants in Fig.2.1.

Products:

Those substances which are formed as a result of the reaction are called products. For example CO_2 is product in Fig.2.1.

2.1.1 Examples

Burning of methane: When methane burns in the excess of air, carbon dioxide and water are formed. During the rearrangement of atoms in burning of methane (natural gas), carbon atom of methane gets attached with two oxygen atoms to give carbon dioxide while hydrogen atoms attach themselves with oxygen atom to give water vapours shown in Figure 2.2.
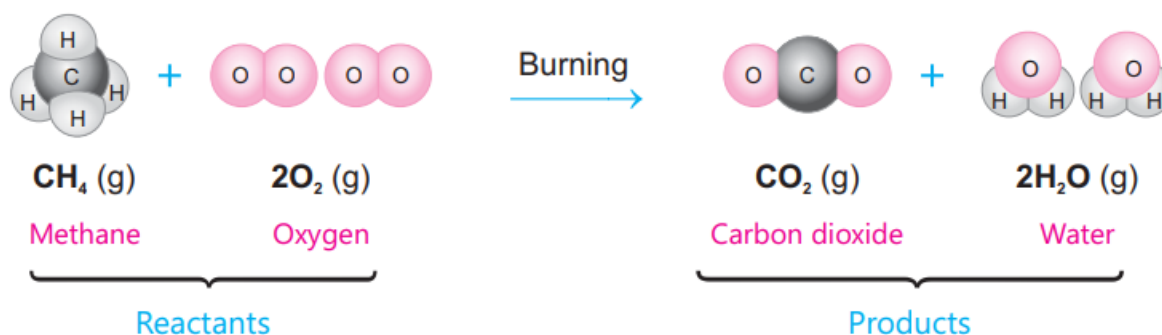
Figure 2.2: Rearrangement of atoms during chemical reaction of methane with oxygen

Chemical reaction of hydrogen with oxygen: The rearrangement of atoms during the chemical reaction of hydrogen with oxygen to form liquid water is shown below in Figure 2.3.

2H₂ (g)   **O₂** (g)   **2H₂O** (g)
Hydrogen   Oxygen   Water

Reactants   Product
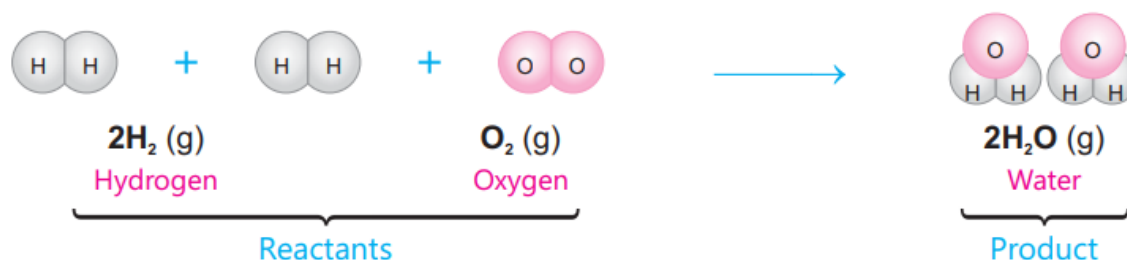
Figure 2.3: Rearrangement of atoms during chemical reaction of hydrogen with oxygen

2.1.2 Applications of Chemical Reactions:

Observing examples of chemical reactions in everyday life is not unusual. Some of the prominent examples are:

Burning:

Fuel (natural gas or petrol, etc.) on its burning in vehicle engine produces different gases. The gases so produced develop pressure to move the piston in the engine and to run the vehicle. Heat produced during burning of fuel in our kitchens is used to cook food. Similarly, the heat produced during burning of fuel in industries is used to produce steam from water.

Photosynthesis:

During photosynthesis in plants, carbon dioxide (CO_2) and water (H_2O) react to produce glucose ($C_6H_{12}O_6$). This reaction takes place in the presence of sunlight and green pigment chlorophyll.

$$6CO_{2\,(g)} + 6H_2O \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} C_6H_{12}O_{6\,(aq)} + 6O_{2\,(g)}$$

Carbon dioxide   Water   Glucose   Oxygen
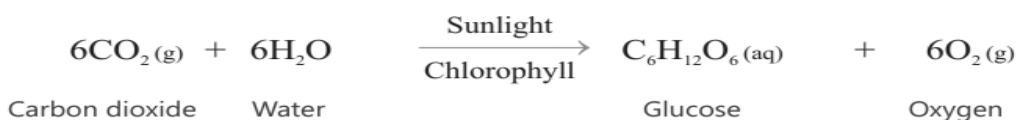
Respiration:

During respiration, the oxygen of air reacts with food (glucose) to produce carbon dioxide and water in the cells of living organisms. The energy produced during this reaction is used to perform all the body functions in living organisms.

$$C_6H_{12}O_{6\,(s)} + 6O_{2\,(g)} \longrightarrow 6CO_{2\,(g)} + 6H_2O + \text{Energy}$$

Glucose   Oxygen   Carbon dioxide   Water

Fermentation:

Conversion of milk into yogurt and formation of baking products involve the chemical changes which are brought about by microorganisms. Such chemical changes or reactions are called fermentation reactions.

Rusting of iron:

Iron (Fe) reacts with oxygen of the air to form iron oxide (Fe_2O_3). The reaction is called rusting of iron and it takes place in the presence of moisture.

$$4Fe_{(s)} + 3O_{2\,(g)} \longrightarrow 2Fe_2O_{3\,(s)}$$
$$\text{Iron} \qquad \text{Oxygen} \qquad\qquad \text{Iron oxide}$$

2.1.3 Types of Chemical Reactions

Chemical reactions are of the following types.

  i.   Synthesis or Combination Reaction
  ii. Decomposition Reaction
  iii. Single-replacement or Single-displacement Reaction
  iv. Double-replacement or Double-displacement Reaction
  v.   Combustion Reaction

2.1.3.1:Synthesis or Combination Reaction:

The reactions in which two or more substances combine to form a single new substance.Combination reactions are called synthesis reactions.

A + B → AB
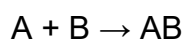
Examples

Solid sodium (Na) metal reacts with chlorine (Cl) gas to produce solid sodium chloride (NaCl)

$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$
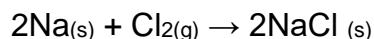
Magnesium rapidly reacts when ignited with oxygen (O_2) to produce a fine powder of magnesium oxide (MgO).

$2 \quad Mg_{(s)} + O_{2\,(g)} \rightarrow 2\,MgO_{(s)}$
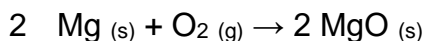
2.1.3.2 :Decomposition Reaction:

A compound breaks down into two or more simple substances in a decomposition reaction.

AB → A + B
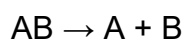
Decomposition reactions are actually reverse to combination reactions.

Examples

Calcium carbonate ($CaCO_3$) decomposes into calcium oxide (CaO) and carbon dioxide (CO_2).

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

Sodium hydroxide (NaOH) decomposes to produce sodium oxide (Na_2O) and water (H_2O).

$$2NaOH_{\ (s)} \rightarrow Na_2O_{\ (s)} + H_2O_{\ (g)}$$

2.1.3.3 Single-replacement or Single-displacement Reaction:
In a single replacement reaction, one element replaces a similar element in a compound.

$A + BC \rightarrow AC + B$
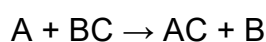
Examples

Zinc (Zn) reacts with hydrochloric acid (HCl) to produce aqueous zinc chloride ($ZnCl_2$) and hydrogen (H_2).

$Zn_{(s)} + 2HCl_{(aq.)} \qquad \rightarrow \qquad ZnCl_{2\ (aq.)} + H_{2\ (g)}$
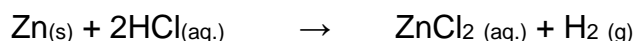
When a strip of magnesium (Mg) metal is placed in an aqueous solution of copper(II)nitrate ($CuNO_3$), it replaces copper, resulting in aqueous magnesium nitrate ($MgNO_3$) and solid copper (Cu) metal.

$Mg_{\ (s)} + Cu(NO_3)_{2\ (aq.)} \rightarrow Mg(NO_3)_{2\ (aq.)} + Cu_{\ (s)}$
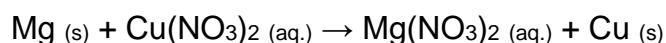
2.1.3.4.Double-replacement or Double-displacement Reaction:
In this type of reaction, the positive and negative ions of two ionic compounds exchange their places to form two new compounds.

$AB + CD \rightarrow AD + CB$
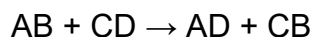
Types

The double-replacement reaction is of two types.

The formation of an insoluble solid in an aqueous solution is called a precipitation reaction. The insoluble solid is called a precipitate.

Examples

When aqueous solutions of potassium iodide (KI) and lead (II) nitrate ($Pb(NO_3)_2$) are mixed, insoluble lead iodide (PbI_2) forms in aqueous potassium nitrate (K_2NO_3).

$2\ KI_{\ (aq.)} + Pb(NO_3)_{2\ (aq.)} \rightarrow K_2NO_{3\ (aq.)} + PbI_{2\ (ppt.)}$
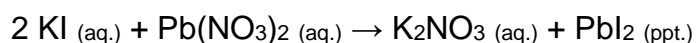
solution of potassium chloride (KCl) and silver nitrate ($AgNO_3$) forms a white insoluble solid, silver chloride (AgCl), in the resulting solution of potassium nitrate (KNO_3).

2 KCl $_{(aq.)}$ + $AgNO_3$ $_{(aq.)}$ → KNO_3 $_{(aq.)}$ + AgCl $_{(ppt.)}$
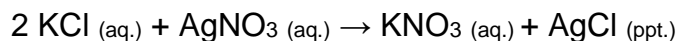
b) Neutralization Reaction

The reaction between an acid and a base is called an acid-base or neutralization reaction. It always forms salt and water.

Example

A mixture of sulfuric acid (H_2SO_4) and sodium hydroxide (NaOH) produces sodium sulfate (Na_2SO_4) and water (H_2O).

H_2SO_4 $_{(aq.)}$ + 2 NaOH $_{(aq.)}$ → Na_2SO_4 $_{(aq.)}$ + 2 H_2O $_{(l)}$
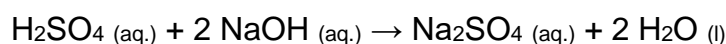
When hydrochloric acid (HCl) reacts with sodium hydroxide (NaOH) . It produces sodium chloride (NaCl) salt and water (H_2O).

HCl $_{(aq.)}$ + NaOH $_{(aq.)}$ → NaCl $_{(aq.)}$ + H_2O $_{(l)}$
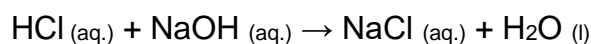
2.1.3.5.Combustion Reaction

A combustion reaction occur when a substance reacts with oxygen gas (O_2), releasing energy through light and heat. Oxygen must be present for a combustion reaction to take place.

Examples

The combustion of hydrogen (H_2) gas in the presence of oxygen (O_2) produces water vapor (H_2O).

$2H_2$ $_{(g)}$ + O_2 $_{(g)}$ → 2 H_2O $_{(g)}$
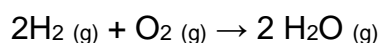
The burning of coal (carbon) in oxygen (O_2) gives carbon dioxide (CO_2).

C $_{(s)}$ + O_2 $_{(g)}$ → CO_2 $_{(g)}$
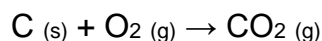
Propane (C_3H_8), a gaseous hydrocarbon, is commonly used as the fuel source in gas grills. It combusts oxygen (O_2) to give carbon dioxide (CO_2) and water (H_2O).

C_3H_8 $_{(g)}$ + 5 O_2 $_{(g)}$ → 3 CO_2 $_{(g)}$ + 4 H_2O $_{(g)}$
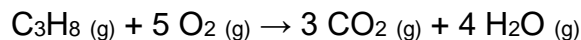
2.1.4 Other Types of Chemical Reactions

1. Redox Reaction:

The Redox reaction is an oxidation-reduction reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing electrons. The atom that loses electrons is said to be oxidized, and the atom that gains electrons is reduced. Redox

reactions can be synthesis, decomposition, single-replacement, or combustion reactions. However, not all combustion reactions are redox reactions.

Example

Iron (Fe) reacts with oxygen (O_2) to give iron (III) oxide. Thus, the oxidation number of iron goes from zero to +3.

Oxidizing agent

gain of electrons (reduction)

$$4\ Fe^0 + 3\ O_2{}^0 \rightarrow 2\ Fe_2{}^{+3}O_3{}^{-2}$$

lose of electron (oxidation)

Reducing agent
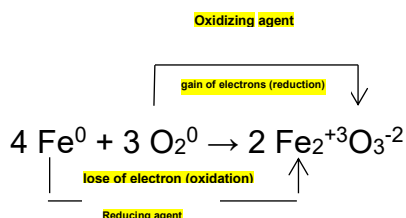
The above reaction is also a combination reaction. Oxygen is called the oxidizing agent since it oxidizes iron to iron (III) oxide.

2. Polymerization:

Polymerization is a process in which relatively small molecules (called monomers) combine chemically to produce a vast chainlike or network of molecules called polymer. This process is known as a chain reaction or polymerization.

Example

The polymerization of ethylene (C_2H_4) produces polyethylene.

$C_2H_4 + C_2H_4 + C_2H_4 \rightarrow (C_2H_4)_n$

Hydrolysis:(Hydro: water -lysis: breakdown)

Hydrolysis is the breakdown of molecule into two parts in the presence of water.

Example

Dissolving sulfuric acid (H_2SO_4) in water (H_2O) yields bisulfate ($HSO_3{}^-$) and hydronium (H_3O^+)

$H_2SO_4 + H_2O \rightarrow HSO_3{}^- + H_3O^+$
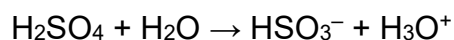
4. Dehydration Synthesis:

Dehydration synthesis is the opposite to hydrolysis. Here, two molecules combine to form a new molecule and eliminate water.

Example

Dehydration of ethanol (C_2H_5OH) at 170 ^0C gives ethene (C_2H_4)

$CH_3 - CH_2 - OH \rightarrow CH_2 = CH_2 + H_2O$
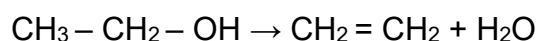
5. Photochemical Reaction:

A photochemical reaction is a type of chemical reaction in which the reactants take in energy in the form of photons from a source of light, like the sun, to form products.

Example

Photography uses the action of light on grains of silver chloride (AgCl) to produce an image. Silver chloride (AgCl) decomposes into silver (Ag) and chlorine (Cl_2) gas.

$2\,AgCl + hv \rightarrow 2Ag + Cl_2$
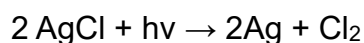
6. Endothermic Reaction:

A chemical reaction is said to be endothermic when it absorbs heat from the surroundings.

Example

Limestone or calcium carbonate ($CaCO_3$) decomposes when heated to a high temperature. Quick lime or calcium oxide (CaO) and carbon dioxide (CO_2) are the products.

$CaCO_{3\,(s)} + heat \rightarrow CaO_{(s)} + CO_{2\,(g)}$
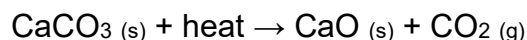
7. Exothermic Reaction:

A chemical reaction is said to be exothermic when it releases heat to the surroundings.

Example

Calcium oxide (CaO) reacts vigorously with water (H_2O) to produce calcium hydroxide [Ca $(OH)_2$] or slacked lime.

$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca\,(OH)_{2\,(aq.)} + heat$
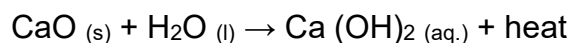
2.2 Chemical Equations and Their Balancing

"A chemical equation is the representation of a chemical reaction in terms of symbols, formulae and signs". In a chemical equation:

   i.   The reactants and products are separated by an arrow.
   ii.   Symbols and formulae of the reactants are written on the left-hand side of the arrow.

iii.   The products are written on the right-hand side of the arrow.
iv.   The arrow is directed towards the products.
v.   Physical states of reactants and products are also expressed along with their formulae or symbols by (s), (g) and (aq.) which stand for solid, gas and aqueous states respectively.

For example:

The chemical equation representing the reaction of sulphur with oxygen to form sulphur dioxide is written as follows.

$S_{(s)} + O_{2\ (g)} \longrightarrow SO_{2\ (g)}$
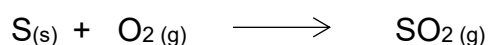
The chemical equation written above shows that sulphur in its solid state reacts with oxygen gas. The product of the reaction, i.e., sulphur dioxide is also a gas. The signs (s) and (g) indicate the physical states of the reactants and the products.

The chemical equation given below indicates that zinc in its solid state reacts with an aqueous solution of sulphuric acid and produces aqueous solution of zinc sulphate and hydrogen gas.

$Zn_{(s)} + H_2SO_{4\ (aq)} \longrightarrow ZnSO_{4\ (aq)} + H_{2\ (g)}$
Zinc   Sulphuric acid   Zinc sulphate   Hydrogen

2.2.1 Balancing the Chemical Equation

The chemical equation in which the number of atoms of each element on both sides of the equation, i.e., reactant side and product side are equal is called a balanced chemical equation.

For example, the chemical equation shown below is a balanced chemical equation.
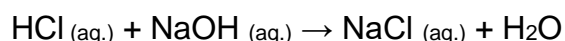
$HCl_{(aq.)} + NaOH_{(aq.)} \rightarrow NaCl_{(aq.)} + H_2O$

The chemical equation in which the number of atoms of each element on both sides of the equation, i.e., reactant side and product side is not equal is called an unbalanced chemical equation. For example, the chemical equation given below is an unbalanced chemical equation.

$H_{2\ (g)} + Cl_{2\ (g)} \longrightarrow HCl_{(g)}$
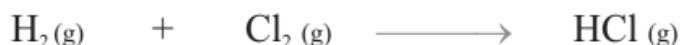
Method:

The trial-and-error method is commonly used. According to this method, the trial and error process of adjusting coefficients before symbols or formulae is continued till the number of atoms of each element on both sides of the equation becomes equal.

The working rules for balancing a chemical equation are as follows:

i. Write the unbalanced equation and count the number of atoms of each element on both sides of the arrow.
ii. Work with one element at a time.
iii. Multiply the symbol or formula with suitable integers (2, 3, 4, 5, etc.) on that side of the equation where the number of atoms of a particular element is less and try to balance the element on both sides of the equation. Start multiplying with relatively small numbers.
iv. Repeat the process for all the elements one by one.
v. Balance the diatomic molecules like H_2, N_2, O_2, Cl_2, etc. at the end.
vi. Some examples for balancing the equation are given below:

Example 2.1- Consider the following equation:

$$N_{2(g)} \quad + \quad H_{2(g)} \quad \longrightarrow \quad NH_{3(g)}$$

Step I

Count the number of atoms of each element on both sides of the arrow.

Reactants	Products	Balanced/Unbalanced
2 N atoms	1 N atom	N is unbalanced
2 H atoms	3 H atoms	H is unbalanced

Step II

Add appropriate coefficient to balance N:

$$N_{2(g)} \quad + \quad H_{2(g)} \quad \longrightarrow \quad 2NH_{3(g)}$$

Reactants	Products	Balanced/Unbalanced
2 N atoms	2 N atoms	N is balanced
2 H atoms	6 H atoms	H is unbalanced

Thus, the equation is balanced.

Example 2.2:

$$CH_{4(g)} \quad + \quad O_{2(g)} \quad \longrightarrow \quad H_2O_{(g)} \quad + \quad CO_{2(g)}$$

Step I

Count the number of atoms of each element or compound on both sides of the arrow:

Reactants	Products	Balanced/Unbalanced
1 C atom	1 C atom	C is balanced
4 H atoms	2 H atoms	H is unbalanced
2 O atoms	3 O atoms	O is unbalanced

Step II

$$CH_{4(g)} \ + \ 2O_{2(g)} \longrightarrow 2H_2O_{(g)} \ + \ CO_{2(g)}$$

Add appropriate coefficients:

Reactants	Products	Balanced/Unbalanced
1 C atom	1 C atom	C is balanced
4 H atoms	4 H atoms	H is balanced
4 O atoms	4 O atoms	O is balanced

Thus the equation is balanced.

Exercise

Q.1 Answer the following questions briefly.

    I.   Define a chemical reaction.
    II.   What are reactants?
    III.   What are products?
    IV.   What is a chemical equation?

Q.2. Differentiate between the following.

    I.   Addition reaction and decomposition reaction.
    II.   Balanced chemical equation and unbalanced chemical equation.
    III.   Exothermic reaction and endothermic reaction.

Q.3. Complete and balance the following incomplete equations.

(i) $Mg_{(s)}$ + $O_{2(g)}$ $\longrightarrow$

(ii) $CH_{4(g)}$ + $O_2(g)$ $\longrightarrow$

(iii) $Fe_{(s)}$ + $S_{(s)}$ $\longrightarrow$

(iv) $N_{2(g)}$ + $H_{2(g)}$ $\longrightarrow$

(v) $Na_{(s)}$ + $Cl_{2(g)}$ $\longrightarrow$
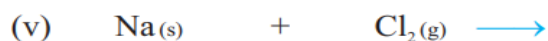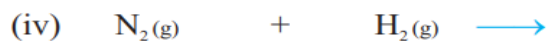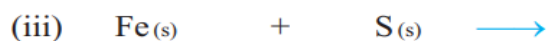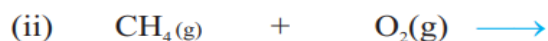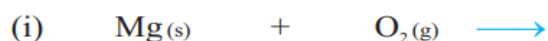
Q.4. Write at least two examples of the following chemical reactions.

    I.   Addition reaction
    II.   Decomposition reaction
    III.   Exothermic reaction
    IV.   Endothermic reaction

Q.5. Describe the applications of chemical reactions.

Q.6. Write down the rules for balancing chemical equations.

Q.7. Describe the importance of exothermic reactions in everyday life.

Q.8. Give two examples of chemical reactions from everyday life which are essential for life.

Q.9. How do the following reactants react together? Write down complete reactions and balance the resulting equations.

    I.   Iron + Hydrochloric acid
    II.   Calcium oxide + Carbon dioxide
    III.   Carbon monoxide + Oxygen
    IV.   Methane + Oxygen
    V.   Carbon dioxide + Water

Q.10. Balance the following equations.

(i) $Ca(HCO_3)_2 +$ HCl $\longrightarrow$ $CaCl_2$ $+ CO_2 + H_2O$

(ii) NaBr $+$ Cl_2 $\longrightarrow$ NaCl $+ Br_2$

(iii) Fe $+$ O_2 $\longrightarrow$ Fe_2O_3

(iv) NH_4OH $+$ $H_2SO_4 \longrightarrow$ $(NH_4)_2SO_4 + H_2O$

(v) Zn $+$ HCl $\longrightarrow$ $ZnCl_2$ $+ H_2$

Chapter # 3

WATER

STUDENTS LEARNING OUTCOMES:

- Describe the occurrence of water and its importance in the environment including industry.
- Review our dependence on water and its importance of maintaining its quality.
- Describe the composition and properties of water.
- Differentiate between types of Hydrolysis.
- Distinguish between types of contaminants.
- Differentiate among soft, temporary and permanent hard water.
- Describe methods for eliminating temporary and permanent hardness of water.
- The importance and significance of water has been recognized by mankind.

3.1 Importance and significance:

The importance and significance of water has been recognized by mankind.

The importance of water is because of two reasons.

- First, it is an essential and major component of every living cell. For example, the human body consists of about 70% water.
- Secondly, it provides an environment for animals and plants that live in water. So, all living organisms owe their life to water.

We use water in daily life for drinking, cooking, and washing purposes. The quality of drinking water has been a major factor in determining human health and welfare for ages. Since World War II, there has been a rapid production and use of synthetic chemicals. Many of these chemicals (fertilizers and pesticides run off from agricultural lands and industrial discharge from industrial units) have polluted water supplies. Besides this, there is also a threat to groundwater from waste chemical dumps and landfills. Currently, waterborne toxic chemicals pose the greatest threat to water supplies, especially in urban areas. The use of this water is causing waterborne diseases. So, the use of polluted water is a concern of every citizen. Therefore, understanding the sources and effects of water pollution is essential for controlling this alarming threat.

3.2 Occurrence of Water

The oceans contain about 97% of the world's water. The rest of the water is in the form of glaciers, ice caps, groundwater, and inland water (rivers, lakes, streams). It is also present in the atmosphere in the form of water vapours. Sea water is unfit for drinking and agricultural purposes due to the high percentage of dissolved salts. Only 0.2% of the total water on the Earth is potable, i.e. fit for drinking purposes.

Distribution of water
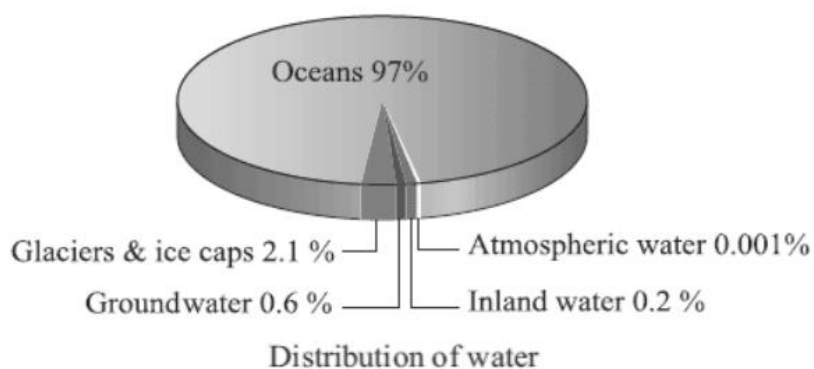
3.3 Properties of Water

3.3.1 Physical Properties:
Water is composed of two elements: oxygen and hydrogen. One atom of oxygen combines with two atoms of hydrogen to form one molecule of water.

I.Appearance:

Water is a colorless, odorless, and tasteless liquid in its natural state. The crystal structure of water in its solid state is hexagonal.

II.Boiling Point of Water:

The boiling point is defined as "the temperature at which the vapor pressure of the liquid is equal to the pressure surrounding the liquid"

At this temperature the liquid changes into vapor.

The boiling point of water is 100°C.

III.Freezing Point of Water:

The freezing point is "the temperature at which the substance changes its state from liquid to solid". The freezing point of water is 0°C or 32°F.

Specific Heat Capacity: Water has a high specific heat capacity of 4.2 joules per gram at 25°C. This is due to the extensive hydrogen bonding between the water molecules.

IV.Density of Water:

The density of a material is defined as its mass per unit volume.

$$\text{Density} = \frac{mass}{volume} = \frac{m}{V}$$

So, the Density of Water can be defined as:

"It is the weight of the water per unit volume, which depends on the temperature of the water". The density of water at 0 ^{0}C is 0.998 g cm^{-3} and at 4 ^{O}C is 1.000 g cm^{-3}.

V.Viscosity of Water:

Viscosity is a measure of a fluid's resistance to flow.

In other words , viscosity is the thickness of the liquid –e.g., syrup and water. The viscosity of water is 0.89 cP (centi-poise).

VI.Surface Tension of Water:

Surface tension is "the tendency of the fluid to shrink in a minimum surface area". Water has a high surface tension of 72 mN/m at room temperature (25^{0}C). Because of this high surface tension of water, insects can walk on the surface of the water without any discomfort.

VII.Refractive Index of Water:

In simpler words, the refractive index is the number that describes how fast the light reached the material. The refractive index of water is 1.333 at 20°C.

VIII.Dielectric Constant of Water:

Dielectric constant can be defined as "the measure of a substance's ability to store electrical energy". The dielectric constant of water is very high, which is at 78.6. This constant plays a very important role in water, being a universal solvent.

3.3.2.Chemical Properties
  i.   It is neutral to litmus.
  ii.   Its freezing point is 0°C and its boiling point is 100 °C at sea level.
  iii. Its maximum density is 1 gcm^{-3} at 4°C.
  iv.   It is an excellent solvent for ionic as well as molecular compounds.
  v.   It has an unusually high heat capacity of about 4.2 Jg^{-1}K^{-1}, which is about six times greater than that of rocks. This specific property of water is responsible for keeping the Earth's temperature within limits. Otherwise, daytime temperature would have been too high to bear and night time temperature would have been too low to freeze everything.

vi.   It has high surface tension. This unique property of water is responsible for its high capillary Capillary action is the process by which water rises from the roots of plants to leaves. This process is vital for the survival of the land plants.

3.4 Water as Deliquescence

Property of water to get in contact with crystalline solids when exposed to the atmosphere at ordinary or at any temperature, they initially become wet then solid loses its crystalline form and finally dissolve in water to form a solution.

Deliquescent can be said as a tendency to become liquid. Deliquescent substances are solids that tend to absorb moisture from the air and dissolve it. For example, $NaOH$, KOH, $MgCl_2$, $CaCl_2$, $FeCl_3$, $LiCl$, $Cu(NO_3)_2$, $NaNO_3$, $LiNO_3$ etc. are deliquescent substances.

3.5 Efflorescence

Certain hydrated crystalline salts, when exposed to the atmosphere at ordinary temperature lose their water of crystallization, molecules either partially or completely becomes anhydrous.

Efflorescence is the automatically loss of water by aqueous salts. For example, $Na_2SO_4.10H_2O$, $FeSO_4.7H_2O$, Blue Vitriol ($CuSO_4.5H_2O$), washing soda ($Na_2CO_3.10H_2O$) lose 9 water molecules first and then become completely anhydrous.

3.6 Hygroscopic substances

Certain substances absorb moisture from the atmosphere at ordinary temperatures but do not dissolve in it. These are called hygroscopic substances, and the property is known as hygroscopy. For example, Sodium trioxonitrate.

3.7 WATER AS A SOLVENT

Water is the universal solvent because it can dissolve almost all the minerals. Its ability to dissolve substances is because of two unique properties of water:

i.   Polarity of water molecule
ii.   Exceptional hydrogen bonding ability

3.7.1 Polar nature of water

Water molecules has polar structure, i.e. one end of the molecule is partially positive while the other end is partially negative because of electronegativity difference between oxygen and hydrogen atoms. All other polar substances are soluble in water, because

the positive end of the substance is attracted by the negative end (Od-) of the water and negative end of the substance is attracted by the positive end (H^+) of the water. The electrostatic attractions among the ions are overcome by the ion-dipole forces of attraction between ion and water molecules.

In this way, positive and negative ions of the compounds are pulled apart as shown in Figure 3.1. Ultimately, these oppositely charged ions are surrounded by water molecules, thus separated and kept in solution. For example, most salts like NaCl, KCl, Na_2SO_4, etc. are soluble in water. On the other hand, many covalent substances like benzene, ether, octane, etc., which do not have polar ends or bonds are not attracted by water molecules. Therefore, non-polar compounds do not dissolve in water.
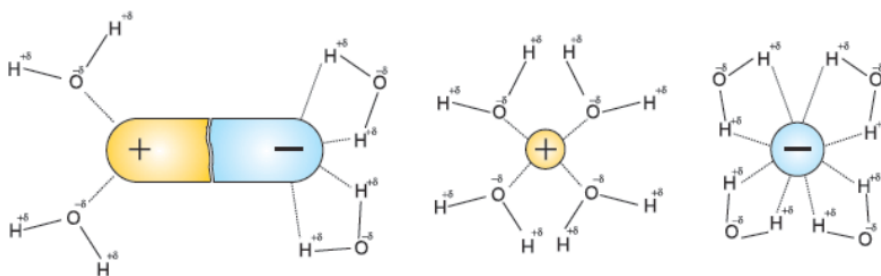
Fig 3.1.Dissolving process of a polar substance in water

3.7.2 Extensive hydrogen bonding ability

Water molecule is composed of oxygen and hydrogen atoms. Because of two O—H bonds and two lone pairs, one H_2O molecule can form hydrogen bonding with four other H_2O molecules, which are arranged tetrahedrally around the H_2O molecule as shown in Figure 3.2. This unique behavior of water enables it to dissolve many polar non-ionic compounds having a hydroxyl group (-OH), like alcohols, organic acids, glucose, sugar, etc. by forming hydrogen bonds with them.
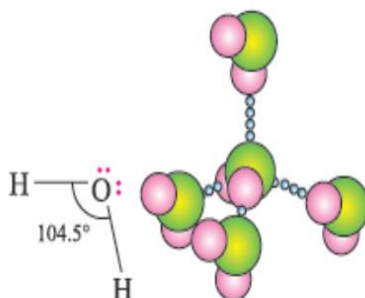
Fig. 3.2 Hydrogen bonding of water molecule

3.8 Hydrolysis:

Hydrolysis is a common form of a chemical reaction where water is mostly used to break down the chemical bonds that exist between a particular substance.

Hydrolysis is derived from the Greek word hydro meaning "water" and lysis that means "breakdown" or to "unbind". Usually in hydrolysis, the water molecules get attached to two parts of a molecule. One molecule of a substance will get the H^+ ion and the other molecule receives the OH^- group. Hydrolysis reaction is mainly used to break down polymer into monomers.

Examples:

Dissolving a salt of a weak acid or base in water or dissolving sulphuric acid in water where hydronium and bisulfate compounds are formed.

Hydrolysis also helps in breaking down proteins, fats, and complex carbohydrates in food.

The general formula of a hydrolysis reaction is:

$H_2O \quad \rightleftharpoons \quad H^+ + OH^-$

3.8.1 Types of Hydrolysis

There are several types of hydrolysis, we will look at them in brief below.

Salts: This is the most common type of hydrolysis. Hydrolysis of salts generally refers to the reaction of salt with water where it involves the interaction between cations or anions of salts and water. During hydrolysis, a salt breaks down to form ions, completely or partially depending upon the solubility factor.
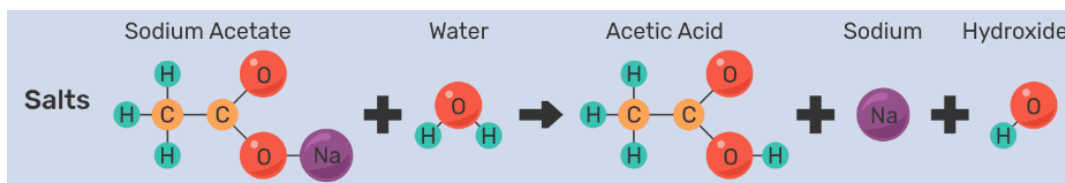
Acid and Base: Acid–base-catalyzed hydrolysis can be found during the hydrolysis of esters or amides. Here, the process of hydrolysis occurs when water or hydroxyl ion reacts with the carbon of the carbonyl group of the ester or amide where new compounds are formed. The products of both hydrolysis are compounds with carboxylic acid groups.
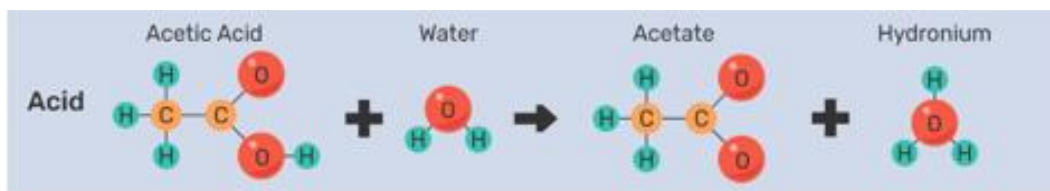
ATP: Most biochemical reactions that occur in living organisms are in the form of ATP hydrolysis which takes place with the help of enzymes acting as catalysts. The catalytic action of enzymes allows the hydrolysis or breaking down of proteins, lipids, oils, fats and carbohydrates.

3.9.Impurities:

Impurities in water are any substances or particles that affect its quality, safety and usability. Impurities in water can be divided into the following categories.

i.   Biological contamination
ii.   Chemical contamination
iii. Physical contamination

3.9.1 Biological contamination
The most-deadly impurities found in water, biological impurities make the water extremely unsafe for use. Biological contaminants such as bacteria, protozoa, viruses, and parasites may lead to fatal water-born diseases and health disorders. The minute organisms present in water lead to diseases such as typhoid fever, dysentery, and many other health problems. Though microorganisms survive different temperatures, these can't be removed from water easily without using the right purification technology.

3.9.2.Chemical contamination
Water also contains some dangerous chemicals that can't be removed by the simple filtration process. The reason is that water is a universal solvent that easily dissolves everything. The total dissolved solids present in water measures the number of dissolved impurities in water and is expressed in 'PPM'. The harmful chemicals present in water include nitrate, mercury, and fluoride which are present in the waste dumped by industries. To remove these chemicals from water, you need to use an RO water purifier.

3.9.3.Physical contamination
The physical impurities mainly affect the appearance and physical properties of water. Sediment or organic materials suspended in water bodies because of soil erosion. The traditional methods of water purification don't remove physical contaminants from water.

RO water purifiers use a multistage purification process to remove physical contaminants from water.

3.10 SOFT AND HARD WATER

Soft water is water which produces a good lather with soap.

Hard water is water which does not produce lather with soap.

3.10.1 Causes of hardness in water

The rain water while coming down absorbs carbon dioxide from the atmosphere. The water mixed with carbon dioxide, when passing through the beds of the soil, converts insoluble carbonates of calcium and magnesium into soluble bicarbonates. It may also dissolve chlorides and sulphates of calcium and magnesium. These salts make the water hard.

$$CaCO_{3(s)} + CO_{2(g)} + H_2O_{(l)} \longrightarrow Ca(HCO_3)_{2(aq)}$$

$$MgCO_{3(s)} + CO_{2(g)} + H_2O_{(l)} \longrightarrow Mg(HCO_3)_{2(aq)}$$
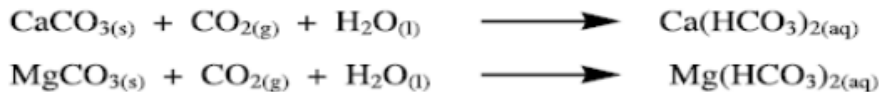
Thus, rain water dissolves many salts of divalent cations like Mg^{2+}, Ca^{2+}, and anions like Cl^-, SO_4^{2-}, HCO_3^- and CO_3^{2-} for example, gypsum ($CaSO_4 .2H_2O$) and limestone ($CaCO_3$). These salts make the water hard. Gypsum is sparingly soluble in water, while limestone is insoluble in water. However, in the presence of carbon dioxide small quantity of limestone is soluble in water according to the above chemical reaction.

3.10.2.TYPES OF HARDNESS:

Hardness is of two types:

Temporary hardness is because of the presence of bicarbonates of calcium and magnesium. It can be remove by using simple physical method i.e. boiling.

Permanent hardness is because of the presence of sulphates and chlorides of calcium and magnesium. It can be remove by using different chemicals. i.e. washing soda and zeolites.

3.10.3 Methods for Removing Hardness:

The removal of Mg^{2+} and Ca^{2+} ions which are responsible for the hardness is called water softening.

3.10.3.1 Removal of temporary hardness
(a) By Boiling

Temporary hardness of water is easily removed by boiling the water. On boiling, calcium bicarbonate $Ca(HCO_3)_2$ decomposes to produce insoluble calcium carbonate, which precipitates out of the solution.

$$Ca(HCO_3)_{2(aq)} \xrightarrow{\Delta} CaCO_{3(s)} + H_2O + CO_{2(g)}$$

(b) Clark's method

A chemical method to remove temporary hardness is by the addition of slaked lime $Ca(OH)_2$. A calculated amount of lime water is added to temporary hard water. Thus, once the magnesium and calcium ions precipitate out water becomes soft.

$$Mg(HCO_3)_{2(aq)} + Ca(OH)_{2(aq)} \longrightarrow Mg\,CO_{3(s)} + Ca\,CO_{3(s)} + 2H_2O_{(l)}$$

$$Ca(HCO_3)_{2(aq)} + Ca(OH)_{2(aq)} \longrightarrow 2CaCO_{3(s)} + 2H_2O_{(l)}$$
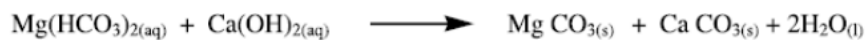
3.10.3.2 Removal of permanent hardness

Permanent hardness can only be removed by using chemicals. Calcium (Ca^{2+}) and magnesium (Mg^{2+}) are removed as insoluble salts by adding washing soda (Na_2CO_3) or sodium zeolite.

(a) By using washing soda

The addition of washing soda removes the calcium and magnesium ions as the insoluble calcium and magnesium carbonates, respectively.

$$Na_2CO_{3(aq)} + CaSO_{4(aq)} \longrightarrow CaCO_{3(s)} + Na_2SO_{4(aq)}$$

$$Na_2CO_{3(aq)} + MgSO_{4(aq)} \longrightarrow MgCO_{3(s)} + Na_2SO_{4(aq)}$$
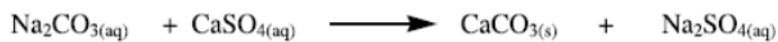
(b) By using Sodium Zeolite (an ion Exchanger)

Sodium zeolite is a naturally occurring resin of sodium aluminium silicate $NaAl(SiO_3)_2$, which can also be prepared artificially. It is used for softening of water on a domestic as well as on an industrial scale.

$$Na_2 \text{ zeolite} + CaSO_{4(aq)} \longrightarrow Ca \text{ zeolite} + Na_2SO_{4(aq)}$$
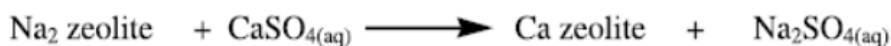
When water is passed through resin, sodium ions of the resin are exchanged with the unwanted
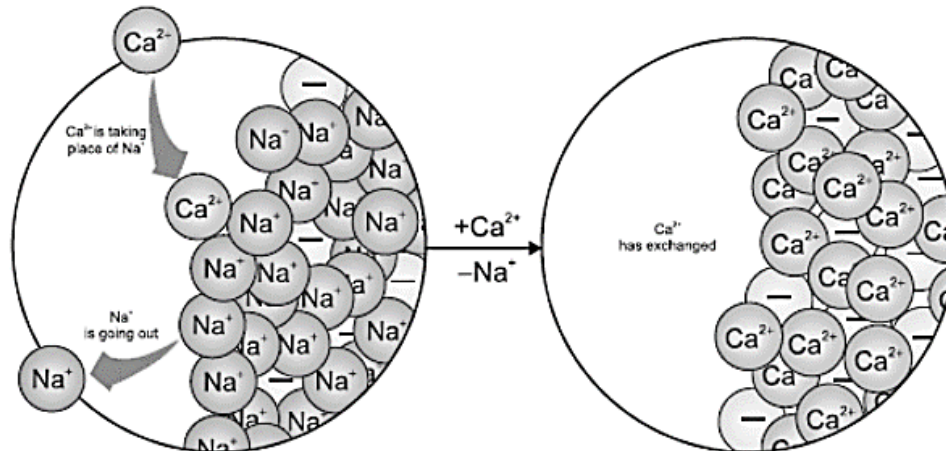
Fig. 3.3 Ion exchange for removal of hard water ions

calcium and magnesium ions of the hard water as shown in figure 3.3.

$$\text{Ca zeolite} \quad + \quad 2\text{NaCl} \quad \longrightarrow \quad \text{Na}_2 \text{ zeolite} \quad + \quad \text{CaCl}_2$$

When resin is fully used up it can be regenerated by flushing it with a concentrated solution of NaCl. The reverse process takes place because of the high concentration of sodium ions.

3.10.4 Disadvantages of Hard Water
I.   Hard water consumes large amounts of soap for washing purposes.
II.   Drinking hard water causes stomach disorders.
III. Hard water is unfit for use in steam engines, boilers, and turbines because insoluble calcium and magnesium salts deposit inside. They are called scales. They are bad conductors of heat, and hence more fuel is used. Insoluble calcium and magnesium sulphates not only reduce the efficiency of the engine but also cause the boiler to burst.

3.11 Water Cycle

The water cycle is an important biochemical cycle involved in the flow or circulation of water through different levels of the ecosystem. The water cycle is defined as a natural process of constantly recycling the water in the atmosphere. It is also known as the hydrological cycle or the hydrologic cycle.

During the process of the water cycle between the earth and the atmosphere, water changes into three states of matter – solid, liquid and gas.
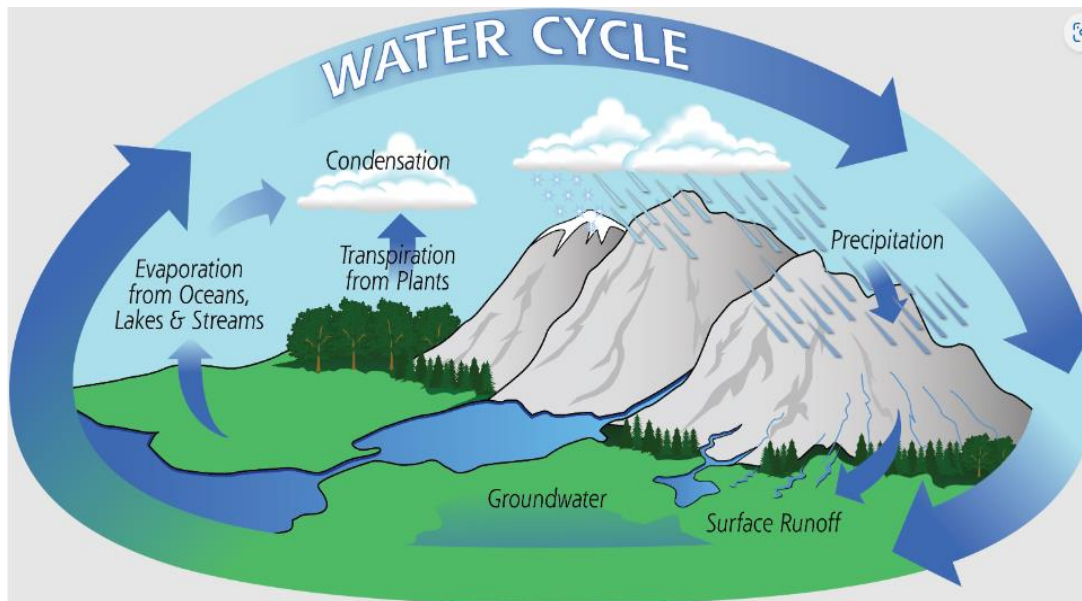
3.11.1:Stages of Water Cycle

The complete water cycle is carried into four stages which are as follows:

- Evaporation
- Condensation
- Precipitation
- Collection.

3.11.1.1:Evaporation

This is the initial stage of the water cycle. The process by which water from its liquid state changes to vapour, a gaseous state, is termed evaporation. During the water cycle, water in the water bodies gets heated up and evaporates in the form of vapour, mixes with the air and disappears.

3.11.1.2.Condensation

When the evaporated water vapor loses its thermal energy, it becomes liquid through the process of condensation. The formations of clouds are examples of condensation.

3.11.1.3 Precipitation

Rain, snow, sleet, and hail are all examples of Precipitation. After the condensation, atmospheric water vapor forms sufficiently large water droplets and falls back to the earth with the help of gravity.

3.11.1.4 Deposition or Collection

This is the final stage of the water cycle. Deposition occurs when evaporated water vapor falls back to earth as precipitation. This water may fall back into the different water

bodies, including oceans, rivers, ponds, and lakes, and even end up on the land, which in turn becomes a part of the groundwater.

Overall, the water cycle process describes how water is balanced in the atmosphere. It also plays an important role in ensuring the availability of water for all living organisms and it has a great impact on our environment.

EXERCISE

Short Questions

i.   How water rises in plants?
ii.   Point out two properties of water that make it an excellent solvent.
iii.   Why is the water molecule polar?
iv.   Which salts are responsible for hardness of water?
v.   Explain the chemistry of removing the temporary hardness by boiling water.
vi.   What is the principle of removing permanent hardness of water?
vii.   What are the reasons for waterborne diseases?
viii.   Differentiate between soft and hard water.
ix.   What are the causes of hardness in water?
x.   What are the effects of temporary hardness in water?
xi.   What are the Disadvantages of Hard Water?
xii.   Differentiate between Deliquescence and Efflorescence.
xiii.   Extensive Questions
xiv.   Write a detail note on water cycle.
xv.   How polarity of water molecule plays its role to dissolve the substances?
xvi.   Explain the methods of removing permanent hardness.
xvii.   What is Hydrolysis? discuss its types.
xviii.   Explain the reasons, why water is considered a universal solvent.
xix.   Differentiate between different types of impurities.

Chapter # 4

Solutions

STUDENTS LEARNING OUTCOMES:

- Define the terms: solution, aqueous solution, solute and solvent and give an example of each.
- Explain the difference between dilute, and concentrated solutions.
- Explain what is meant by the concentration of a solution.
- Define percentage solution.
- Define molarity, molality, and mole fractions.
- Define percentage solution.
- Solve problems involving the molarity of solution.
- Explain the formation of solutions (mixing gases into gases, gases into liquids, gases into solids) and give an example of each.
- Explain the formation of solutions (mixing liquids into gases, liquids into liquids, liquids into solids) and give an example of each.
- Explain the formation of solutions (mixing solids into gases, solids into liquids, solids into solids) and give an example of each.
- Explain what is meant by the concentration of a solution.
- Concept of Solubility and Solubility curves.
- Use the rule that "like dissolves like" to predict the solubility of one substance in another.

Every sample of matter with uniform properties and a fixed composition is called a phase. For example, water at room temperature and normal pressure exists as a single liquid phase, that is, all the properties of water are uniform throughout this liquid phase.

If a small amount of sugar is added to this sample of water, the sugar dissolves but the sample remains in a single liquid phase. However, the properties and composition of this new liquid phase, now the sugar solution, are different from those of pure water. As this solution of sugar in water contains two substances (binary solution), so it is a mixture and since its properties are uniform, therefore, it is homogeneous.

4.1 Solution:

A solution is a homogeneous mixture of two or more kinds of different molecular or ionic substances.

Solvent: The substance which is present in large quantities is called a solvent

Solute: The other component in small quantity is called a solute.
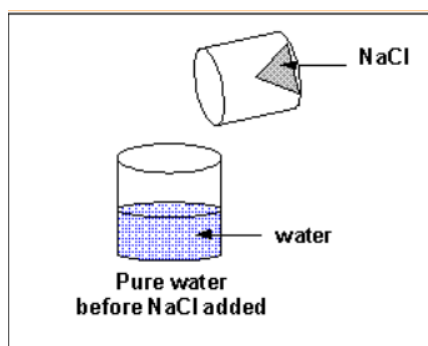
Fig 4.1: CONCEPT OF A SOLUTION

For a given solution, the amount of solute dissolved in a unit volume of solution (or a unit amount of solvent) is termed as the concentration of the solution.

Dilute Solutions: Solutions containing relatively lower concentrations of solute are called dilute solutions.

Concentrated Solutions: Those containing relatively higher concentrations of solutes are called concentrated solutions.

4.2 CONCENTRATION UNITS OF SOLUTIONS

There are various types of concentration units of solutions.

4.2.1 Percentage composition:
The amounts of solute and solvent can be expressed in percentage composition by four different ways.

      a. Percentage weight/weight
      b. Percentage weight/volume
      c. Percentage volume/weight
      d. Percentage volume/volume

(a) Percentage weight/weight:

It is the weight of a solute dissolved per 100 parts by weight of solution. 5% w/w sugar solution will contain 5 g of sugar dissolved in 100 g of solution in water. This solution contains 95 g of water.

$$\% \text{ by weight} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

Example 4.1:

Calculate the percentage by weight of NaCl, if 2.0g of NaCl is dissolved in 20 g of water.

Solution:

Weight of NaCl = 2.0g

Weight of solvent = 20.0g

Weight of solution = 20+2 = 22g

$$\% \text{ by weight} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

% of NaCl by weight = $\frac{2.0\ g}{22g}$ x100 = 9.09 % $\frac{w}{w}$

(b) Percentage Weight/Volume:

It is the weight of a solute dissolved per 100 parts by volume of solution. 10 g of glucose dissolved in 100 cm^3 of solution is 10% w/v solution of glucose. The quantity of the solvent is not exactly known. In such solutions, the total volume of the solution is under consideration.

c) Percentage Volume /Weight :

It is the number of cm^3 of a solute dissolved per 100 g of the solution. If we dissolve 10 cm^3 of alcohol in water and the total weight of the solution is 100 g, then it is 10% v/w solution of alcohol in water. In such type of solutions, we don't know the total volume of the solution.

(d) Percentage Volume / Volume:

It is the volume of a solute dissolved per 100 cm3 of the solution. This unit of concentration is best applicable to the solutions of liquids in liquids. A 12 % alcohol beverage is 12 cm^3 of alcohol per 100 cm^3 of solution. In such solutions, the total volume of the solution may not be necessarily equal to the sum of volumes of solute and the solvent.

4.2.2.Molarity (M)

Molarity is the number of moles of solute dissolved per dm3 of the solution.

To prepare one molar solution of glucose in water, we take 180 g of glucose and add sufficient water to make the total volume 1 dm^3 (liter) in a measuring flask.

In case of one molar solution of sucrose, 342 g of sucrose dissolved in water to make it $1dm^3$. Since the volume of 342 g of sucrose is greater than 180 g of glucose so the volume of water in a 1 molar sucrose solution is less than that of 1 molar glucose solution. Anyhow, to calculate the volume of the solvent, we need to know the density of the solute. The following formula is used to prepare the solution of any molarity.

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Volume of soultion (dm}^3)}$$

or

$$\text{Molarity(M)} = \frac{\text{Number of moles of solute}}{\text{Volume of soultion (dm}^3)}$$

Examples 4.2:

Calculate the molarity of a solution containing 20.7 g of K_2CO_3 dissolved in 500 cm^3 of the given solution.

Solution:

Mass of K_2CO_3 = 20.7g

Molar mass of K_2CO_3 = 138gmol^{-1}

volume of solution = 500cm^3 = 0.5dm^3

Formula: $$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Volume of soultion in dm}^3}$$

$$\text{Molarity} = \frac{20.7g}{138gmol^{-1}} \times \frac{1}{0.5dm^3} = 0.3\text{mol dm}^{-3} = \boxed{0.3\text{mol dm}^{-3}} \text{ Answer}$$

4.2.3.Molality (m)
Molality is the number of moles of solute in 1000 g (1 kg) of the solvent.

To prepare molal solutions, we don't have to take any flask. 180 g of glucose when dissolved in 1000 g of water gives one molal solution of glucose. The total mass of the solution is 1180 g. We don't know the volume of the solution. To know the volume, we need the density of the solution. For one molal sucrose solution, 342 g of sucrose are dissolved in 1000 g of H2O.

So, 1 molal solution of different solutes in water have its own masses and volumes. To get the molality of any solution, we use the following equation.

$$\text{Molality(m)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Mass of solvent in kg}}$$

or $$\text{Molality(m)} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$$

Example 4.3:

What is the molality of a solution prepared by dissolving 5g of toluene (C7H8) in 250g of benzene.

Solution:

Mass of toluene = 5g

Mass of benzene = 250g = 0.25kg

Molar mass of toluene = 12x7+1x8=92

Formula used

$$\text{Molality(m)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Mass of solvent in kg}}$$

$$\text{Molality(m)} = \frac{5g}{92g \ mol^{-1}} \times \frac{1}{0.250 \ kg}$$

$$= \frac{5}{92 \times 0.25} mol \ kg^{-1} = \boxed{0.217 mol \ kg^{-1}} \ \text{Answer}$$

$$= 0.217m$$

The molality of a solution is an indirect expression of the ratio of the moles of the solute to the moles of the solvent. The molal aqueous solution of a solute say glucose or NaOH is dilute in comparison to its molar solution. The reason is that in molal solution the quantity of the solvent is comparatively greater. The value of concentration given in the units of molality does not change with temperature but that of molarity does. The reason is that the volume of liquids are affected by the variation in temperature.

4.2.4-Mole Fraction (x)

The mole fraction of any component in a mixture is the ratio of the number of moles of it to the total number of moles of all the components present.

This unit of concentration is for any type; of solution i.e. gas in gas, liquid in liquid or solid in liquid, etc. This unit is also applicable to a solution having more than two components.

Let there be three components A, B, C making a solution. The number of moles are n_A, n_B, n_C respectively. If the mole fraction of A, B and C are denoted by x_A, x_B, x_C respectively, Then,

$$X_A = \frac{n_A}{n_A + n_B + n_C}$$

$$X_B = \frac{n_B}{n_A + n_B + n_C}$$

$$X_C = \frac{n_C}{n_A + n_B + n_C}$$

The sum of the mole fractions of all the components of a solution must be equal to one. There are no formal units of mole fraction. Anyhow, we sometimes multiply mole fraction by 100 to get mole percent.

Example 4.4:

Calculate the mole fraction and mole percent of each component in a solution having 92 g of ethyl alcohol, 96 g of methyl alcohol and 90 g of water.

Solution:

First of all get the number of moles of each component.

$$\text{Number of moles of the substance} = \frac{\text{Mass in grams of the substance}}{\text{Molecular mass in grams}}$$

Molar mass of ethyl alchohol (C_2H_5OH) = 46 gmol^{-1}

$$\text{Number of moles of ethyl alchohol} = \frac{92g}{46 \text{ gmol}^{-1}} = 2 \text{ mol}$$

Molar mass of methyl alchohol (CH_3OH) = 32 gmol^{-1}

$$\text{Number of moles of methyl alchohol} = \frac{96g}{32 \text{ gmol}^{-1}} = 3 \text{ mol}$$

Molar mass of water(H_2O) = 18 gmol^{-1}

$$\text{Number of moles of water} = \frac{90g}{18 \text{ gmol}^{-1}} = 5 \text{ mol}$$

The mole fraction of any components is ratio of its moles to total number of moles.

$$X_{\text{ethyl alcohol}} = \frac{2}{2+3+5} = \frac{2}{10} = \boxed{0.2} \text{ Answer}$$

$$X_{\text{methyl alcohol}} = \frac{3}{2+3+5} = \frac{2}{10} = \boxed{0.3} \text{ Answer}$$

$$X_{\text{H2O}} = \frac{5}{2+3+5} = \frac{5}{10} = 0.5 \text{ Answer}$$

Now, multiply the mole fractions with 100, to get mole percent.

Moles % of ethyl alcohol = 0.2x100 = $\boxed{20}$ Answer

Moles % of methyl alcohol = 0.3x100 = $\boxed{30}$ Answer

Moles % of H_2O = 0.5x100 = $\boxed{50}$ Answer

In the case of mixture of gases, one can determine the mole fraction from the partial pressure data of the mixture. Hence

$$X_A = \frac{P_A}{P_A + P_B + P_C} \qquad ,X_B = \frac{P_B}{P_A + P_B + P_C} \qquad ,X_C = \frac{P_C}{P_A + P_B + P_C}$$

Where p_A , p_B , p_C are the partial pressures of various gases in the mixture. Generally, we can say that

$$\text{Mole fraction of any gas} = \frac{\text{Partial pressure of that gas}}{\text{Total pressure of the mixture of gases}}$$

4.2.5-Parts Per Million (ppm):

It can define as the number of parts (by weight or volume) of a solute per million parts (by weight or volume) of the solution. This unit is used for very low concentrations of solutions, e.g. to express the impurities of substances in water.

Mass of solute in Parts per million (ppm) = x 10^6 Mass of solution.

Example 4.5:

Seawater has 5.65 x 10^{-3} g of dissolved oxygen in one kg of water. Calculate the concentration of oxygen in sea water in parts per million.

Solution:

$$\text{ppm of oxygen in sea water} = \frac{5.65 \times 10^{-3} g}{1000 g} \times 10^6 = \boxed{5.65 \text{ ppm}} \text{ Answer}$$

4.3 TYPES OF SOLUTIONS

Most commonly, we come across solutions, where solute is a solid and the solvent is a liquid. All the three states of matter i.e. solid, liquid or gas can act as solute or solvent. Examples of nine possible types of solution are given in Table (4.1).

Common types and examples of solutions

State of Solute	State of Solvent	Example
Gas	Gas	Air
Gas	Liquid	O_2 in water, CO_2 in water.
Gas	Solid	H_2 adsorbed by palladium
Liquid	Gas	Mist, fog, clouds, liquid air pollutants.
Liquid	Liquid	Alcohol in water, milk, benzene in toluene.
Liquid	Solid	Mercury in silver, butter, cheese.
Solid	Liquid	Sugar in water, jellies, paints.
Solid	Gas	Dust particles in smoke.
Solid	Solid	Metal alloys pearls, opals, carbon in iron (steel).

4.3.1 Solutions of Solids in Liquids

When a solid encounter a suitable liquid, it dissolves forming a solution i.e. a homogeneous mixture. This process of dissolution can be explained in terms of attraction between the particles of a solute and that of a solvent. The molecules or ions in solids are arranged in such a regular pattern that the inter-molecular or inter-ionic forces are at a maximum. The process of dissolution is to overcome these forces of attraction holding together the solute molecules or ions in the crystal lattice, by the solute-solvent forces.

Dissolution of Molecular Solids:

In molecular crystals, the intermolecular forces of attraction are either dipole-dipole or London dispersion type. These forces are relatively weak and can easily be overcome. Hence, non-polar or less polar molecular crystals usually dissolve in non-polar solvents like benzene.

Dissolution in Ionic Solids:

In the crystal lattice, the inter-molecular or inter-ionic forces of attraction between high polar molecules or ions are quite strong; hence, the polar solids fail to dissolve in nonpolar solvents. These strong electrostatic forces cannot be overcome or shattered by the weak solute-solvent attractions. Take the case of sugar. Due to hydrogen bonding, it

has tightly bound molecules, so it will not be dissolved by solvents like kerosene oil, petrol, benzene, etc. It can dissolve readily in water; because water attracts sugar molecules almost in the same way, as the sugar molecules attracts one another.

The inter-ionic forces of attraction are very strong in ionic solids so, equally strong polar solvents are needed to dissolve them. Such solids cannot be dissolved by moderate polar solvents e.g. acetone. A moderately polar solvent fails to dissolve sodium chloride, which is an ionic solid. Thus the solubility principle is that " like dissolves like' .

4.3.2 Solutions of Liquids in Liquids
The solutions of liquids in liquids may be divided into three classes.

4.3.2.1 Completely Miscible Liquids:
Liquids like alcohol and water or alcohol and ether mix in all proportions. However, the properties of such solutions are not strictly additive. Generally, the volume decreases on mixing but in some cases it increases. These types of solutions can usually be separated by fractional distillation.

- Water and ethanol
- Ethanol and ether

Properties:

- These solutions do not strictly obey additive property.
- Generally volume decreases on mixing due to the development of strong solute-solvent forces but in some cases it increases.
- Heat may be absorbed or evolved during mixing.
- Most of the solutions of this type can separate by fractional distillation.

4.3.2.2 Partially Miscible Liquids:
A large number of liquids are known which dissolve into one another up to a limited extent. For example, ether $C_2H_5OC_2H_5$ dissolves water to the extent of about 1.2 % and water dissolves ether up to the extent of about 6.5%. As the mutual solubilities are limited, the liquids are only partially miscible. On shaking equal volumes of water and ether, two layers are formed. Each liquid layer is a saturated solution of the other liquid. Such solutions are called conjugate solutions. The mutual solubility of these conjugate solutions is affected by temperature changes. Typical examples of such systems are:

Phenol-water system

- Triethylamine-water system
- Nicotine-water system

Consulate Temperature:

The temperature at which two conjugate solutions merge into one another is called critical temperature or upper consulate temperature.

Example

Water-aniline system has a single layer at 167.00 C with 15% water.

Methanol-cyclohexane system has consulate temperature of 49.1^0C with 29% methanol.

4.3.2.3.Liquids Practically Immiscible
Those liquids, which do not dissolve into each other in any proportion, are immiscible.

Examples:

Water and benzene ($H_2O + C_6H_6$)

Water and carbon disulphide ($H_2O + CS_2$)

4.4 SOLUBILITY:

"The concentration of the solute in the solution when it is in equilibrium with the solid substance at a particular temperature."

   OR

''The number of grams of solute, which can be dissolved in 100g of solvent at the given temperature in the presence of excess of un dissolved solids, called solubility''.

4.4.1 UNITS:
The solubility is expressed in various units:

- Number of g of solute dm^3 solution
- $Moldm^3$

4.4.2 Solubility Curves
Temperature has a marked effect on the solubility of many substances. A graphical representation between temperature and solubility of solution is called solubility curves. There are two types of solubility curves.

(a) Continuous solubility curves

(b) Discontinuous solubility curves

4.4.2.1 Continuous Solubility Curves
Continuous solubility curves don't show sharp breaks anywhere. According to Fig. (4.2), $KClO_3$, $K_2Cr_2O_7$, Pb $(NO_3)_2$ and $CaCl_2$ show continuous solubility curves. The solubility curves of KCl, NaCl and $NaNO_3$ give the straight lines. NaCl shows a very small change

of solubility from 0°C to 100°C increase in temperature. $Ce_2(SO_4)_3$ shows exceptional behavior whose solubility decreases with the increase in temperature and becomes constant from 40°C onwards. Anyhow, it shows a continuous solubility curve.
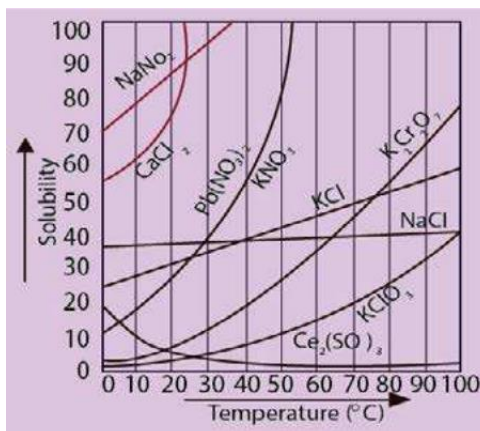
Fig. (4.2). Continuous solubility curves

4.4.2.2 Discontinuous Solubility Curves:

The solubility curves show sudden changes of solubilities, and these curves are called discontinuous solubility curves. The best examples in this reference are $Na_2SO_4.10 H_2O$, $CaCl_2.6 H_2O$. These curves are the combination of two or more solubility curves. At the break a new solid phase appears, and another solubility curve of that new phase begins. It is the number of molecules of water crystallization that changes and hence solubility changes, Fig (4.3).
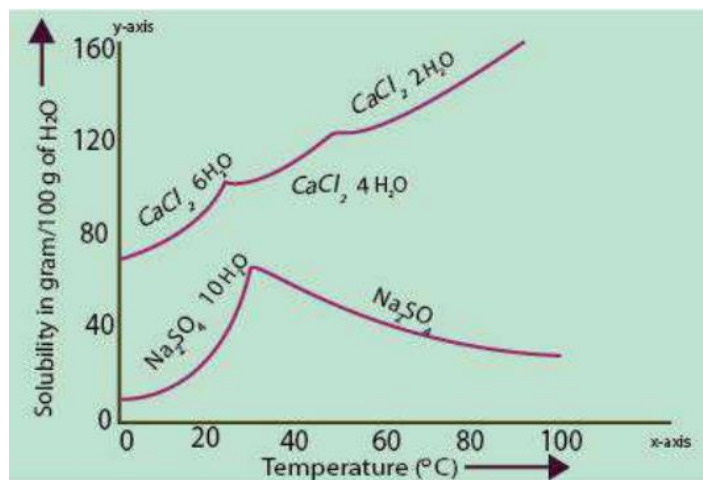
Fig (4.3). Discontinuous solubility curves

4.5 COLLIGATIVE PROPERTIES OF SOLUTIONS

The colligative properties are the properties of solution that depend on the number of solute and solvent molecules or ions. Following are colligative properties of dilute solution.

    I.   Lowering of vapour pressure
    II.   Elevation of boiling point
   III. Depression of freezing point
   IV. Osmotic pressure

The study of colligative properties has provided us with methods of molecular mass determination and has also contributed to the development of solution theory.

4.5.1 Why Some of the Properties are Called Colligative?

It can be explained by considering three different solutions. Let us take 6 g of urea, 18 g of glucose and 34.2 g of sucrose and dissolve them separately in 1 kg of H_2O. This will produce 0.1 molal solution of each substance.

Variations in vapour pressure:

Pure H_2O has a certain value of vapour pressure at a given temperature. In these three solutions, the vapour pressures will be lowered. The reason is that the molecules of a solute present upon the surface of a solution decrease the evaporating capability. It seems that sucrose solution should show the maximum lowering of vapour pressure while urea should have the minimum lowering of vapor pressure. The reality is that the lowering of vapor pressure in all these solutions will be same at a given temperature. The number of particles of the solute in all the solutions is equal. We have added 1/10th of Avogadro's number of particles (6.02×10^{22}). The lowering of vapor pressure depends upon the number of solute particles and not upon their molar mass and structures. Well, it should be kept in mind that these three solutes are non-volatile and non-electrolyte.

Examples of boiling point elevation and freezing point depression:

The boiling points of these solutions are higher than that of pure solvent. It is observed that the boiling point elevation of these three solutions is 0.052 °C. Similarly, freezing points will be depressed for these solutions and the value of depression in these three cases is 0.186°C. The reason again is that the elevation of boiling point and the depression of freezing point depend upon number of particles of solute.

Ebulliscopic constant (K_b) And Cryoscopic constant (K_f)

If we dissolve 60 g of urea, 180 g of glucose and 342 g of sugar separately in 1 kg of water. It is possible that the elevation of boiling point and depression of freezing point of water will be 0.52 °C and 1.86 °C, respectively. All the three solutions will boil at 100.52 °C and freeze at -1.86 °C . These values of elevation of boiling point and depression of freezing point are called molal boiling point constants and molal freezing point constants of H_2O denoted by K_b and K_f respectively. These are also named ebullioscopic and cryoscopic constants, respectively. These constants depend upon the nature of solvent and not upon the nature of solute. Following Table give the values of K_b and K_f for some common solvents.

Kb and Kf values for some solvents

Solvent	B.P.(^{0}C)	K_b(^{0}C/m)	F.P.(^{0}C)	K_f(^{0}C/m)
H_2O	100	0.52	0	1.86
Ether	34.4	2.16	-116.3	1.79
Aceticacid	118	3.07	17	3.90
Ethanol	79	1.75	-114.5	1.99
Benzene	80	2.70	5.5	5.10

To observe the colligative properties, the following condition should be fulfilled by the solutions.

- Solution should be dilute.
- Solute should be non-volatile.
- Solute should be non-electrolyte.

Now, let us discuss these colligative properties one by one.

4.5.2 Lowering of Vapor Pressure

The particles can escape from all over the surface of a pure solvent Fig. (4.4a). When the solvent is containing dissolved non-volatile non-electrolyte solute particles, the escaping tendency of solvent particles from the surface of the solution decreases and its vapor pressure is lowered Fig (4.4b)

(a)   (b)

● Solvent particle
● Solute particle
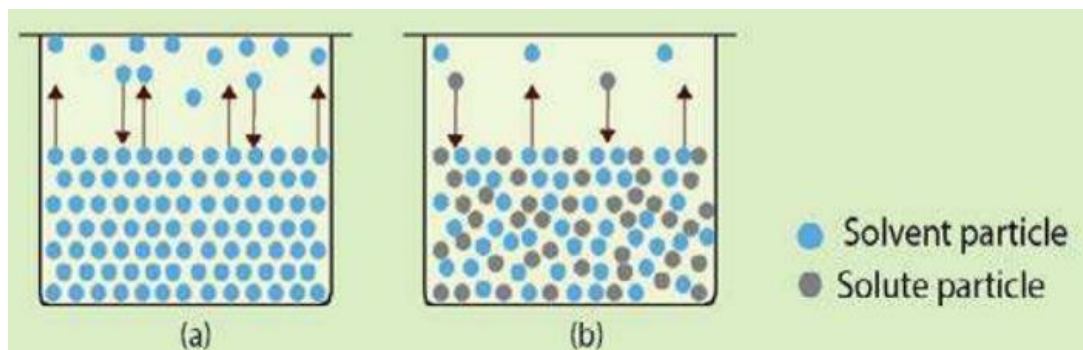
Fig (4.4) Lowering of vapour pressure

A quantitative relationship between the change of vapour pressure of a solvent due to addition of non-volatile and non-electrolyte solute and the mole fraction of solute has been given by Raoult. According to equation (3), Raoult says that relative lowering of vapour pressure is equal to the mole fraction of solute.

$$\frac{\Delta p}{p^o} = x_2$$

If n_2 and n_1 are the number of moles of a solute and solvent respectively , then.

$$x_2 = \frac{n_2}{n_1 + n_2}$$

$$\frac{\Delta p}{p^o} = \frac{n_2}{n_1 + n_2}$$

For a dilute solution, n2 can be ignored in the denominator.

$$\frac{\Delta p}{p^o} = \frac{n_2}{n_1}$$

The number of moles of solute and solvent are obtained by dividing their masses in grams with their respective relative molecular masses. If W1 and W2 are the masses of solvent and solute while M1 and M2 are their relative molecular masses receptively, then.

$$n_1 = \frac{W_1}{M_1} \qquad \text{and} \qquad n_2 = \frac{W_2}{M_2}$$

$$\frac{\Delta p}{p^o} = \frac{\dfrac{W_2}{M_2}}{\dfrac{W_1}{M_1}}$$

$$\frac{\Delta p}{p^{\circ}} = \frac{W_2}{M_2} \times \frac{M_1}{W_1} \qquad \cdots\cdots\cdots \quad (5)$$

$$M_2 = \frac{p^{\circ}}{\Delta p} \times \frac{W_2 M_1}{W_1} \qquad \cdots\cdots\cdots \quad (6)$$

The molecular mass (M) of a non-volatile solute can be calculated from the equation (6).

Example 4.10:

Pure benzene has a vapour pressure of 122.0 torr at 32°C. When 20g of a non- volatile solute were dissolved in 300g of benzene, a vapour pressure of 120 torr was observed. Calculate the molecular mass of the solute. The molecular mass of benzene being 78.1.

Solution:

Let the molecular mass of the solute be = M_2

Mass of solute dissolved (W_2) = 20 g

Vapor pressure of pure solvent (p°) = 122.0 torr

Vapor pressure of solution (p) = 120.0 torr

Lowering of vapor pressure (Δp) = 122.0 - 120.0 = 2.0 torr

Mass of solvent (W_1) = 300 g Molar mass of solvent (M_1) = 78.1

Formula applied
$$\frac{\Delta p}{p^{\circ}} = \frac{W_2}{M_2} \times \frac{M_1}{W_1}$$

$$M_2 = \frac{p^{\circ}}{\Delta p} \times \frac{W_2 M_1}{W_1}$$

Putting the values

$$M_2 = \frac{122.0}{2.0} \times \frac{20 \times 78.1}{300} = \boxed{317.6 \text{ g mol}^{-1}} \text{ Answer}$$

4.5.3 Elevation of Boiling Point
The presence of a non-volatile non-electrolyte solute in the solution decreases the vapour pressure of the solvent. The greater, the concentration of solute, greater will be

the lowering of vapour pressure. Therefore, the temperature at which a solvent in the solution state boils is increased.

To understand it, determine the vapour pressures of a solvent at various temperatures. Plot a graph between temperatures on x-axis and vapour pressures on y-axis. A rising curve is obtained with the increase of temperature. The slope of the curve at high temperature is greater, which shows that at high temperature the vapour pressure increases more rapidly.
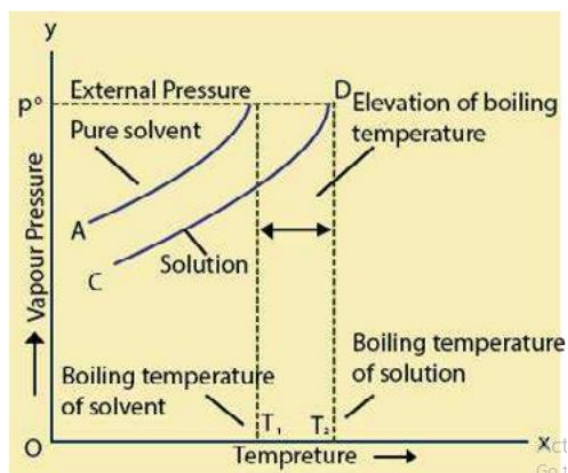
Fig 4.5. Elevation of boiling temperature curve

Temperature T1 on the curve AB which is for the pure solvent, corresponds to the boiling point of the solvent. The solvent boils when its vapour pressure becomes equal to the external pressure represented by p°. When the solute is added in the solvent and vapour pressures are plotted vs temperatures, then a curve CD is obtained. This curve is lower than the curve AB because vapour pressures of solution are less than those of pure solvent. Solution will boil at higher temperature T2 to equalize its pressure to p°. The difference of two boiling points gives the elevation of the boiling point ΔTb . The higher the concentration of solute, the greater will be the lowering in vapour pressure of solution and hence higher will be its boiling point. So, elevation of boiling point ΔTb is directly proportional to the molality of solution.

$$\Delta T_b = K_b m \qquad \text{...............} \quad (8)$$

Where Kb is called the ebullioscopic constant or molal boiling point constant.

According to equation (8), molality of any solute determines the elevation of boiling point of a solvent. You may dissolve 6 g of urea in 500 g of H_2O or 18 g of glucose in 500 g of H2O both give 0.2 molal solution and both have same elevation of boiling points i.e. 0.1 °C, which is l/5th of 0.52°C. We say that ΔTb (not T) is a colligative property.

We know that

$$\text{Molality(m)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Mass of solvent in kg}}$$

or

$$m = \frac{W_2}{M_2} \frac{1}{W_1/1000} = \frac{1000\,W_2}{M_2 W_1} \qquad \text{.........} \quad (9)$$

Putting the value of m from equation (9) into equation (8)

$$\Delta T_b = K_b \frac{1000\,W_2}{M_2 W_1} \qquad \text{.........} \quad (10)$$

Rearranging equation (10)

$$\text{Molecular mass (M}_2) = \frac{K_b}{\Delta T_b} \times \frac{W_2}{W_1} \times 1000 \qquad \text{............} \quad (11)$$

Equation (11) can be used to determine the molar mass of a non-volatile and nonelectrolyte solute in a volatile solvent.

4.5.4 Depression of the Freezing Point of a Solvent by a Solute

The freezing point of a substance is the temperature at which the solid and liquid phases of the substance co-exist.

Freezing point is also defined as the temperature at which its solid and liquid phases have the same vapor pressures. When a non-volatile solute is added to a solvent, its vapor pressure is decreased. At the freezing point, there are two things in the vessel i.e. liquid solution and the solid solvent. The solution will freeze at that temperature at which the vapor pressures of both liquid solution and solid solvent are same. It means that a solution should freeze at lower temperature than pure solvent.

To understand it, plot a graph between vapour pressure temperature for pure solvent and that of solution. The curve ABC is for the pure solvent. The solvent freezes at temperature T1 corresponding to point B when the vapour pressure of freezing solvent is p°. The portion of the curve BC is for the solid solvent. This portion has a greater slop, showing that the change of vapour pressure with the change of temperature is more rapid Fig (4.6).
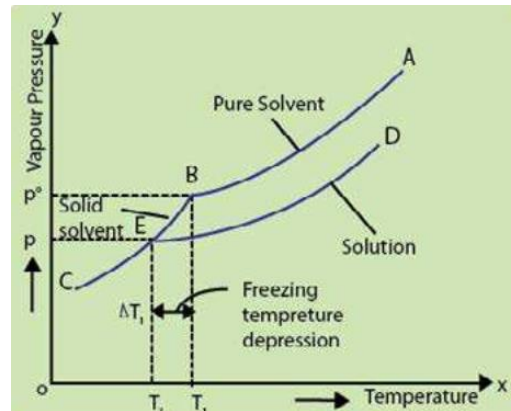
Fig 4.6 Depression of freezing point curve

The curve DEC for the solution will meet the curve BC at point E. This is the freezing point of solution T$_2$, and corresponds to the vapour pressure p which is lower than p°. The reason is that vapour pressure of solution is less than the pure solvent.

Depression of freezing point = freezing point of pure solvent - freezing point of solution.

$$\Delta T_f = T_1 - T_2$$

This depression in freezing point ΔTf is related to the molality (m) of the solution. The relationship is like that of elevation of the boiling point.

$$\Delta T_f \propto m$$

$$\Delta T_f = K_f m \qquad \text{............} \quad (12)$$

K$_f$ is called the molal freezing point constant or the cryoscopic constant and m is the molality of the solution. To get the inal expression, let us put the following expression (9) of molality into the equation (12) i.e.

$$m = \frac{1000 \; W_2}{M_2 x W_1} \qquad \text{.........} \quad (9)$$

We get $\qquad \Delta T_f = K_f \dfrac{1000 \; W_2}{M_2 x W_1} \qquad \text{...............} \quad (13)$

Where W2 = mass of solute and M2 = molar mass of the solute, W1 = mass of solvent in kg Rearranging equation (13)

$$\text{Molar mass of solute } (M_2) = \frac{K_f}{\Delta T_f} \frac{1000 \, W_2}{W_1} \qquad \text{.........} \quad (14)$$

4.5.5 Applications of Boiling Point Elevation and Freezing Point Depression Phenomena:

i.   It can be applied for molecular mass determination of a substance (Solute).
ii. The presence of a solute increases the liquid range of the solution both by raising the boiling point and lowering the freezing point.
iii. The most important application of this phenomenon is the use of antifreeze in the radiator of an automobile. The solute is ethylene glycol, which is not only completely miscible with water but has a very low vapor pressure and is non-volatile. When mixed with water, it lowers the freezing point as well as raises the boiling point. During winter it protects a car by preventing the liquid in the radiator from freezing, as water alone, if it were used instead. In hot summer, the antifreeze solution also protects the radiator from boiling over.
iv. Another common application is the use of NaCl or KNO_3 to lower the melting point of ice. One can prepare a freezing mixture for use in an ice cream machine.

EXERCISE

1. Define and explain the followings with one example in each case.
   (a) A homogeneous phase
   (b) A concentrated solution
   (c) A solution of solid in a solid
   (d) A consulate temperature
   (e) Conjugate solution
2. What are the concentration units of solutions. Compare molar and molal solutions.
3. One has one molal solution of NaCl and one molal solution of glucose.
   (a) Which solution has a greater number of particles of solute?
   (b) Which solution has a greater amount of the solvent?
   (c) How do we convert these concentrations into weight-by-weight percentages?
4. Explain the following with reasons.
   i.   The concentration in terms of molality is independent of temperature but molarity depends upon temperature.
   ii. The sum of mole fractions of all the components is always equal to unity for any solution.
   iii. 100 g of 98 % H_2SO_4 has a volume of 54.34 cm3 of H_2SO_4 .(Density = 1.84 g cm-3) iv) Relative lowering of vapour pressure is independent of the temperature.
   iv. Colligative properties are obeyed when the solute is non-electrolyte, and when the solutions are dilute.
5. Give graphical explanation for elevation of boiling point of a solution.
6. Freezing points of solutions are depressed when non-volatile solutes are present in volatile solvents. Justify it. Plot a graph to elaborate your answer.
7. Calculate the molarity of glucose solution when 9 g of it is dissolved in 250 cm3 of solution. (Ans: 0.2 mol dm^{-3})
8. Calculate the mass of urea in 100 g of H_2O in 0.3 molal solution. (Ans: 1.8g)
9. Calculate the concentration of a solution in terms of molality, which is obtained by mixing 250 g of 20% solution of NaCl with 200 g of 40 % solution of NaCl. (Ans: 6.94m)
10. An aqueous solution of sucrose has been labeled as 1 molal. Find the mole fraction of the solute and the solvent. (Ans: 0.0176, 0.9823)
11. You are provided with 80% H_2 SO_4 w/w having a density 1.8 g cm-^{3} . How much volume of this H_2SO_4 sample is required to obtain one dm3 of 20% w/w H_2SO_4 , which has a density of 1.25 g cm^{-3} . (Ans: 173.5cm^3)

Chapter # 5

Acid, Base and Salt

STUDENTS LEARNING OUTCOMES:

- Define and give examples of Arrhenius acids and bases.
- Use the Bronsted-Lowry theory to classify substances as acids or bases, as proton donors or proton acceptors.
- Classify substances as Lewis acids or bases.
- Complete and balance a neutralization reaction.

Acids, bases and salts are three distinct classes in which almost all the organic and inorganic compounds are classified. A famous Muslim Chemist Jabir Bin Hayan prepared nitric acid (HNO_3), hydrochloric acid (HCl) and sulphuric acid (H_2SO_4).

In 1787, Lavoisier named binary compounds of oxygen such as CO_2 and SO_2 as acids which on dissolution in water gave acidic solutions.

Later in 1815, Sir Humphrey Davy discovered that there are certain acids which are without oxygen, e.g., HCl. Davy proved the presence of hydrogen as the main constituent of all acids. It was also discovered that all water-soluble metallic oxides turn red litmus blue, which is a characteristic of bases. The word acid is derived from the Latin word 'Acidus' meaning sour. The first acid known to man was acetic acid, i.e., in the form of vinegar.

We all have a little concentration of hydrochloric acid in our stomach, which helps to break down the food. Sometimes, the amount of stomach acid becomes too much, which causes 'acidity'. This uncomfortable feeling is easily treated by taking an alkaline medicine. The alkali neutralizes the acid, producing a harmless chemical called a salt.

5.1:CONCEPTS OF ACIDS AND BASES

Table 5.1 Acids and bases are recognized by their characteristic properties, such as:

Acids	Bases
1. Acids have sour taste. For example, unripe citrus fruits or lemon juice.	1. Bases have bitter taste and feel slippery, for example, soap is slippery to touch.
2. They turn blue litmus red.	2. They turn red litmus blue.
3. They are corrosive in concentrated form.	3. They are non-corrosive except concentrated forms of NaOH and KOH.
4. Their aqueous solutions conduct electric current	4. Their aqueous solutions conduct electric current.

5.2:Different concepts of acids and bases:

5.2.1 Arrhenius Concept of Acids and Bases

According to Arrhenius concept (1787):

"Acid is a substance which dissociates in aqueous solution to give hydrogen ions". In general, the ionization of acids takes place as follows.

$$HA_{(aq)} \underset{}{\overset{water}{\rightleftharpoons}} H^+_{(aq)} + A^-_{(aq)}$$
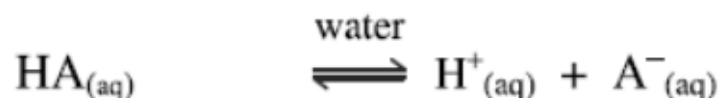
For example, substances such as HCl , HNO$_3$, CH$_3$COOH, etc., are acids because they ionize in aqueous solutions to provide H$^+$ ions.

$$HCl_{(aq)} \underset{}{\overset{water}{\rightleftharpoons}} H^+_{(aq)} + Cl^-_{(aq)}$$

$$HNO_{3(aq)} \underset{}{\overset{water}{\rightleftharpoons}} H^+_{(aq)} + NO^-_{3(aq)}$$

$$CH_3COOH_{(aq)} \overset{water}{\rightleftharpoons} CH_3COO^-_{(aq)} + H^+_{(aq)}$$
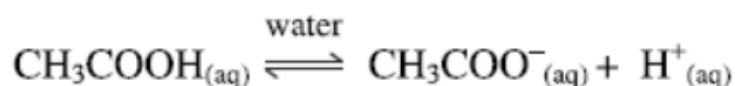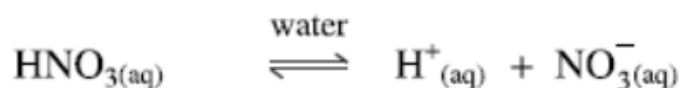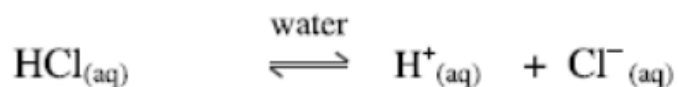
On the other hand, "a base is a substance that dissociates in an aqueous solution to give hydroxide ions". The general ionization of bases takes place as follows.

$$BOH_{(aq)} \overset{water}{\rightleftharpoons} B^+_{(aq)} + OH^-_{(aq)}$$
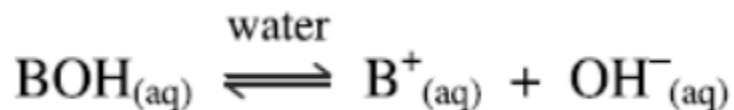
The substances such as NaOH, KOH, NH$_4$OH, Ca(OH)$_2$ etc. are bases because these compounds ionize in aqueous solutions to provide OH ions

$$NaOH_{(aq)} \overset{water}{\rightarrow} Na^+_{(aq)} + OH^-_{(aq)}$$
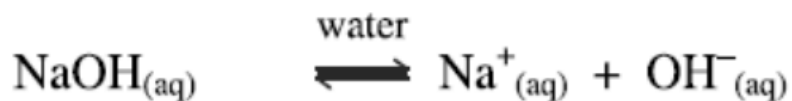
Thus, according to Arrhenius Concept:

Acids give H$^+$ ions in water, bases give OH ions in water.

Examples of some important acids and bases are given in Table 5.2.

Acids	Bases
Hydrochloric acid, HCl	Sodium hydroxide, NaOH
Nitric acid, HNO_3	Potassium hydroxide, KOH
Sulphuric acid, H_2SO_4	Calcium hydroxide, $Ca(OH)_2$
Phosphoric acid, H_3PO_4	Aluminium hydroxide, $Al(OH)_3$

5.2.1.1:Limitations of Arrhenius Concept:

1. This concept is applicable only in aqueous medium and does not explain the nature of acids and bases in non-aqueous medium.

2. According to this concept, acids and bases are only those compounds which contain hydrogen (H^+) and hydroxide (OH^-) ions, respectively. It can't explain the nature of compounds like CO_2, NH3, etc. which are acid and base, respectively.

Although this concept has limited scope yet, it led to the development of more general theories of acid-base behaviour.

5.2.2 Bronsted-Lowry Concept:

In 1923, the Danish chemist Bronsted and the English chemist Lowry independently presented their theories of acids and bases based on proton-transfer.

According to this concept:

"An acid is a substance (molecule or ion) that can donate a proton (H+) to another substance".

A base is a substance that can accept a proton (H+) from another substance. For example, HCl acts as an acid while NH3 acts as a base:

$$HCl_{(aq)} \quad + \quad NH_{3(aq)} \quad \rightleftharpoons \quad NH_4^+{}_{(aq)} \quad + \quad Cl^-{}_{(aq)}$$

Similarly, when HC1 dissolves in water; HC1 acts as an acid and H_2O as a base.

$$\underset{\text{Acid}}{HCl_{(aq)}} \quad + \quad \underset{\text{Base}}{H_2O_{(aq)}} \quad \rightleftharpoons \quad \underset{\text{Conjugate acid}}{H_3O^+{}_{(aq)}} \quad + \quad \underset{\text{Conjugate base}}{Cl^-{}_{(aq)}}$$

whereas H_2O is a base as it accepts a proton. In the reverse reaction, Cl⁻ ion is a base as it accepts a proton from acid H_3O^+ ion. Cl ion is called a conjugate base of acid HCl and H_3O^+ ion is called a conjugate acid of base H_2O. It means every acid produces a conjugate base and every base produces a conjugate acid such that there is a conjugate acid-base pair. Conjugate means joined together as a pair.

Conjugate acid is a species formed by accepting a proton by a base.

A conjugate base is a species formed by donating a proton by an acid.

Thus, conjugate acid-base pairs differ from one another only by a single proton. Similarly

$$CH_3COOH_{(aq)} + H_2O_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

Acid · · · · Base · · · · Conjugate base · · · · Conjugate acid

According to Bronsted-Lowry concept, an acid and a base always work together to transfer a proton. That means, a substance can act as an acid (proton donor) only when another substance simultaneously behaves as a base (proton acceptor). Hence, a substance can act as an acid as well as a base, depending upon the nature of the other substance. For example, H2O acts as a base when it reacts with HCl as stated above and as an acid when it reacts with ammonia such as:

$$H_2O_{(l)} + NH_{3(aq)} \rightleftharpoons NH_4^+_{(aq)} + OH^-_{(aq)}$$

Such a substance that can behave as an acid, as well as a base is called amphoteric.

5.2.2.1:Limitations of Bronsted-Lowry concept:
It has been observed that there are certain substances which behave as acids though they cannot donate a proton, e.g., SO₃. Similarly, CaO behaves as a base, but it cannot accept a proton. These observations prove the limitations of Bronsted-Lowry concept of acids and bases.

Conjugate acid-base pairs of common species

Acid		Base		Conjugate acid		Conjugate base
$HNO_{3(aq)}$	+	$H_2O_{(l)}$	$\rightleftharpoons$	$H_3O^+_{(aq)}$	+	$NO_3^-_{(aq)}$
$H_2SO_{4(aq)}$	+	$H_2O_{(l)}$	$\rightleftharpoons$	$H_3O^+_{(aq)}$	+	$HSO_4^-_{(aq)}$
$HCN_{(aq)}$	+	$H_2O_{(l)}$	$\rightleftharpoons$	$H_3O^+_{(aq)}$	+	$CN^-_{(aq)}$
$CH_3COOH_{(aq)}$	+	$H_2O_{(l)}$	$\rightleftharpoons$	$H_3O^+_{(aq)}$	+	$CH_3COO^-_{(aq)}$
$H_2O_{(l)}$	+	$NH_{3(aq)}$	$\rightleftharpoons$	$NH_4^+_{(aq)}$	+	$OH^-_{(aq)}$
$H_2O_{(l)}$	+	$CO_3^{2-}_{(aq)}$	$\rightleftharpoons$	$HCO_3^-_{(aq)}$	+	$OH^-_{(aq)}$
$HCl_{(l)}$	+	$HCO_3^-_{(aq)}$	$\rightleftharpoons$	$H_2CO_{3(aq)}$	+	$Cl^-_{(aq)}$

5.2.3 Lewis Concept of Acids and Bases
The Arrhenius and Bronsted-Lowry concepts of acids and bases are limited to substances which contain protons. G.N. Lewis (1923) proposed a more general and broader concept of acids and bases. According to this concept: An acid is a substance (molecule or ion) which can accept a pair of electrons, while a base is a substance
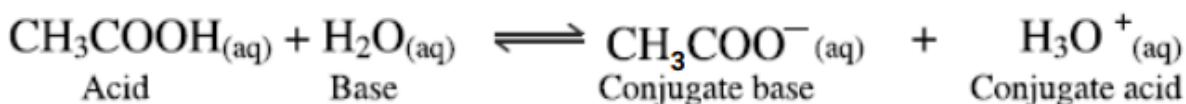
(molecule or ion) which can donate a pair of electrons. For example, a reaction between ammonia and boron trifluoride takes place by forming a coordinate covalent bond between ammonia and boron trifluoride by donating an electron pair of ammonia and accepting that electron pair by boron trifluoride.
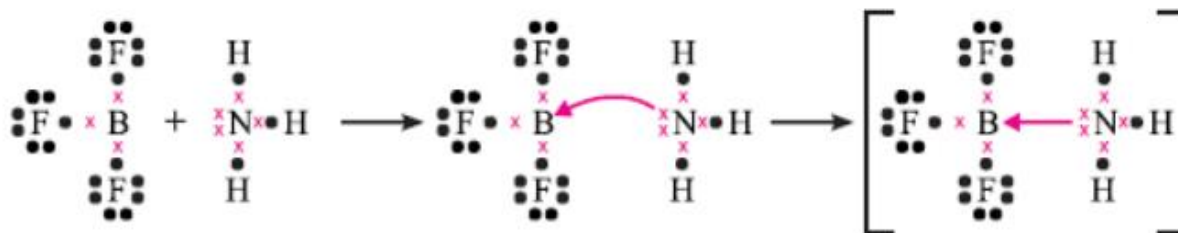
The cations (proton itself or metal ions) act as Lewis's acids. For example, a reaction between H^+ and NH_3, where H^+ acts as an acid and ammonia as a base.
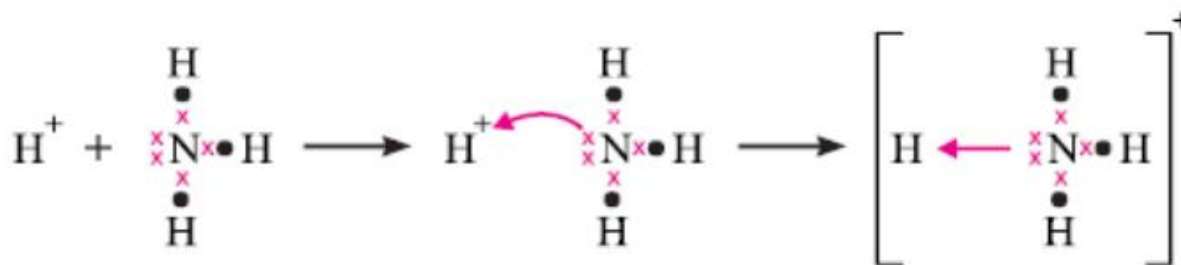
Adduct:

The product of any Lewis acid-base reaction is a single specie, called an adduct .So, a neutralization reaction according to Lewis concept is donation and acceptance of an electron pair to form a coordinate covalent bond in an adduct.

Acids are electron pair acceptors while bases are electron pair donors. Thus, it is evident that any substance which has an unshared pair of electrons can act as a Lewis base while a substance which has an empty orbital that can accommodate a pair of electrons acts as Lewis acid. Examples of Lewis acids and bases are given below:

5.2.3.1:Characteristics of Lewis acids:
According to Lewis concept, the following species can act as Lewis acids:

i. Molecules in which the central atom has incomplete octet. For example, in BF_3 , $AlCl_3$, $FeCl_3$, the central atoms have only six electrons around them, therefore, these can accept an electron pair.
ii. Simple cations can act as Lewis acids. All cations act as Lewis acids since they are deficient in electrons. However, cations such as Na^+, K^+, Ca^{2+} ions, etc., have

a very little tendency to accept electrons. While the cations like H^+, Ag^+ ions, etc., have a greater electron accepting tendency therefore, act as Lewis acids.

5.2.3.2:Characteristics of Lewis bases:

According to Lewis concept, the following species can act as Lewis bases:

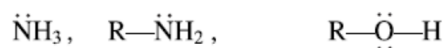(i) Neutral species having at least one lone pair of electrons. For example, ammonia, amines, alcohols etc. act as Lewis bases because they contain a lone pair of electrons:

$$\ddot{N}H_3, \quad R\!-\!\ddot{N}H_2, \qquad R\!-\!\ddot{O}\!-\!H$$

(ii) Negatively charged species or anions. For example, chloride, cyanide, hydroxide ions, etc., act as Lewis bases:

$$CN^-, \ Cl^-, \ OH^-, \ \text{etc.}$$

It may be noted that all Bronsted bases are also Lewis bases but all Bronsted acids are not Lewis acids. According to Bronsted concept, a base is a substance which can accept a proton, while according to Lewis concept, a base is a substance which can donate a pair of electrons. Lewis bases generally contain one or more lone pair of electrons and therefore, they can also accept a proton (Bronsted base). Thus, all Lewis bases are also Bronsted bases. On the other hand, Bronsted acids are those which can give a proton. For example, HCl, H2 SO4 are not capable of accepting a pair of electrons. Hence, all Bronsted acids are not Lewis acids.

Summary of the Concepts.

Concept	Acid	Base	Product
Arrhenius	give H^+	gives OH^-	salt + H_2O
Bronsted-Lowry	donate H^+	accepts H^+	conjugate acid base pair
Lewis	electron pair acceptor	electron pair donor	adduct

5.2.4 General Properties of Acids

5.2.4.1 Physical Properties

Physical properties of acids have been described in the beginning of the chapter.

5.2.4.2 Chemical Properties:

i.Reaction with Metals:

Acids react explosively with metals like sodium, potassium, and calcium. However, dilute acids (HCl, H2SO4) react moderately with reactive metals like: Mg, Zn, Fe and Al to form their respective salts with the evolution of hydrogen gas.

$$Zn_{(s)} \quad + \quad H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} \quad + \quad H_{2(g)} \uparrow$$

$$2Al_{(s)} \quad + \quad 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} \quad + \quad 3H_{2(g)} \uparrow$$

ii.Reaction with Carbonates and Bicarbonates

Acids react with carbonates and bicarbonates to form corresponding salts with the evolution of carbon dioxide gas.

$$CaCO_{3(aq)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + CO_{2(g)} \uparrow + H_2O_{(l)}$$
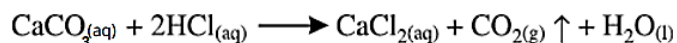
iii.Reaction with Bases

Acids react with bases (oxides and hydroxides of metal and ammonium hydroxide) to form salts and water. This process is called neutralization.

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

$$CuO_{(s)} + H_2SO_{4(aq)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)}$$
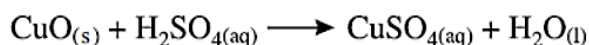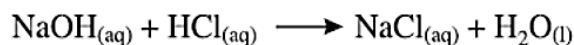
iv.Reaction with Sulphites and Bisulphites:

It react with sulphites and bisulphites to form salts with the liberation of sulphur dioxide gas.

$$CaSO_{3(aq)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + SO_{2(g)} \uparrow + H_2O_{(l)}$$

$$NaHSO_{3(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + SO_{2(g)} \uparrow + H_2O_{(l)}$$
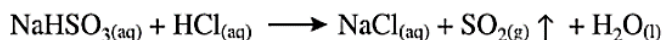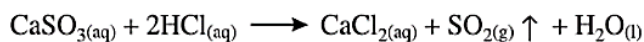
v.Reaction with Sulphides:

Acids react with metal sulphides to liberate hydrogen sulphide gas.

$$FeS_{(s)} + H_2SO_{4(aq)} \longrightarrow FeSO_{4(aq)} + H_2S_{(g)} \uparrow$$
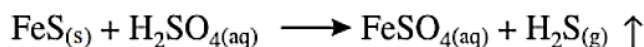
5.2.5 Uses of Acids

i. Sulphuric acid is used to manufacture fertilizers, ammonium sulphate, calcium superphosphate, explosives, paints, dyes, drugs. It is also used as an electrolyte in lead storage batteries.

ii. Nitric acid is used in manufacturing of fertilizer (ammonium nitrate), explosives, paints, drugs and etching designs on copper plates.

iii. Hydrochloric acid is used for cleaning metals, tanning and in printing industries.

iv. Benzoic acid is used for food preservation.

v. Acetic acid is used for flavoring food and food preservation. It is also used to cure the sting of wasps.

	Naturally Occurring Acids	
	Acid	**Source**
i	Citric acid	Citrus fruits i.e., lemon, oranges
ii	Lactic acid	sour milk
iii	Formic acid	Stings of bees and ants
iv	Butyric acid	Rancid butter
v	Tartaric acid	Tamarind, grapes, apples
vi	Malic acid	Apples
vii	Uric acid	Urine
viii	Stearic acid	Fats

5.2.6 General Properties of Bases

5.2.6.1 Chemical Properties:

i. Reaction with Acids

Bases react with acid to form salt and water. It is a neutralization reaction.

$$2KOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$$
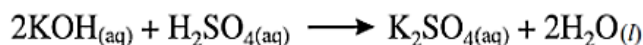
ii. Reaction with Ammonium Salts

Alkalis react with ammonium salts to liberate ammonia gas:

$$NH_4Cl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + NH_{3(g)}\uparrow + H_2O_{(l)}$$

$$(NH_4)_2SO_{4(aq)} + Ca(OH)_{2(aq)} \longrightarrow CaSO_{4(aq)} + 2NH_{3(g)}\uparrow + 2H_2O_{(l)}$$
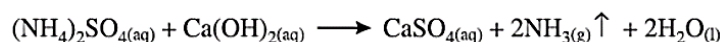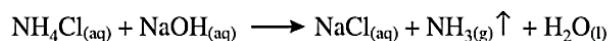
iii. Precipitation of Hydroxides

Alkalis precipitate insoluble hydroxides when added to solutions of salts of heavy metals such as copper, iron, zinc, lead and calcium.

$$FeCl_{3(aq)} + 3NaOH_{(aq)} \longrightarrow \underset{\text{Brown ppt.}}{Fe(OH)_{3\,(s)}} + 3NaCl_{(aq)}$$

$$Pb(NO_3)_{2(aq)} + 2NaOH_{(aq)} \longrightarrow \underset{\text{White ppt.}}{Pb(OH)_{2(s)}} + 2NaNO_{3(aq)}$$

$$CaCl_{2(aq)} + 2NaOH_{(aq)} \longrightarrow \underset{\text{White ppt.}}{Ca(OH)_{2(s)}} + 2NaCl_{(aq)}$$
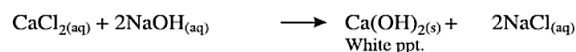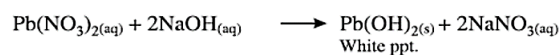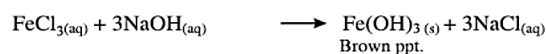
5.2.7 Uses of Bases:

1. Sodium hydroxide is used for manufacturing of soap.

2. Calcium hydroxide is used for manufacturing of bleaching powder, softening of hard water and neutralizing acidic soil and lakes due to acid rain.

3. Potassium hydroxide is used in alkaline batteries.

4. Magnesium hydroxide is used as a base to neutralize acidity in the stomach. It is also used for the treatment of bee's stings.

5. Aluminum hydroxide is used as foaming agent in fire extinguishers.

6. Ammonium hydroxide is used to remove grease stains from clothes.

5.3 SALTS

Salts are ionic compounds generally formed by the neutralization of an acid with a base. Salts are made up of positive ions (cations) and negative ions (anions). A cation is metallic ion derived from a base, therefore, it is called basic radical. While anion is derived from an acid, therefore, it is called acid radical. A salt gets its name from the names of the metal and the acid as shown in Table

Metal	Acid	Salt name
Sodium (Na)	Hydrochloric acid (HCl)	Sodium chloride (NaCl)
Potassium (K)	Nitric acid (HNO_3)	Potassium nitrate (KNO_3)
Zinic (Zn)	Sulphuric acid (H_2SO_4)	Zinc sulphate ($ZnSO_4$)
Calcium (Ca)	Phosphoric acid (H_3PO_4)	Calcium phosphate $Ca_3(PO_4)_2$
Silver (Ag)	Acetic acid (CH_3COOH)	Silver acetate (CH_3COOAg)

5.3.1 Characteristic properties of salts:
   i.   Salts are ionic compounds found in crystalline form.
   ii.   They have high melting and boiling points.
  iii.   Most of the salts contain water of crystallization which is responsible for the shape of the crystals. Number of molecules of water are specific for each salt and they are written with the chemical formula of a salt.For example, Copper sulphate $CuSO_4 .5H_2O$; Calcium sulphate $CaSO_4.2H_2O$
  iv.   Salts are neutral compounds. Although, they do not have equal number of positive and negative ions, but have equal number of positive and negative charges

5.3.2 Preparation:

Salts may be water soluble or insoluble. The methods used for the preparation of salts are based on their solubility in water. Salts may be water-soluble or insoluble. The methods used for the preparation of salts are based on their solubility in water.

5.3.2.1 General Methods for the Preparation of Salts

There are five general methods for the preparation of salts. Four methods make soluble salts but one prepares insoluble salts.

5.3.2.1.1:Preparation of soluble salts:

Soluble salts are often prepared in water. Therefore, they are recovered by evaporation or crystallization.

i.By the reaction of an acid and a metal: (Direct Displacement method)

This is direct displacement method in which hydrogen ion of acid is replaced by a reactive metal. Such as calcium, magnesium, zinc and iron, e.g.

$$\text{Acid} + \text{Metal} \longrightarrow \text{Salt} + \text{Hydrogen gas}$$

$$2HCl_{(aq)} + Mg_{(s)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$$
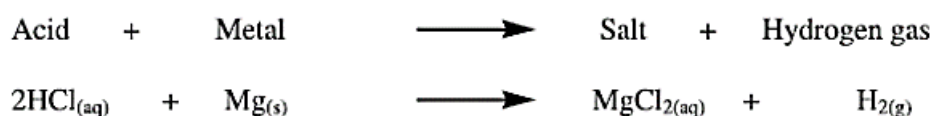
ii.By the reaction of an acid and a base: (Neutralization method)

It is a neutralization reaction in which acid and base react to produce a salt and water.

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$
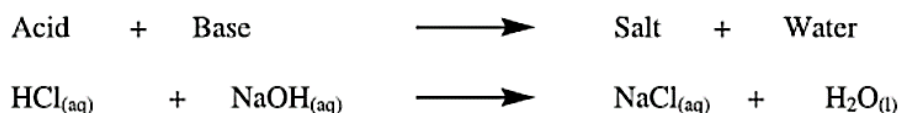
iii.By the reaction of an acid and metallic oxide:

Mostly the insoluble metallic oxides react with dilute acids to form salt and water

$$\text{Acid} + \text{Metallic oxide} \longrightarrow \text{Salt} + \text{Water}$$

$$H_2SO_{4(aq)} + CuO_{(s)} \longrightarrow CuSO_{4(aq)} + H_2O_{(aq)}$$
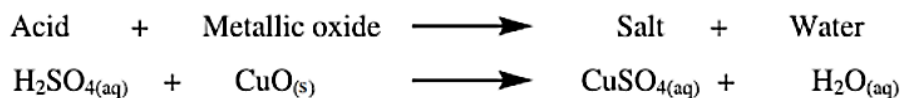
iv.By the reaction of an acid and a carbonate:a

Dilute acids react with metallic carbonates to produce salts, water and carbon dioxide gas.

$$2HNO_{3(aq)} + Na_2CO_{3(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}$$

5.3.2.1.2:Preparation of insoluble salts

In this method, usually solutions of soluble salts are mixed. During the reaction exchange of ionic radicals (i.e., metallic radicals exchange with acidic radicals) takes place to produce two new salts. One of the salts is insoluble and the other is soluble. The insoluble salt precipitates (solidify in solution).

$$AgNO_{3(aq)} + NaCl_{(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

$$Na_2CO_{3(aq)} + CuSO_{4(aq)} \longrightarrow CuCO_{3(s)} + Na_2SO_{4(aq)}$$

5.3.3 Types of Salts

Following are the main classes of salts.

i. Normal salts
ii. Acidic salts
iii. Basic salt
iv. Double salts
v. Mixed salts
vi. Complex salts

i.Normal salts

A salt formed by the total replacement of ionizable H+ ions of an acid by a positive metal ion or NH^{+}_{4} ions is called normal or neutral salt. These salts are neutral to litmus, that is,

$$HCl_{(aq)} + KOH_{(aq)} \longrightarrow KCl_{(aq)} + H_2O_{(l)}$$

$$H_2SO_{4(aq)} + ZnO_{(s)} \longrightarrow ZnSO_{4(aq)} + H_2O_{(l)}$$

$$H_3PO_{4(aq)} + 3NaOH_{(aq)} \longrightarrow Na_3PO_{4(aq)} + 3H_2O_{(l)}$$

$$HNO_{3(aq)} + NH_4OH_{(aq)} \longrightarrow NH_4NO_{3(aq)} + H_2O_{(l)}$$
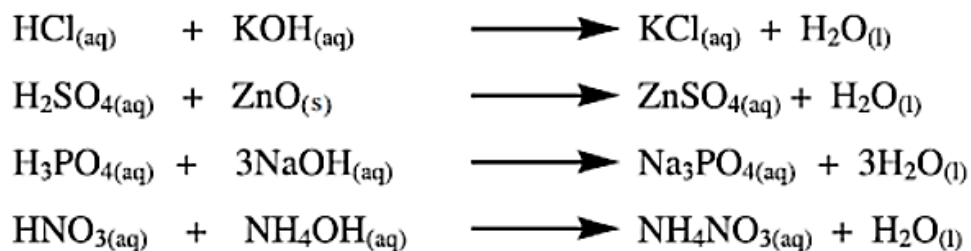
ii.Acidic Salts

These salts are formed by partial replacement of a replaceable H+ ions of an acid by a positive metal ion.

$$H_2SO_{4(aq)} + KOH_{(aq)} \longrightarrow KHSO_{4(aq)} + H_2O_{(l)}$$

$$H_3PO_{4(aq)} + NaOH_{(aq)} \longrightarrow NaH_2PO_{4(aq)} + H_2O_{(l)}$$
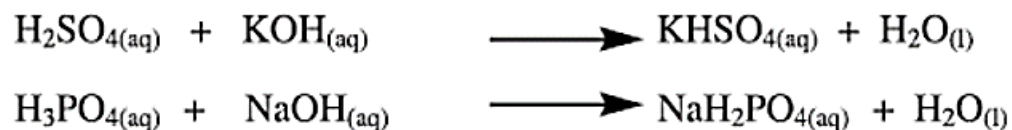
These salts turn blue litmus red.

Acidic salts react with bases to form normal salts.

$$KHSO_{4(aq)} + KOH_{(aq)} \longrightarrow K_2SO_{4(aq)} + H_2O_{(l)}$$

$$NaH_2PO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_3PO_{4(aq)} + 2H_2O_{(l)}$$

iii. Basic Salts

Basic salts are formed by the incomplete neutralization of a polyhydroxy base by an acid.

$$Al(OH)_{3(s)} + HCl_{(aq)} \longrightarrow Al(OH)_2Cl_{(aq)} + H_2O_{(l)}$$

$$Pb(OH)_{2(s)} + CH_3COOH_{(aq)} \longrightarrow Pb(OH)CH_3COO_{(aq)} + H_2O_{(l)}$$

$$Zn(OH)_{2(s)} + HNO_{3(aq)} \longrightarrow Zn(OH)NO_{3(aq)} + H_2O_{(l)}$$

These salts further react with acids to form normal salts.

$$Al(OH)_2Cl_{(aq)} + HCl_{(aq)} \longrightarrow Al(OH)Cl_{2(aq)} + H_2O_{(l)}$$

$$Al(OH)Cl_{2(aq)} + HCl_{(aq)} \longrightarrow AlCl_{3(aq)} + H_2O_{(l)}$$

$$Pb(OH)CH_3COO + CH_3COOH_{(aq)} \longrightarrow Pb(CH_3COO)_{2(aq)} + H_2O_{(l)}$$

$$Zn(OH)NO_{3(aq)} + HNO_{3(aq)} \longrightarrow Zn(NO_3)_{2(aq)} + H_2O_{(l)}$$
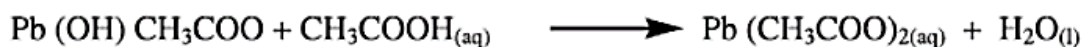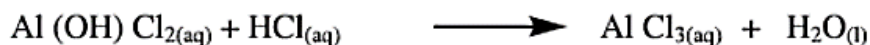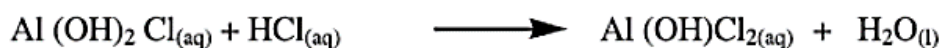
iv. Double Salts

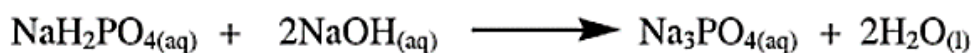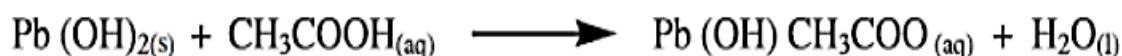Double salts are formed by two normal salts when they are crystallized from a mixture of equimolar saturated solutions. The individual salt components retain their properties. The anions and cations give their respective tests. Mohr's salt FeSO4 (NH4)2 SO4 6H2 O; Potash alum K2SO4. Al2(SO4)3. 24H2O; Ferric alum K2SO4 . Fe2(SO4)3. 24H2O, are examples of double salts.

v. Mixed Salts

Mixed salts contain more than one basic or acid radicals. Bleaching powder Ca(OCl) Cl, is an example of mixed salts.

vi. Complex Salts

Complex salts on dissociation provides a simple cation and a complex anion or vice versa. Only the simple ions yields the characteristics test for cation or anion. For example: Potassium ferrocyanide K4 [Fe(CN)6) gives on ionization, a simple cation K+ and complex anion [Fe (CN)6]4

5.3.4 Uses of Salts

Salts have vast applications in industries and in our daily life. Some common salts and their uses are given in Table

Name of salts	Common and Industrial Uses
Sodium chloride (NaCl)	It is commonly used as a table salt and for cooking purposes, it is also used for de-icing roads in winter and for the manufacture of sodium metal, caustic soda, washing soda.
Sodium carbonate Na_2CO_3) Soda ash	It is used for the manufacture of glass, detergents, pulp and paper and other chemicals.
Sodium carbonate Na_2CO_3, $10H_2O$) Washing soda	It is used as cleaning agent for domestic and commercial purposes, for softening of water, in manufacture of chemicals like caustic soda (NaOH), borax, glass, soap and paper.
Sodium sulphate (Na_2SO_4)	It is used for the manufacture of glass, paper and detergents.
Sodium silicate (Na_2SiO_3)	It is used for the manufacture of detergents, cleaning agents and adhesives.
sodium chlorate ($NaClO_3$)	It is used for manufacture of explosives, plastics and other chemicals.
Sodium tetraborate ($Na_2B_4O_7$ $10H_2O$)	It is used for manufacture of heat resistance glass (pyrex), glazes and enamels, in leather industry for soaking and cleaning hides.
Calcium chloride ($CaCl_2$)	It is used for de-icing roads in winter, as a drying agent of chemical reagents and as freezing agent.
Calcium oxide (CaO) Quick lime	It is used as drying agent for gases and alcohol and in steel making, water treatment and other chemicals like slaked lime, bleaching powder, calcium carbide. For purification of sugar, a mixture of CaO and NaOH called soda lime is used to remove carbon dioxide and water vapours from air.
Calcium sulphate ($CaSO_4$. $2H_2O$)	Gypsum is used as fertilizer, to prepare plaster of Paris which is used for making statues, casts, etc.
Potassium Nitrate (KNO_3)	It is used as fertilizer and for the manufacture of flint glass.

EXERCISE

Short Questions:

i.   Define a base and explain that all alkalies are bases, but all bases are not alkalies.

ii.   Define Bronsted-Lowry base and explain with an example that water is a Bronsted-Lowry base.

iii.   How can you justify that Bronsted-Lowry concept of acid and base is applicable to non- aqueous solutions?

iv.   Which kind of bond is formed between Lewis acid and a base?

v.   Why H+ ion acts as a Lewis acid?

vi.   Name two acids used in the manufacture of fertilizers.

vii.   Define the followings:
   a) Normal salt
   b) Basic salt

viii.   Na2SO4 is a neutral salt while NaHSO4 an acid salt. Justify.

ix.   Give a few characteristic properties of salts.

x.   How are the soluble salts recovered from water?

xi.   How are the insoluble salts prepared?

xii.   Why is a salt is neutral, explain with an example?

xiii.   Name an acid used in the preservation of food.

xiv.   Name the acids present in: i. Vinegar ii. Ant sting iii. Citrus fruit iv. Sour milk

xv.   How can you justify that Pb(OH)NO3 is a basic salt?

xvi.   You are in a need of an acidic salt. How can you prepare it?

xvii.   Which salt is used to prepare plaster of Paris?

Q.2: Extensive Question

1. Define an acid and a base according to Bronsted-Lowry concept and justify with examples that water is an amphoteric compound.
2. Explain the Lewis concept of acids and bases.
3. Define a salt and give the characteristic properties of salts.
4. Explain with examples how are soluble salts prepared?
5. Give the characteristics of an acidic salt.
6. Give four uses of calcium oxide.
7. You are having a strong acid (HNO3) and strong base (NaOH) on mixing
   a) What type of salt you will have?
   b) What type of reaction will it be?
   c) Will it be soluble or insoluble salt?
   d) If it is soluble, how will it be recovered?
8. Explain why:

a) HCl forms only one series of salts.
b) H_2SO_4 forms two series of salts.
c) H_3PO_4 form three series of salts. Give necessary equations.

9. Classify the following salts as soluble or insoluble salts:
   a) Sodium chloride
   b) Silver nitrate
   c) Lead chloride
   d) Copper sulphate
   e) Barium sulphate
   f) Ammonium chloride
   g) Sodium carbonate
   h) Calcium carbonate
   i) Ferric chloride
   j) Magnesium sulphate

10. Complete and balance the following equations:

   i.   Aluminium + Hydrochloride acid $\longrightarrow$
   ii. Copper oxide + Sulphuric acid $\longrightarrow$
   iii. Iron sulphide + Sulphuric acid $\longrightarrow$
   iv. Ammonium chloride + Sodium hydroxide $\longrightarrow$
   v.   Ferric chloride + Sodium hydroxide $\longrightarrow$

Chapter # 6

pH Scale and buffer system

STUDENTS LEARNING OUTCOMES:

- Analyze pH scale
- Write the equation for the self-ionization of water
- Given the hydrogen or hydroxide ion concentration, classify a solution as neutral, acidic, or basic.
- Calculate Acidic and Basic strength
- Distinguish between different types of indicators.

6.1 pH SCALE:

Concentration of hydrogen ion [H⁺] in pure water is the basis for the pH scale.

Water is a weak electrolyte because it ionizes very slightly into ions in a process called auto ionization or self-ionization.

$$H_2O \rightleftharpoons H^+ + OH^-$$

The equilibrium expression of this reaction may be written as

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

As concentration of water (H_2O) is almost constant. The above equation may be written as

$$K_c[H_2O] = [H^+][OH^-]$$

A new equilibrium constant known as ionic product constant of water 'Kw' is used instead of product of equilibrium constant and [H₂O]. Therefore,

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \quad at \quad 25°C$$

As we know, one molecule of water produces one H⁺ ion and one OH⁻ ion on dissociation so,

$$[H^+] = [OH^-] \quad Or \quad [H^+]^2 = 1.0 \times 10^{-14}$$
$$[H^+] = \sqrt{1.0 \times 10^{-14}}$$
$$[H^+] = 1.0 \times 10^{-7}\,M \quad at \quad 25°C$$

As it is difficult to deal with such small figures having negative exponents, it is convenient to convert these figures into a positive figure using a numerical system. It is taking the common (base-10) logarithm of the figure and multiplying it with -1. 'p' before a symbol means' negative logarithm of the symbol. So 'p' before H means negative logarithm of [H$^+$]. Therefore, pH is the negative logarithm of molar concentration of the hydrogen ions. That is,

$$pH = -\log [H^+]$$

With reference to this equation, a scale develops according to the molar concentration of H$^+$ ions that is called pH scale.

It ranges from 0 to 14. According to this scale, pH of water is calculated as:

$$pH = -\log [H^+]$$
$$pH = -\log (1.0 \times 10^{-7}) = 7$$

Similarly

$$pOH = -\log [OH^-]$$
$$pOH = -\log (1.0 \times 10^{-7}) = 7$$

pH value normally varies from 0 to 14. Therefore:

$$pH + pOH = 14$$

So, the sum of the pH and pOH of the solution is always 14 at 25 °C. Such as;

	Highly acidic			Slightly acidic		neutral		Slightly basic				Highly basic		
pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
pOH	14	13	12	11	10	9	8	7	6	5	4	3	2	1	0

A solution of a compound of pH 7 or pOH 7 is considered a neutral solution. Solutions of pH less than 7 are acidic and more than 7 are basic as are also shown in figure

	$[H_3O+]$	pH	$[OH^-]$	pOH
BASIC	1×10^{-14}	14.0	1×10^{-0}	0.0
	1×10^{-13}	13.0	1×10^{-1}	1.0
	1×10^{-12}	12.0	1×10^{-2}	2.0
	1×10^{-11}	11.0	1×10^{-3}	3.0
	1×10^{-10}	10.0	1×10^{-4}	4.0
	1×10^{-9}	9.0	1×10^{-5}	5.0
	1×10^{-8}	8.0	1×10^{-6}	6.0
NATURAL	1×10^{-7}	7.0	1×10^{-7}	7.0
ACIDIC	1×10^{-6}	6.0	1×10^{-8}	8.0
	1×10^{-5}	5.0	1×10^{-9}	9.0
	1×10^{-4}	4.0	1×10^{-10}	10.0
	1×10^{-3}	3.0	1×10^{-11}	11.0
	1×10^{-2}	2.0	1×10^{-12}	12.0
	1×10^{-1}	1.0	1×10^{-13}	13.0
	1×10^{-0}	0.0	1×10^{-14}	14.0

A solution of a compound of pH 7 or pOH 7 is considered a neutral solution. Solutions of pH less than 7 are acidic and more than 7 are basic as are also shown in figure.

Since the pH scale is logarithmic, a solution of pH 1 has 10 times higher concentration of [H+] than that of a solution of pH 2; 100 times than that of a solution of pH 3 and so on. Hence, low pH value means strong acid while high pH value means a strong base and vice versa.

Conclusion:

pH of a neutral solution is always 7.

Acidic solutions have pH less than 7.

Basic solutions have pH value greater than 7.

pH and pOH values range from 0 to 14.

6.1.1:Uses of pH:

i. It is used to determine acidic or basic nature of a solution.
ii. It is used to produce medicines, culture at a microbiological particular concentration of H^+ ion.
iii. It is used to prepare solutions of required concentrations necessary for certain biological reactions.

6.2 Indicators:

Indicators are organic compounds. They have different colours in acidic and alkaline solutions. Litmus is a common indicator. It is red in acidic solutions and blue in alkaline solutions.

Each indicator has a specific colour in acidic medium which changes at a specific pH to another colour in basic medium. For example, phenolphthalein is colourless in strongly acidic solution and red in strongly alkaline solution. It changes colour at a pH of about 9. This means phenolphthalein is colourless in a solution with pH less than 9. If the pH is above 9, phenolphthalein is red as is shown in the figure.

A few commonly used indicators in titrations are given in Table 10.3

Table 10.4 Few important indicators

Indicator	Colour in strongly acidic solution	pH a which colour changes	Colour in strongly alkaline solution
Methyl orange	red	4	Yellow
Litmus	red	7	blue
Phenolphthalein	colourless	9	red

6.2.1 Measuring pH of a Solution

There are several methods for measuring pH, each with its own advantages and limitations. Here are some common methods:

i.pH Indicator Paper:

pH indicator paper or pH test strips are one of the simplest and most affordable methods for measuring pH. These strips are impregnated with pH-sensitive compounds that

change color in response to the pH of the solution. Users compare the color change with a provided color chart to determine the pH of the solution. Suitable for quick and qualitative pH measurements but may not provide high precision.

ii.pH Electrodes:

pH electrodes are widely used in laboratories and various industries for accurate pH measurement. These electrodes consist of a glass membrane that is sensitive to hydrogen ions and a reference electrode. When immersed in a solution, the glass membrane generates a voltage that is proportional to the pH of the solution. pH meters are used to measure this voltage and convert it into a pH value. It offers high accuracy and precision, making it suitable for scientific and industrial applications.

iii.pH Meter:

A pH meter is an electronic device used with pH electrodes for precise pH measurement. It measures the voltage generated by the electrode and displays the pH value on a digital screen. pH meters often include features like automatic temperature compensation (ATC) to correct for variations in temperature that can affect pH readings.

iv.Colorimetric Methods:

Colorimetric methods involve using chemical reagents that produce color changes in response to pH. A pH indicator dye is added to the sample solution, and the resulting color change is compared to a color chart or measured with a spectrophotometer. This method can be accurate when used with appropriate equipment but may require more time and resources than simpler methods.

v.Ion-Selective Electrodes (ISEs) :

ISEs are specialized electrodes that measure the activity of specific ions, including hydrogen ions (H+). They can provide accurate pH measurements and are particularly useful in situations where traditional pH electrodes may not work well, such as in highly concentrated or non-aqueous solutions.

vi.Litmus Paper:

Litmus paper is a type of pH indicator paper that uses natural dyes extracted from lichens. It changes color in response to pH, turning red in acidic solutions and blue in alkaline solutions. Litmus paper is primarily used for qualitative pH testing and is less precise than other methods.

vii.Potentiometric Methods: Potentiometric pH measurement involves using a high-impedance voltmeter to measure the voltage across a specialized pH-sensitive electrode. This method is similar to pH meters but may offer increased sensitivity and flexibility for research and specialized applications.

The choice of pH measurement method depends on the specific requirements of the application, the level of accuracy needed, and the available budget and equipment. pH meters with glass electrodes are the standard for high-precision measurements, while simpler methods like pH test strips are convenient for quick, qualitative assessments.

Problem 6.1:

Find out the pH and pOH of 0.001M solution of KOH?
Solution: Potassium hydroxide solution is a strong base. It ionizes completely such that one mole of KOH gives one mole of OH⁻ ions.

$$KOH_{(aq)} \longrightarrow K^+_{(aq)} + OH^-_{(aq)}$$

Therefore, 0.001M solution of KOH produces 0.001M OH⁻ ions.

$$[OH^-] = 0.001 \text{ M} \quad \text{or} \quad 10^{-3} \text{ M}$$
$$pOH = -\log 10^{-3} = 3$$
$$pH = 14 - 3 = 11$$

6.3 Buffer System:

6.3.1 Buffer:

A buffer, in the context of chemistry and biochemistry, is a solution that resists changes in its pH (acidity or alkalinity) when an acid or base is added to it. Buffers play a crucial role in maintaining the stability of pH levels in various chemical and biological systems. They are essential in many laboratory experiments, industrial processes, and biological functions.

6.3.2 Composition:

A buffer solution typically consists of a weak acid and its corresponding conjugate base (or a weak base and its corresponding conjugate acid). This combination of a weak acid and its conjugate base is referred to as a "buffer pair." The weak acid can donate hydrogen ions (H+) to the solution, and the conjugate base can accept hydrogen ions, which helps to neutralize changes in pH.

6.3.3 Characteristics of buffer solution:

(i) It has a definite pH.

(ii) Its pH does not change on standing for long periods of time.

(iii) Its pH does not change on dilution.

(iv) Its pH is slightly changed by the addition of small quantity of an acid or base.

6.3.4 Types of buffer solutions:

(a) Acidic Buffer:

It is formed by the mixture of weak acid and salt with a strong base.

Examples:

a) $CH_3COOH + CH_3COONa$,
b) $HCN + NaCN$
c) Boric acid + Borax etc.

(b) Basic Buffer:

It is formed by the mixture of a weak base and its salt with strong acid.

Examples:

a) $NH_4OH + NH_4Cl$,
b) $NH_4OH + NH_4NO3$
c) Glycine + Glycine hydrochloride

(c) Simple Buffer:

It is formed by a mixture of acid salt and normal salt of a polybasic acid,

Example

a) $Na_2HPO_4 + Na_3PO_4$
b) Or a salt of weak acid and a weak base. Example: CH_3COONH_4

6.3.5 Buffer capacity

Buffer capacity is defined as the number of moles of acid or base added in one litre of solution as to change the pH by unity.

Buffer capacity (Φ) = No. of moles of acid or base added to 1 litre solution/change in pH

$\Phi = \partial b / \partial(pH)$

Where ∂b – No. of moles of acid or base added to 1 litre

$\partial(pH)$ – change in pH

6.3.6 Applications of Buffer in chemistry

There are many applications of buffer solutions in chemistry. Few are given below:

i.   **Stable Reactions**: Buffers help keep chemical reactions from going too acidic or too basic, ensuring they happen smoothly.

ii.   **Precise Measurements:** In chemistry tests, buffers make sure we get accurate measurements of how acidic or basic something is.

iii.   **Enzyme Studies:** Buffers help scientists to study enzymes (special proteins) by keeping the right pH for them to work well.

iv.   **Separating Stuff:** In processes like filtering or sorting chemicals, buffers help keep the pH steady for precise separation.

v.   **pH Testing:** Buffers are used to check if pH meters and sensors are working correctly in labs and industries.

vi.   **Biotech and Medicine:** In labs making medicines or working with cells, buffers make sure things stay stable and safe.

vii.   **Metal Coating:** Buffers help control the pH in processes like adding a shiny metal layer to objects.

viii.   **Food and Drinks:** Buffers keep the right pH in foods and drinks so they taste good and stay safe.

ix.   **Clean Water**: In water treatment, buffers help make tap water safe by keeping the pH right and preventing pipe corrosion.

EXERCISE

Give short answers.

i.   Name three common household substances having
ii.   pH value greater than 7 b. pH value less than 7 c. pH value equal to 7
iii. Define pH. What is the pH of pure water?
iv.   How many times a solution of pH 1 will be stronger than that of a solution having pH 2?

Numericals

1. Calculate the pH and pOH of 0.2 M H2SO4?
2. Calculate the pH of 0.1 M KOH?
3. Calculate the pOH of 0.004 M HNO3?
4. Complete the following Table:

Solution	$[H^+]$	$[OH]$	$[pH]$	$[pOH]$
(i) 0.15 M HI				
(ii) 0.040 M KOH				
(iii) 0.020 M Ba(OH)$_2$				
(iv) 0.00030 M HClO$_4$				
(v) 0.55 M NaOH				
(iv) 0.055 M HCl				
(vii) 0.055 M Ca(OH)$_2$				

Chapter # 7

Electrolytes and Electrolysis

STUDENTS LEARNING OUTCOMES:

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen.
- Define oxidation and reduction in terms of loss or gain of electrons.
- Define oxidizing and reducing agents in a redox reaction.
- Define oxidation state.
- State the common rules used for assigning oxidation numbers to free elements, ions (simple and complex), molecules, atoms .
- Determine the oxidation number of an atom of any element in a compound. List the possible uses of an electrolytic cell.
- Sketch a Daniel cell, labelling the cathode, the anode, and the direction of flow of the electrons.
- Describe how a battery produces electrical energy.
- Identify the half-cell in which oxidation occurs and the half-cell in which reduction occurs given a voltaic cell.
- Distinguish between electrolytic and voltaic cells.
- Describe the methods of preparation of alkali metals.
- Describe the manufacture of sodium metal from fused NaCl.
- Identify the formation of by products in the manufacture of sodium metal from fused NaCl.

Electrochemistry is concerned with the conversion of electrical energy into chemical energy in electrolytic cells as well as the conversion of chemical energy into electrical energy in galvanic or voltaic cells.

In an electrolytic cell, a process called electrolysis takes place. In this process electricity is passed through a solution or the fused state of electrolyte to break it into its components. The electricity provides sufficient energy to cause non-spontaneous oxidation-reduction reaction to take place.

A galvanic cell, on the other hand, provides a source of electricity. This source of electricity results from a spontaneous oxidation-reduction reaction taking place in the solution. First of all, we should learn the theoretical background of oxidation and reduction reaction and try to understand the balancing of equations.

7.1 OXIDATION STATE AND BALANCING OF REDOX EQUATIONS

7.1.1 Oxidation Number or oxidation State:

It is the apparent charge on an atom of an element in a molecule or an ion. It may be positive or negative or zero.

7.1.2 Rules for Assigning Oxidation Number:

i. The oxidation number of all elements in the free state is zero. This is often shown as a zero written on the symbol. For example, H_2 ^{o}.

ii. The oxidation number of an ion, consisting of a single element, is the same as the charge on the ion. For example, the oxidation numbers of K^+ , Ca^{2+}, Al^{3+} , Br^- , S^{2-} are + 1, +2, +3, -1, -2, respectively.

iii. The oxidation number of hydrogen in all its compounds except metal hydrides is +1. In metal hydrides it is -1. ($Na^{+,}$ H^- , Mg^{2+} H_2 $^{-1/2}$)

iv. The oxidation number of oxygen in all its compounds except in peroxides, OF_2 and in super oxides is -2. It is -1 in peroxides +2 in OF_2 and -1/2 in super oxides.

v. In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero.

vi. In ions, the algebraic sum of oxidation number equals the charge on the ion.

vii. In any substance the more electronegative atom has the negative oxidation number.

The oxidation number or state of any atom of an element present in a compound or a radical can be determined by making use of the above said rules.

Example 7.1:

Calculate the oxidation number (O.N) of manganese in $KMnO_4$.

Solution.

(Oxidation number of K) + (oxidation number of Mn)+4 (oxidation number of O) = 0

Where	oxidation number of K	= +1
	oxidation number of O	= -2
Let	oxidation number of Mn	= x

Putting these values in the above equation.

$$(+1) + x + 4(-2) = 0$$

or $$x = +7$$

Thus the oxidation state of Mn in $KMnO_4$ is + 7.

7.1.3 Balancing of Redox Equations by Oxidation Number:

Method Carry out the following steps for balancing of redox equations by oxidation number method.

i. Write down the skeleton equation of the redox reaction under consideration.
ii. Identify the elements which change their oxidation number during the reaction.
iii. Record the oxidation number above the symbols of the element, which have changed the oxidation number.
iv. Indicate the change in oxidation number by arrows joining the atoms on both sides of the equation. It shows the number of electrons gained or lost.
v. Equate the increase or decrease in the oxidation number, i.e. electrons gained or lost by multiplying with a suitable digit.
vi. Balance the rest of the equation by inspection method.

Example 7.2 :

**Balance the following equation by oxidation number method.**

$$K_2Cr_2O_7 + HCl \rightarrow KCl + CrCl_3 + Cl_2 + H_2O$$
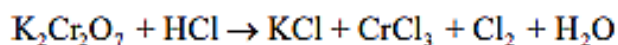
Let us balance the equation stepwise:

i. Write the equation with the oxidation number of each element

$$\overset{(+1)2\ (+6)2\ (-2)7}{K_2\ Cr_2\ O_7} + \overset{+1\ \ -1}{HCl} \rightarrow \overset{+1\ \ -1}{KCl} + \overset{+3\ \ (-1)3}{CrCl_3} + \overset{}{Cl_2^{\circ}} + \overset{(+1)2\ \ -2}{H_2O}$$
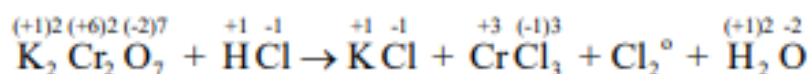
ii. Identify, those elements whose oxidation numbers have changed. Equation shows that Cr goes from + 6 to +3 and it is reduced. Cl goes from -1 to zero and is oxidized. Moreover, the oxidation number of chlorine remains the same, i.e from -1 to -1 when KCl and CrCl3 are produced. So, we should write HCl, twice on the left-hand side. One of HCl on left side shows those Cl atoms which do not change their oxidation numbers Other HCl shows those Cl atoms which changes their oxidation numbers.
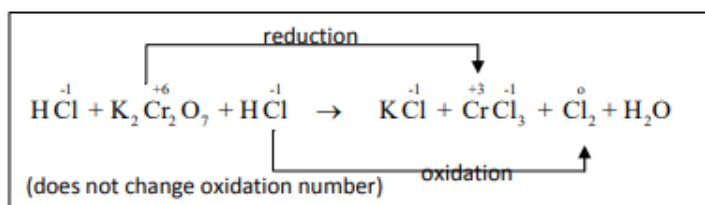

iii. Draw the arrows between the same elements whose oxidation numbers have changed. Also, point out the change in oxidation number. Cr has changed its oxidation number from + 6 to + 3 and chlorine has changed from -1 to zero. It means 6 electrons have been gained by two Cr atoms and 1 electron has been lost by 1 chlorine atom.
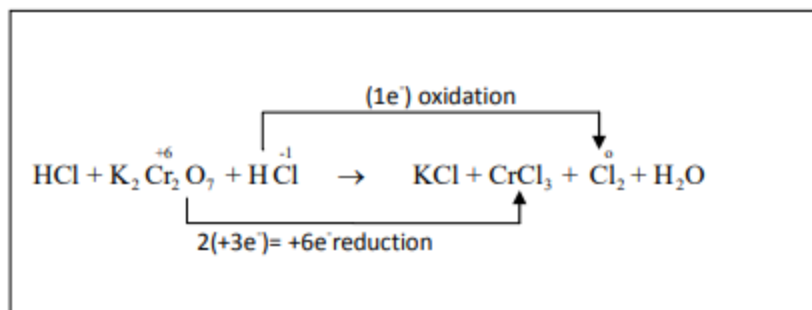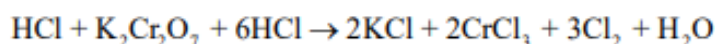
iv. To balance the number of electrons lost and gained multiply HCl with six. In this way, the 6 electrons lost by 6 Cl- will be gained by 2Cr+6 to give 2Cr+3. But do not multiply other HCl molecules with anything at this moment.

$$HCl + K_2Cr_2O_7 + 6HCl \rightarrow KCl + CrCl_3 + Cl_2 + H_2O$$
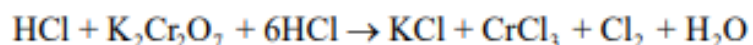
v. Balance Cr atoms by multiplying CrCl3 by 2. Balance Cl2 on right-hand side, whose oxidation number has changed by multiplying it with 3. In this way, the atoms which have been oxidized and reduced get balanced.

$$HCl + K_2Cr_2O_7 + 6HCl \rightarrow KCl + 2CrCl_3 + 3Cl_2 + H_2O$$
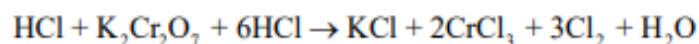
vi. To balance K atoms, multiply KCl by 2.

$$HCl + K_2Cr_2O_7 + 6HCl \rightarrow 2KCl + 2CrCl_3 + 3Cl_2 + H_2O$$

vii. Now balance those atoms of chlorine which have not been oxidized or reduced. There are 8 such chlorine atoms on the right hand side with KC1 and 2CrCl3. So multiply HCl with eight. This HCl has produced KCl and CrCl3.

$$8HCl + K_2Cr_2O_7 + 6HCl \rightarrow 2KCl + 2CrCl_3 + 3Cl_2 + H_2O$$
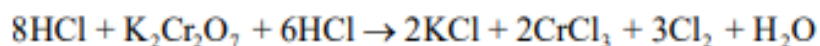
viii. Balance the rest of the equation by inspection method. To balance O atoms, multiply H2O with 7.

$$8HCl + K_2Cr_2O_7 + 6HCl \rightarrow 2KCl + 2CrCl_3 + 3Cl_2 + 7H_2O$$

This is the final balanced equation.

7.1.4 Balancing of Redox Equations by Ion-Electron Method

The balancing of redox equations by the loss and gain of electrons usually involves quite a few ions, which do not undergo change in valence and which are not necessary for the

process of balancing. The ion-electron method eliminates all the unnecessary ions and retains only those which are essential. Following are the general rules for balancing the redox equations by ion-electron method.

1. Write a skeleton equation that shows only those substances that are involved in the reaction.

2. Split the equation into two half reactions, one showing oxidation half reaction and the other reduction half reaction.

3. The element should not be written as a free atom or ion unless it exists as such. It should be written as a real molecular or ionic species.

4. Balance each partial equation as to the number of atoms of each element. In neutral or acidic solution, H_2O or H^+ ions may be added for balancing oxygen and hydrogen atoms. Oxygen atoms are balanced irst. If the solution is alkaline, OH- may be used for each excess oxygen on one side of the equation.

5. Balance each half reaction as to the number of charges by adding electrons to either the left or the right side of the equation.

6. Multiply each half reaction by a number chosen so that the total number of electrons lost by the reducing agent equals the number of electrons gained by the oxidizing agent

7. Add the two half reactions. Count the number of atoms of each element on each side of the equation and check the net charge on each side, which should be equal on both sides.

Balancing of redox equations by ion-electron method, making use of the above rules. There are two types of such reactions Le. in acidic medium and basic medium. Now, let us discuss one example of each.

Example7.3 : (Acidic medium)

Balance the equation for the reaction of HCl with $KMnO_4$ where Cl^- is oxidized to Cl_2 and MnO^{4-} is reduced to Mn^{2+}. The skeleton equation which does not contain either H^+ or H_2O, is

Solution

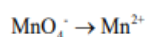$$Cl^- + MnO_4^- \rightarrow Cl_2 + Mn^{2+}$$
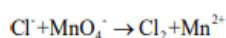
It is clear that Cl^- is oxidized to Cl_2 and MnO^{1-}_4 reduces to Mn^{2+}
Splitting the equation into half-reactions,
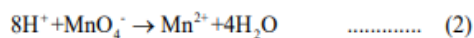Oxidation half reaction

$$Cl^- \rightarrow Cl_2$$

Reduction half reaction

$$MnO_4^- \rightarrow Mn^{2+}$$

Balancing atoms on both sides of oxidation half reaction.
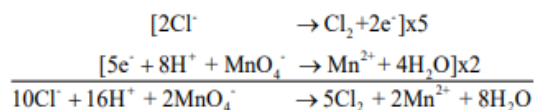
$$2Cl^- \rightarrow Cl_2 \qquad\qquad .... \quad (1)$$

Now, balance the reduction half reaction. To balance O-atoms, add $4H_2O$ on R.H.S. and to balance H-atoms add $8H^+$ on L.H.S. The reason is that medium is acidic.

$$8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O \qquad .... \quad (2)$$

Balancing the charges by adding electrons in equation (1) and (2), we get (3) and (4).

$$2Cl^- \rightarrow Cl_2 + 2e^- \qquad\qquad .... \quad (3)$$

$$8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O \qquad .... \quad (4)$$
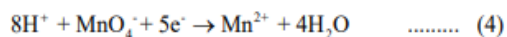
For making the number of electrons lost in first equation equal to the number of electron gained in the second equation, multiply the first equation by 5 and second by 2. After adding both equations and cancelling the common species on both sides, balanced equation is obtained.
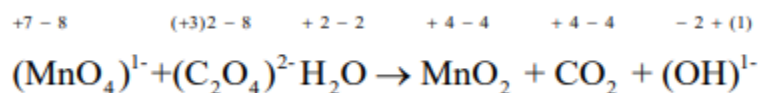
$$
\begin{array}{l}
[2Cl^- \qquad\qquad \rightarrow Cl_2 + 2e^-]x5 \\
\underline{[5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O]x2} \\
10Cl^- + 16H^+ + 2MnO_4^- \quad \rightarrow 5Cl_2 + 2Mn^{2+} + 8H_2O
\end{array}
$$

Example:7.4 (basic medium)

Balance the following equation in basic aqueous solution by ion-electron method.

$$MnO^{1-}_4(aq) + C_2O_4^{2-}(aq) + H_2O \rightarrow MnO_2(s) + CO_2(g) + OH^-(aq)$$
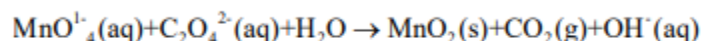
Solution:

The following steps are involved in balancing of equation in basic aqueous solution by ionelectron method.

i. Identify those elements which undergo change in oxidation number by writing number above each element.

$$
\overset{+7\ -8}{(MnO_4)^{1-}} + \overset{(+3)2\ -8}{(C_2O_4)^{2-}} \overset{+2\ -2}{H_2O} \rightarrow \overset{+4\ -4}{MnO_2} + \overset{+4\ -4}{CO_2} + \overset{-2\ +(1)}{(OH)^{1-}}
$$

ii.   Split the reaction into two half reactions, the oxidation and reduction half reactions.

$$C_2O_4^{2-} \rightarrow CO_2 \qquad \text{(oxidation half reaction)}$$

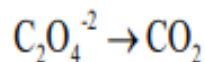$$MnO_4^{-1} \rightarrow MnO_2 \qquad \text{(reduction half reaction)}$$

The elements undergoing a change in oxidation number are Mn and C.

iii.   Split the reaction into two half reactions, the oxidation and reduction half reactions.

$$C_2O_4^{2-} \rightarrow CO_2 \qquad \text{(oxidation half reaction)}$$

$$MnO_4^{-1} \rightarrow MnO_2 \qquad \text{(reduction half reaction)}$$

iv.   Balancing of Oxidation Half Reaction:

$$C_2O_4^{-2} \rightarrow CO_2$$

v.   Balancing the C atoms in both sides of the half reaction.

$$C_2O_4^{-2} \rightarrow 2CO_2$$
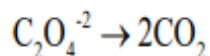
vi.   Balancing the charges on both sides of the half reaction by adding the appropriate number of electrons to the more position side.

$$C_2O_4^{-2} \rightarrow 2CO_2 + 2e^- \qquad \text{....... (1)}$$

The oxidation half reaction is balanced.

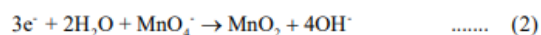<u>Balancing of Reduction Half Reaction:</u>

$$MnO_4^- \rightarrow MnO_2$$
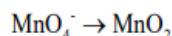
Balance in O-atoms by adding OH ions on the side needing the oxygen. Add two OH ions for each oxygen atom needed. So, we have to add 4OH on R.H.S:

$$MnO_4^- \rightarrow MnO_2 + 4OH^-$$

Balance the hydrogen, by adding H_2O on the other side of the half reaction. Add one H_2O for each two OH ion. In this way, oxygen and hydrogen atoms are balanced.
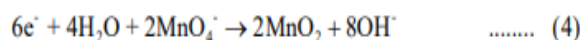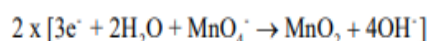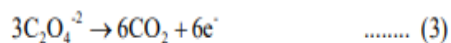
$$2H_2O + MnO_4^- \rightarrow MnO_2 + 4OH^-$$
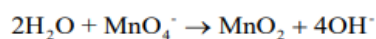
Balance the charges by adding three electrons to L.H.S. of equation

$$3e^- + 2H_2O + MnO_4^- \rightarrow MnO_2 + 4OH^- \qquad \text{........ (2)}$$
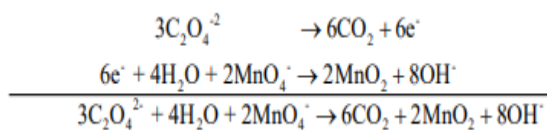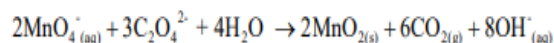
The reduction half reaction is balanced.

(iii)   Multiply each half reaction by an appropriate number, so that the number of electrons on both the half reactions becomes equal. For this purpose, multiply the oxidation half reaction by 3 and the reduction half reaction by 2.

$$3C_2O_4^{-2} \rightarrow 6CO_2 + 6e^- \qquad \text{........ (3)}$$

$$2 \times [3e^- + 2H_2O + MnO_4^- \rightarrow MnO_2 + 4OH^-]$$

$$6e^- + 4H_2O + 2MnO_4^- \rightarrow 2MnO_2 + 8OH^- \qquad \text{........ (4)}$$

(iv)   Add the two half-reactions to get the net ionic equation and cancel out anything appearing on both sides of the equation. For this purpose, add equation and equation (4).

$$3C_2O_4^{-2} \qquad \rightarrow 6CO_2 + 6e^-$$
$$\underline{6e^- + 4H_2O + 2MnO_4^- \rightarrow 2MnO_2 + 8OH^-}$$
$$3C_2O_4^{2-} + 4H_2O + 2MnO_4^- \rightarrow 6CO_2 + 2MnO_2 + 8OH^-$$

Hence, the balance ionic equation is

$$2MnO_{4~(aq)}^- + 3C_2O_4^{2-} + 4H_2O \rightarrow 2MnO_{2(s)} + 6CO_{2(g)} + 8OH_{(aq)}^-$$

7.2 Electrolytic Cells:

Look at the arrangement as shown in Fig. (1). It represents an electrolytic cell; The electrolyte consists of positive and negative ions which are free to move in the solution.

When a direct current (D.C) source is connected to the electrodes of the cell, each electrode acquires an electric charge. Thus, on applying electric potential, the positive ions migrate towards the negative electrode, called cathode and the negative ions move towards the positive electrode, called the anode.
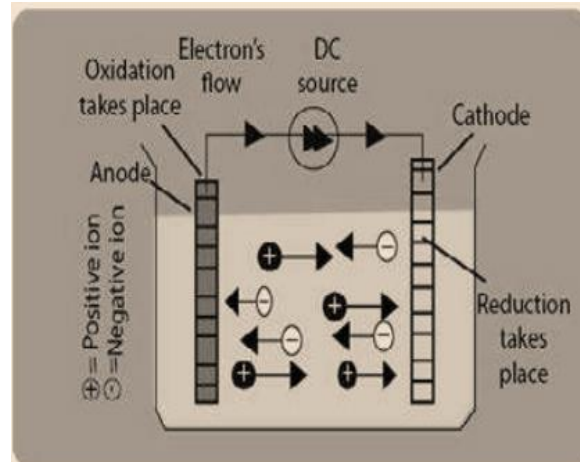

Fig:7.1.The migration of ions in electrolytic cell

This movement of ionic charges through the liquid brought by the application of electricity is called electrolytic conduction and the apparatus used is known as electrolytic cell. When electrolytic conduction occurs, electrochemical reactions take place. The ions in the liquid encounter the electrodes.

At the anode the negative ions give up electrons and are, therefore, oxidized.

At the cathode the positive ions pick up electrons and are reduced.

During electrolytic conduction, oxidation takes place at the anode and the reduction takes place at the cathode. The liquid will continue to conduct electricity only as long as oxidation-reduction reactions, occurring at the electrodes, continue. The electrochemical reactions that occur at the electrodes during the electrolytic conduction constitute the phenomenon of electrolysis.

When a molten salt is electrolyzed, the products are predictable. When an aqueous solution of a salt is electrolyzed, hydrogen and oxygen appear at the cathode and anode, respectively in certain cases. The products formed from a few electrolytes are shown in Table.

Electrolyte	Cathode	Anode
$PbBr_{2(molten)}$	Pb(s)	$Br_2(g)$
$NaCl_{(molten)}$	Na(s)	$Cl_2(g)$
NaCl(aq)	$H_2(g)$	$Cl_2(g)$
$CuCl_2(aq)$	Cu(s)	$Cl_2(g)$
$CuSO_4(aq)$	Cu(s)	$O_2(g)$
$KNO_3(aq)$	$H_2(g)$	$O_2(g)$
NaOH(aq)	$H_2(g)$	$O_2(g)$
$H_2SO_4(aq)$	$H_2(g)$	$O_2(g)$

Products of electrolysis (using inert electrodes of platinum or graphite)

Electrolyte	Copper cathode	Copper anode
$CuSO_4(aq)$	Cu deposits	Cu (s) dissolves to form CU^{2-} ions
Electrolyte	Silver cathode	Silver anode
$AgNO_3(aq)$ and $HNO_3(aq)$	Ag deposits	Ag (s) dissolves to form Ag^+ ions

Products of electrolysis (when electrodes take part in the reaction)

## 7.2.1:Explanation of Electrolysis

Fused Salts

When a fused salt is electrolyzed, the metal ions called cations arrive at the cathode which being negatively charged supply electrons to them and thus discharge the cations. The anions move towards the anode, give up their electrons and are thus discharged. In the case of fused lead chloride, the equations for electrode processes are given as below.

At anode: $Pb^{2+}_{(\ell)} + 2e^- \rightarrow Pb_{(s)}$ (oxidation)

At cathode: $2Cl^-_{(\ell)} \rightarrow Cl_{2(g)} + 2e^-$ (reduction)
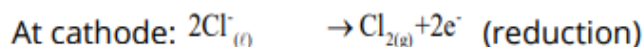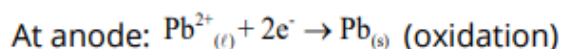
So, oxidation happens at anode and reduction at the cathode. Similarly, for fused NaCl and fused PbBr₂ the electrolytes are decomposed during electrolysis. Fused Pb and Na are deposited at cathode and Cl₂ (g) and Br₂ at anode. Electrons low through the external circuit from anode to cathode. The electric current is conducted through the cell by the ions and through the external circuit by the electrons.

Aqueous Solutions of Salts:

The electrolysis of aqueous solutions is somewhat more complex. Its reason is the ability of water to be oxidized as well as reduced. Hence, the products of electrolysis are not precisely predictable. Some metal cations are not discharged from their aqueous solutions. While, electrolyzing aqueous sodium nitrate ($NaNO_3$) solution, sodium ions present are not discharged at the cathode. A small concentration of hydronium and hydroxyl ions arises from the dissociation of water:

$$NaNO_3 \rightarrow Na^+ + NO_3^-$$

$$2H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$
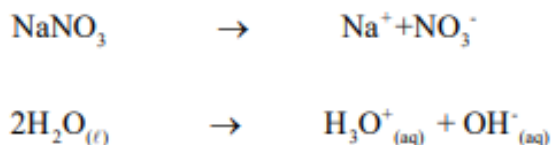
Hydronium ions accept electrons from the cathode to form hydrogen atoms:

$$\text{At cathode: } H_3O^+_{(aq)} + e^- \rightarrow H_{(g)} + H_2O_{(\ell)} \text{ (reduction)}$$
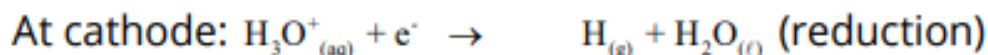
Subsequently, hydrogen atoms combine rapidly to form hydrogen molecules at the cathode.

$$H_{(g)} + H_{(g)} \rightarrow H_{2(g)}$$
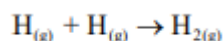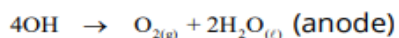
So, H_2 gas evolves at the cathode.

The concentration of hydronium ions is only 10-7 moles dm-3 in pure water. When these are discharged then more are formed by further dissociation of water molecules. This gives a continuous supply of such ions to be discharged. Sodium ions remain in solution, while hydrogen is evolved at the cathode. Thus, the reduction of the solute cations depends on the relative ease of the two competing reactions.

At the anode, both nitrate and hydroxide ions are present. Hydroxide ions are easier to discharge than nitrate ions. Nitrate ions remain in solution while the electrode reaction is:

$$\text{At anode: } OH^-_{(aq)} \rightarrow OH_{(aq)} + e^- \text{ (oxidation)}$$
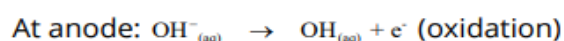
The OH groups combine to give O_2 gas as follows.

$$4OH \rightarrow O_{2(g)} + 2H_2O_{(\ell)} \text{ (anode)}$$

So, O_2 gas evolves at the anode.

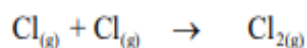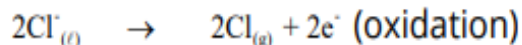But remember that the expected order of the discharge of ions may also depend upon their concentrations.

7.2.2:Electrolytic Processes of Industrial Importance

Various types of electrolytic cels are employed on industrial scale. Some of the important ones are given here.
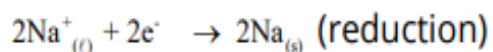
(i) **Extraction of sodium** by the electrolysis of fused sodium chloride is carried out in Down's cell. In this case, molten sodium chloride is electrolyzed between iron cathode and graphite anode. The cell is planted to get sodium metal commercially chlorine is obtained as a byproduct.
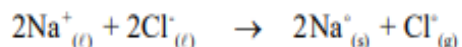
$$NaCl_{(s)} \rightarrow Na^+_{(\ell)} + Cl^-_{(\ell)}$$

At anode

$$2Cl^-_{(\ell)} \rightarrow 2Cl_{(g)} + 2e^- \text{ (oxidation)}$$

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

At cathode

$$2Na^+_{(\ell)} + 2e^- \rightarrow 2Na_{(s)} \text{ (reduction)}$$

By adding the two reactions at anode and cathode, the overall reaction is

$$2Na^+_{(\ell)} + 2Cl^-_{(\ell)} \rightarrow 2Na^\circ_{(s)} + Cl^\circ_{(g)}$$

(ii) **Caustic soda** is obtained on industrial scale by the electrolysis of concentrated aqueous solution of sodium chloride using titanium anode and mercury or steel cathode This electrolysis is carried out in Nelson cell and Castner- Kellner cell or Hg- cell.

$$NaCl_{(s)} \quad \square \quad Na_{(aq)}^+ + Cl^-_{(aq)}$$

At anode $\qquad 2Cl^-_{(g)} \quad \rightarrow \quad Cl_{2(g)} + 2e^- \qquad$ (oxidation)

At cathode $\qquad 2H_2O_{(\ell)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)} \qquad$ (reduction)
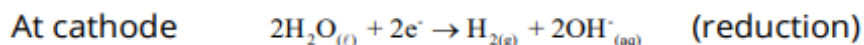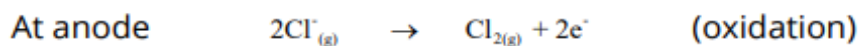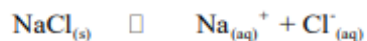
By combining, the electrode reactions and including Na⁺ ions, the overall reaction is

$$2Na^+_{(aq)} + 2Cl^-_{(aq)} + 2H_2O_{(l)} \rightarrow Cl_{2(g)} + H_{2(g)} + 2Na^+_{(aq)} + 2OH^-_{(aq)}$$
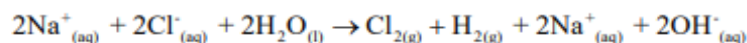
Here, chlorine and hydrogen are obtained as by products, and Na+ is not discharged at cathode.

(iii) Magnesium and calcium metals are extracted by the electrolysis of their fused chlorides. Mg and Ca are collected at cathodes while Cl_2 at anodes.

(iv) Aluminium is extracted by electrolyzing fused bauxite, $Al_2O_3.2H_2O$ in the presence of fused cryolite, Na_3AlF_6. This process is called Hall-Beroult process.

Anodized aluminium is prepared by making it an anode in an electrolytic cell containing sulphuric acid or chromic acid, which coats a thin layer of oxide on- it. The aluminium oxide layer resists attack from corrosive agents. Freshly anodized aluminium is hydrated and can absorb dyes.

(v) Electrolytic cells can also be used for the **purification of copper**. Impure copper makes the anode, and a thin sheet of pure copper is made the cathode. Copper sulphate solution is used as an electrolyte. The atoms of Cu from impure Cu- anode are converted to Cu2+ions and migrate to cathode which is made up of pure Cu. In this way Cu anode is puriied. Impurities are left at the anode.

Copper, silver, nickel, and chromium plating is done by various types of electrolytic cells. One metal is deposited at the surface of another metal.

7.3 Voltaic or Galvanic Cell:

A voltaic or a galvanic cell consists of two half-cells that are electrically connected. Each half cell is a portion of the total cell in which a half reaction takes place. Fig. (7.2) shows such a galvanic cell. The left half-cell consists of a strip of zinc metal dipped in 1.0 M solution of zinc sulphate giving the following equilibrium:

$$Zn_{(s)} \quad \rightarrow \quad Zn^{2+}_{(aq)} + 2e^-$$
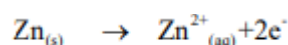
The right half-cell is a copper metal strip that dips into 1.0 M copper sulphate solution and the equilibrium here is represented as follows:

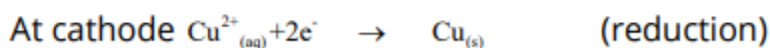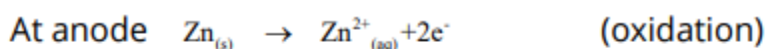$$Cu(s) \rightarrow Cu^{2+}+2e^-$$
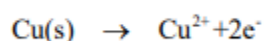
These half-cells in Fig (2) are connected electrically by a salt bridge. If the solutions were to mix, direct chemical reactions would take place, destroying the half-cells. The salt bridge contains an aqueous solution of potassium chloride in a gel. Zinc tends to lose electrons more readily than copper.

Fig (7.2) A Galvanic cell consisting of Zn and Cu electrodes at 25°C and unit concentration of electrolytic solutions.
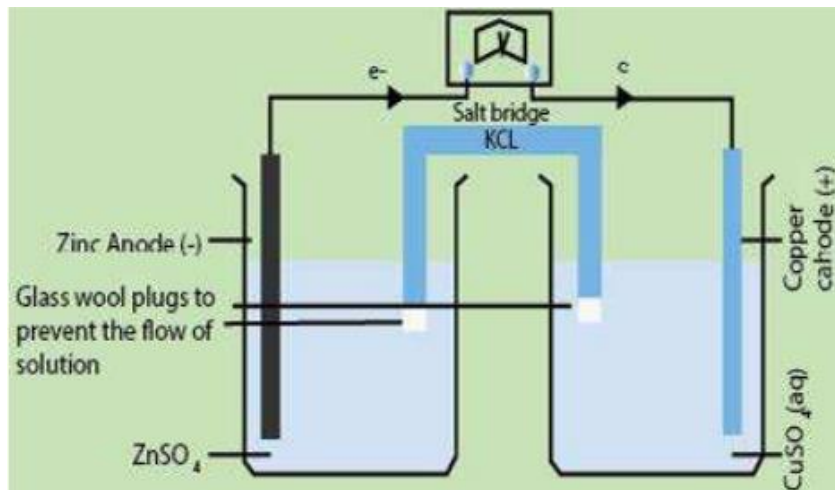
Zn electrode takes on a negative charge relative to the copper electrode. If the external circuit is closed by connecting the two electrodes as shown in the igure, electrons low from the zinc through the external circuit to copper electrode. The following half-cell reactions occur at two electrodes and cell potential at standard conditions is 1.1volts . It is denoted by E°.

At anode $\quad Zn_{(s)} \rightarrow Zn^{2+}_{(aq)}+2e^-$ $\qquad$ (oxidation)

At cathode $Cu^{2+}_{(aq)}+2e^- \rightarrow Cu_{(s)}$ $\qquad$ (reduction)

The overall voltaic cell reaction is the sum of these two half cell reactions.

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)} \qquad E°=1.1V$$
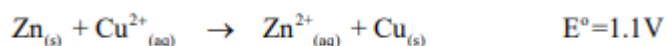
This voltaic cell can be represented as follows;

$$Zn_{(s)}/Zn^{2+}_{(aq)}1M \Box Cu^{2+}_{(aq)}+1M/Cu_{(s)} \qquad E°=1.1V$$
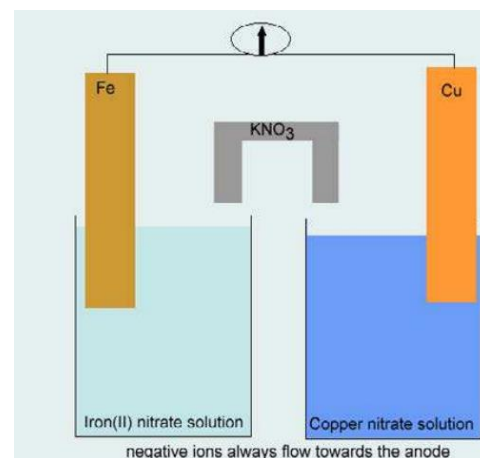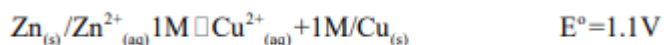
Note that reduction occurs at the copper electrode and oxidation occurs at the zinc electrode. Sign ⬜ shows the presence of salt bridge.

7.3.1:Function of Salt Bridge:

Let us, examine the purpose of the salt bridge. Since, zinc ions are produced as electrons leave the anode, we have a process which tends to produce a net positive charge in the left beaker. Actually, the concentration of Zn^{2+} ions increases in the left compartment. Similarly, the arrival of the electrons at the copper cathode and their reaction with copper ions tend to produce a net negative charge in the right beaker. The purpose of the salt bridge is to prevent any net charge accumulation in eith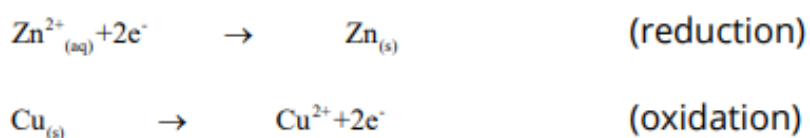er beaker by allowing negative ions to leave the right beaker, diffuse through the bridge and enter the left beaker. If this diffusional exchange of ions does not occur, the net charge accumulating in the beakers would immediately stop the flow of electrons through the external circuit and the oxidation-reduction reaction would stop. Many other oxidation-reduction reactions can be carried out successfully in galvanic cells using different electrodes. It is natural to think of these cell processes as separated into two half-reactions which occur at the two electrodes. In a voltaic cell the electric current in the external circuit can be used to light a bulb, drive a motor and so on.

7.3.2:Voltaic Cell is Reversible Cell:

On the other hand, if the external circuit is replaced by a source of electricity that opposes the voltaic cell, the electrode reactions can be reversed. Now, the external source pushes the electrons in the opposite direction and supplies energy or work to the cell so that the reverse non-spontaneous reaction occurs. Such a cell is called a reversible cell. For the zinc-copper cell, the half-cell reactions are reversed to give.

$$Zn^{2+}_{(aq)} + 2e^- \quad \rightarrow \quad Zn_{(s)} \qquad \text{(reduction)}$$

$$Cu_{(s)} \quad \rightarrow \quad Cu^{2+} + 2e^- \qquad \text{(oxidation)}$$

and the overall reaction being reversed, becomes

$$Zn^{2+}_{(aq)} + Cu_{(s)} \quad \rightarrow \quad Zn_{(s)} + Cu^{2+}_{(aq)}$$
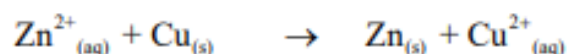
Oxidation occurs at the copper electrode and reduction takes place at the zinc electrode and the cell operates as an electrolytic cell in which energy from an external source drives a nonspontaneous reaction. When a cell operates as a voltaic the electrode at which reduction occurs is called the cathode while the electrode at which oxidation takes place is called the anode. Hence in voltaic cell, Zn acts as an anode and Cu acts as a cathode.

Exercise

I. Define oxidation in terms of electrons. Give an example.

II. Define reduction in terms of loss or gain of oxygen or hydrogen. Give an example.

III. What is difference between valency and oxidation state?

IV. Differentiate between oxidizing and reducing agents

V. Differentiate between strong and weak electrolytes.

VI. How electroplating of tin on steel is carried out?

VII. Why steel is plated with nickel before the electroplating of chromium.

VIII. How can you explain, that following reaction is oxidation in terms of increase of oxidation number?

IX. How can you prove with an example that conversion of an ion to an atom is an oxidation process?

X. Why does the anode carries negative charge in galvanic cell but positive charge in electrolytic cell? Justify with comments.

XI. Where do the electrons flow from Zn electrode in Daniel's cell?

XII. Why do electrodes get their names 'anode' and cathode in galvanic cell?

XIII. What happens at the cathode in a galvanic cell?

XIV. Which solution is used as an electrolyte in Nelson's cell?

XV. Name the by-products produced in Nelson's cell?

XVI. Discuss the construction and working of a cell in which electricity is produced

XVII. How can we prepare NaOH on commercial scale? Discuss its chemistry along with the diagram.

XVIII. Describe the rules for assigning the oxidation state.

XIX. Find out the oxidation numbers of the underlined elements in the following compounds.

(a) $Na_2\underline{S}O_4$ (b) $Ag\underline{N}O_3$ (c) $K\underline{Mn}O_4$

(d) $K_2\underline{Cr}_2O_7$ (e) $H\underline{N}O_2$

XX. How can a non-spontaneous reaction be carried out in an electrolytic cell? Discuss in detail.

Chapter # 8

Amines And Amides

STUDENTS LEARNING OUTCOMES:

- Basic concept and properties of Amines and Amides.
- Classification of Amines/Amides.

8.1 Introduction:

Amines and amides are important classes of organic compounds, each with distinct chemical structures and properties. Amines contain nitrogen atoms bonded to hydrogen and/or organic groups, and they play essential roles in biological molecules like amino acids, neurotransmitters, and alkaloids. Amines can be found in various natural and synthetic compounds, influencing the odor and flavor of many substances. On the other hand, amides feature a nitrogen atom bonded to a carbonyl group, and they are crucial components of proteins and peptides. Amides also serve as building blocks in the synthesis of pharmaceuticals, plastics, and other industrially significant materials.

Properties	Amines	Amides
Chemical Structure:	Amines have a nitrogen atom bonded to one or more carbon atoms, with hydrogen atoms and/or organic groups (alkyl or aryl) attached to the nitrogen atom.	Amides consist of a nitrogen atom bonded to a carbonyl group (C=O), with hydrogen atoms and/or organic groups attached to the nitrogen atom.
General formula:	The general formula for amines is $R-NH_2$, where R represents an organic group.	The general formula for amides is $RCONH2$, where R represents an organic group
Basicity	Amines are weak bases because they can accept a proton (H+) to form an ammonium ion (NH4+). The basicity of amines increases with the number of alkyl or aryl substituents attached to the nitrogen atom.	Amides are much less basic than amines. The presence of the carbonyl group reduces their basicity, making them poor proton acceptors.

Biological Significance	Amines are essential components of biological molecules such as amino acids, neurotransmitters, and alkaloids, playing critical roles in biological processes and signaling.	Amides are important in biological molecules like proteins and peptides, where they form peptide bonds between amino acids.

8.2 Amines:

Amines are one of the most important classes of organic compounds which can be derived when we replace one or more hydrogen atoms of an ammonia molecule with an alkyl group.

An amine is generally a functional group with a nitrogen atom having a lone pair. Amines resemble ammonia structurally where nitrogen can bond up to 3 hydrogen atoms. It is also characterized by various properties that are based on carbon connectivity.

Methylamine CH_3NH_2 Trimethylamine $(CH_3)_3N$

Ethylamine $C_2H_5NH_2$ Aniline

Diethylamine   Benzylamine
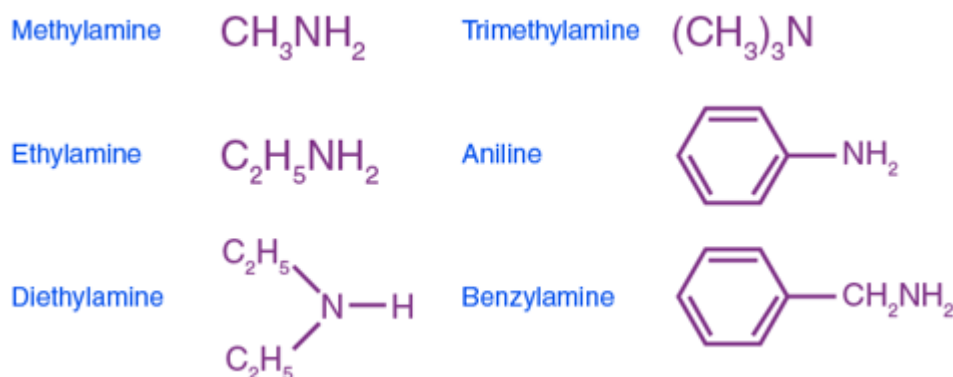
Amines are organic compounds that contain nitrogen atoms with a lone pair. Basically, they are derived from ammonia (NH_3) in which one or more hydrogen atoms is replaced by an alkyl or aryl group, and so they are known as alkylamines and arylamines respectively.

8.2.1 Amine Structure:
Nitrogen has 5 valence electrons and so is trivalent with a lone pair. According to VSEPR Theory nitrogen present in amines is sp³ hybridized and due to the presence of a lone pair, it is pyramidal instead of tetrahedral shape which is a general structure for most sp³ hybridized molecules. Each of the three sp3 hybridized orbitals of nitrogen overlap with orbitals of hydrogen or carbon depending upon the configuration of amines. Due to the presence of a lone pair, the C-N-H angle in amines is less
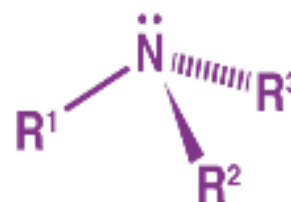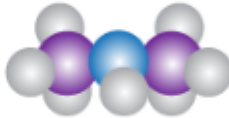
than 109 degrees which is a characteristic angle of tetrahedral geometry. The angle of the amines is near about 107 degrees.

8.2.2 Occurrence of Amines:

Naturally, amines occur in proteins, vitamins, hormones, etc. and they are also prepared synthetically to make polymers, drugs, and dyes.

8.2.3 Types of Amines:

On the basis of how the hydrogen atoms are replaced by an ammonia molecule, amines can be divided into 4 types.

| Ammonia | Primary amine | Secondary amine | Tertiary amine |

8.2.3.1:Primary Amines:

When one of the hydrogen atoms of the ammonia molecule is replaced by an alkyl or aryl group.Eg: Methylamine CH_3NH_2, Aniline $C_6H_5NH_2$

8.2.3.2:Secondary Amines:

Two organic substituents replace the hydrogen atoms of the ammonia molecule forming an amine.Eg: Dimethylamine $(CH_3)2NH$, Diphenylamine $(C_6H_5)_2NH$

8.2.3.3:Tertiary Amines:

When all 3 of the hydrogen atoms are replaced by an organic substituent, it could be an aryl or aromatic group.Eg: Trimethylamine $N(CH_3)_3$, Ethylenediaminetetraacetic acid (EDTA)

8.2.3.4:Cyclic Amines:

These are secondary or tertiary amines in an aromatic ring structure. Eg: Piperidine $(CH_2)_5NH$, Aziridines C_2H_5N

8.2.4 Preparation of Amines:

Some processes for preparing primary amines are mentioned below.

Preparation of Primary Amines

1.From halogenoalkanes:

This process will be carried out in a sealed tube. Here haloalkanes will be heated with the concentrated solution of ammonia in ethanol. The mixture cannot be heated under the reflux as ammonia would move out in the form of gas from a container.

$$CH_3CH_2Br \;+\; NH_3 \longrightarrow CH_3CH_2NH_3^+ \, Br^-$$
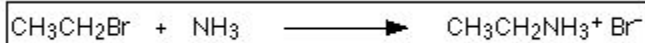
The preparation of primary amine from halogenoalkane, the reaction takes place in two stages. Salt will be formed at the first stage. Here ethyl ammonium bromide is the salt. It is like ammonium bromide except for the fact that one of the hydrogens in the ammonium atom is replaced by an ethyl group.

$$CH_3CH_2NH_3^+ \, Br^- \;+\; NH_3 \;\rightleftharpoons\; CH_3CH_2NH_2 \;+\; NH_4^+ \, Br^-$$
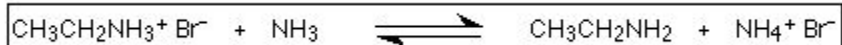
A reverse reaction can occur between ammonia and salt. It is illustrated in the above reaction.

2. Reduction of nitriles:

We can get primary amines when nitriles are reduced with lithium aluminium hydride. This method is mainly used for the preparation of amines which contain one carbon atom more than the starting amine.

$$R - C \equiv N \;\xrightarrow[\text{2) } H_2O]{\text{1) } LiAlH_4}\; R - CH_2 - NH_2$$

Nitrile   1° Amine
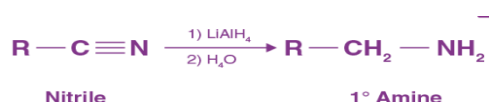
3. Gabriel phthalimide synthesis:

Primary amines can easily prepare by Gabriel synthesis. In this process, on the treatment of phthalimide with ethanolic potassium hydroxide, we get potassium salts of phthalimide. When this is further heated with alkyl halide followed by alkaline hydrolysis then primary amine is produced. We cannot prepare aromatic primary amines because aryl halides do not undergo nucleophilic substitution with the anion which is formed by phthalimide.
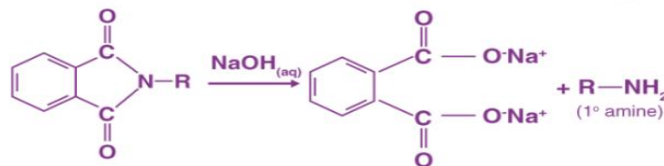
8.2.5 Basicity of Amines:

Like ammonia, primary & secondary amines have protic hydrogens and thus they showcase a degree of acidity. Whereas tertiary amines have no protic hydrogen and thus do not possess a degree of acidity.

pK$_a$ value for primary & secondary amines is about 38, which makes them a real weak acid. Whereas if we take the pK$_b$, it is about 4. This makes the amines much more basic than acidic. Thus, an aqueous solution of an amine is strongly alkaline.

8.2.6 Uses of Amines:

Amines have a widespread application in our daily lives.Some uses of amines are listed below:

i. It is used in water purification, medicine manufacturing and development of insecticides and pesticides.
ii. It is involved in the production of amino acids which is the building block of proteins in living beings. Many varieties of vitamins are also made by amines.
iii. Serotonin is an important amine that functions as one of the primary neurotransmitters. It controls the feelings of hunger and is critical for the speed with which the brain operates in general.
iv. Pain-relieving medicines such as Morphine and Demerol which are also known as analgesics are made from amines.

8.3 Amides

Amides are compounds that have a carbonyl functional group that is linked to both an amine and a hydrocarbon group (or hydrogen atom). A carbonyl functional group is made up of a double-bonded carbon atom and an oxygen atom.

An amide's general structure is R(C = O)NR'R", where R, R' and R" represent a hydrocarbon or a hydrogen substituent.
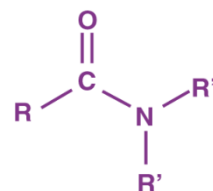
8.3.1 Synthesis of Amides

i.Amine Acylation

An amide is formed by the reaction of an amine with an acyl group (a carboxylic acid derivative).

In amines, the nitrogen atom has a lone pair (non-bonded) of electrons that attacks the electron-deficient carbonyl carbon and forms a bond with it. The elimination of a proton (from the amine substituent) and the leaving group yields the corresponding amide.
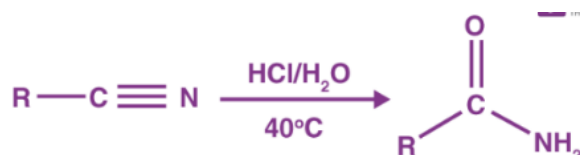
ii.Nitrile Hydrolysis

The formation of amides is caused by the partial hydrolysis of nitriles (RCN) under acidic or basic conditions.

iii.Beckmann Rearrangement

Beckmann rearrangement is one of the most important amide synthesis methods.

The reaction entails the formation of an amide from an oxime group. The rearrangement of oximes in the presence of an activating agent, such as H_2SO_4 or PCl_3, results in the formation of an amide.

8.3.2 Properties of Amides

An amide is formed when the nitrogen atom of an amine is attached to a carboxyl group.

i. The amide bond is polar, and therefore, it can form hydrogen bonds with water. As a result, amides are water-soluble (but only the lower members). The solubility of an amide decreases as the molar mass of its hydrocarbon substituent increases.
ii. The amide's high boiling and melting points are due to its solubility.
iii. Amides are less basic than amines because the lone pair of electrons on the nitrogen atom of the amide is conjugated with the carbonyl group, making the electrons less available for donation.

8.3.3 Examples of Amides

Acetamide $H_3C-CONH_2$, benzamide $C_6H_5CONH_2$, and dimethylformamide $HCON(CH_3)_2$ are examples of amides.

Acetamide   Benzamide   Dimethylformamide
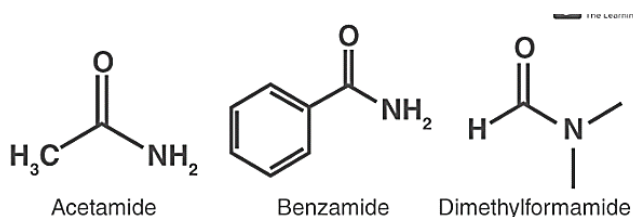
8.3.4 Classification of Amides:

Amides are divided into three types based on the substituents attached to the amide's nitrogen atom.

Primary, secondary, and tertiary amides are the three types of amides.

A **primary (1°) amide** is one in which the substituent on the nitrogen atom of the amide is only hydrogen atoms.

A **secondary (2°) amide** is one in which the amide's nitrogen atom is linked to a hydrocarbon substituent.

The nitrogen in a **tertiary (3°) amide** is attached to three carbons.

Lactam is the name given to a cyclic amide.

Primary amide	Secondary amide	Tertiary amide	Lactam
Nitrogen attached to one carbon	Nitrogen attached to two carbons	Nitrogen attached to three carbons	Generic name for a cyclic amide

8.3.5Applications:

Amides are widely used in technology as structural materials.

An amide connection is simple to form, resistant to hydrolysis and provides structural rigidity.

The most durable materials are polyamides and nylons.

A variety of drugs are amides, including penicillin, LSD, and paracetamol.

Furthermore, plant N-alkylamides have a diverse set of biological functions.

Exercise

I. Classify amides on the basis of substitutes.
II. What is the general formula for amines.
III. Write properties of amines.
IV. Write a few applications of amides.
V. Write a note on Beckmann Rearrangement.
VI. Differentiate between amines and amides.

Chapter # 9

PROTEINS

STUDENTS LEARNING OUTCOMES:

- Importance of proteins.
- Basic structure and sources of proteins.
- Identify different types of proteins.
- Uses of Proteins

Proteins are extremely complicated molecules of living things. They are nitrogeneous compounds made up of a variable number of amino acids. The human body probably contains at least 10,000 different kinds of proteins. The name protein is derived from the Greek word proteios meaning of prime importance.

Proteins are present in all living organisms and without proteins, life would not be possible. They are present in muscles, skin, hair, and other tissues that make up the bulk of the body's non-bony structure. All proteins contain the elements carbon, hydrogen, oxygen, and nitrogen. They may also contain phosphorus and traces of other elements like iron, copper, iodine, manganese, sulphur and zinc. Proteins are very high molecular weight macromolecules. All proteins yield amino acids upon complete hydrolysis. Thus, proteins may be defined as high molecular weight organic materials, which upon complete hydrolysis, yield amino acids.

Amino acids Amino acids are organic compounds consisting of both amino and carboxyl groups. They have the general formula:

$$\text{(side chain)} \; R-\underset{\underset{\displaystyle NH_2 \; \text{(amino group)}}{|}}{CH}-COOH \; \text{(carboxylic group)}$$
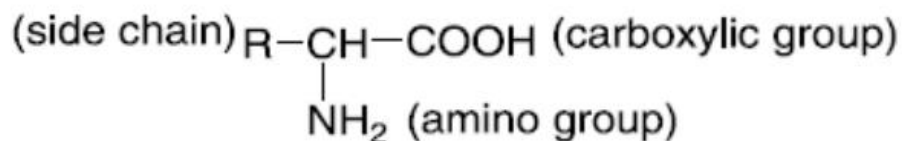
Side chain 'R' is different for different amino acids.

There are 20 amino acids. Ten out of twenty amino acids can be synthesized by the human body. These amino acids are called non-essential amino acids. While the other ten which cannot be synthesized by our bodies are called essential amino acids. Essential amino acids are required by our bodies and must be supplied through diet.

9.1 Amino acids are Building Blocks of Proteins:

Two amino acids linked through peptide linkage. Peptide linkage (bond) is formed by the elimination of water molecule between the amino group of one amino acid and carboxyl acid group of another, such as:

$$H_2N-CH-COOH \ + \ HNH-CH-COOH \xrightarrow{-H_2O} H_2N\left[CH-\overset{\overset{O}{\underset{||}{}}}{C}-NH\right]CH-COOH$$

(Dehydration)   Peptide linkage
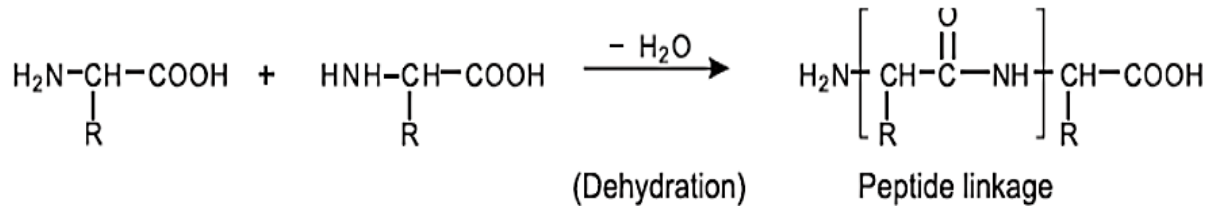
When thousands of amino acids polymerize, they form protein.

9.2 Classification of Proteins

Based on the physicochemical properties, proteins may be classified into three types.

  i.   Simple proteins
  ii. Compound or Conjugated proteins
  iii. Derived proteins

9.2.1 Simple Proteins

These proteins on hydrolysis yield only amino acids or their derivatives. For example, albumins, globulins, legumin, collagen, etc. Globulins are insoluble in water but soluble in dilute salt solutions. They are found in animals, e.g; lactoglobulin is found in muscles and in plants. Legumin and collagen proteins are present in the connective tissues throughout the body. They are the most abundant proteins in the animal kingdom forming some 25 to 35% of body protein.

9.2.2 Compound or Conjugated Proteins

In these molecules the protein is attached or conjugated to some non-protein groups which are called prosthetic groups. For example, phospho-proteins are conjugated with phosphoric acid, lipoproteins are conjugated with lipid substances like lecithin, cholesterol, and fatty acids.

9.2.3 Derived Proteins

This class of protein includes substances that are derived from simple and conjugated proteins. For example, protease enzymes, peptones, oligopeptides, polypeptides, etc. Based on their functions, proteins may also be classified as regulatory or hormonal proteins, structural proteins, transport proteins, genetic proteins, etc.
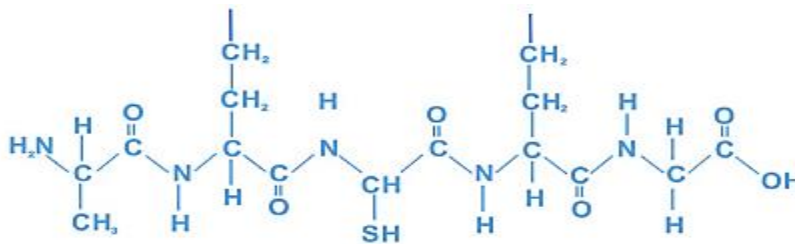
9.3 Structure of Proteins

Many proteins are compact, highly convoluted molecules with the position of each atom relative to the others determined with great precision. To describe the structure of a protein in an organism it is necessary to specify the three- dimensional shape that the polypeptide chain.

i.   Primary structure
ii. Secondary structure
iii. Tertiary structures

Some proteins also possess a fourth structure called the quaternary structure.

The sequence of the amino acids combined in a peptide chain is referred to as the primary structure.

The secondary structure of a protein is a regular coiling or zigzagging of polypeptide chains caused by hydrogen bonding between NH and C = 0 groups of amino acids near each other in the chains. The three-dimensional twisting and folding of the polypeptide chain results in the tertiary structure of proteins.

9.4 Denaturation of Proteins:

The structure of proteins can be disrupted easily by heat, change in pH, and under strongly oxidizing or reducing conditions. Under such conditions the proteins undergo denaturation. The most familiar example of denaturation is the change that takes place in albumin, the principal component of egg white, when it is cooked. In this case, the change is irreversible. The sequence of the amino acids combined in a peptide chain is referred to as the primary structure.

9.5 Sources and Uses of Proteins:

9.5.1 Animal Sources:
i.   Meat: Beef, chicken, pork, and lamb are rich sources of animal protein. Fish and Seafood: Salmon, tuna, shrimp, and other seafood provide high-quality protein.
ii. Dairy Products: Milk, cheese, yogurt, and eggs are excellent sources of protein.
iii. Poultry: Chicken and turkey are lean sources of animal protein.
iv. Eggs: Eggs are a complete protein source with all essential amino acids.

9.5.2 Plant Sources:

i.   Legumes: Beans (e.g., black beans, kidney beans, lentils), peas, and chickpeas are rich in plant-based protein.
ii. Nuts and Seeds: Almonds, peanuts, chia seeds, and flaxseeds contain substantial protein.
iii. Grains: Quinoa, oats, and whole wheat are grains with decent protein content.
iv. Vegetables: Some vegetables like broccoli, spinach, and peas provide protein in smaller quantities.

9.5.3 Processed Foods:

i.   Processed foods like protein bars, protein shakes, and meat substitutes (e.g., tofu-based products, seitan) are engineered to be protein-rich.

9.6 Uses of proteins:

Nutrition:

Proteins are essential macronutrients that are crucial for the growth, repair, and maintenance of tissues in the human body. They are a fundamental component of a balanced diet.

Muscle Development:

Protein is essential for building and repairing muscle tissue. Athletes and individuals involved in strength training often consume extra protein to support muscle growth.

Enzymes:

Many enzymes, which are biological catalysts, are proteins. They facilitate chemical reactions in the body, enabling processes like digestion, metabolism, and DNA replication.

Hormones:

Some hormones, such as insulin and growth hormone, are proteins. They regulate various physiological functions, including blood sugar levels and growth.

Immune Function:

Antibodies, which are proteins, play a critical role in the immune system by recognizing and neutralizing foreign invaders like bacteria and viruses.

Transport:

Hemoglobin, a protein in red blood cells, transports oxygen from the lungs to body tissues, while other proteins transport molecules across cell membranes.

Structure:

Collagen is a protein that provides structural support to connective tissues like skin, tendons, and bones.

Energy Source:

In the absence of carbohydrates and fats, proteins can be used as an energy source, although this is not their primary function.

Biotechnology and Industry:

Proteins are used in various industrial applications, such as the production of enzymes for detergents, biodegradable plastics, and pharmaceuticals.

Food Production:

Proteins are used in food processing to enhance texture, taste, and nutritional value. They are also used as food additives and preservatives.

Cosmetics and Personal Care:

Proteins are used in cosmetics and skincare products to improve skin and hair health.

9.7 Importance of Proteins:

i. Proteins take an essential part in the formation of protoplasm which is the essence of all forms of life.
ii. Nucleoproteins which are complexes of proteins with nucleic acids serve as carriers of heredity from one generation to the other.
iii. Enzymes which are biological catalysts are protein in nature. Without them life is not possible.
iv. Many proteins have specialized functions. Haemoglobin acts as a carrier of O_2. Some proteins act as hormones which have regulatory functions, for example; insulin, thyroxine etc.
v. Industrially proteins have great importance. We are familiar with the use of leather made by tanning of hides. This is essentially a precipitation of the proteins with tannic acid. Gelatin is obtained by heating bones, skin and tendons in water. It is used in bakery goods. Caesein is another protein used in the manufacture of buttons and buckles.

EXERCISE

Give short answers.

    i.   Which elements are found in proteins?
    ii.   How are amino acids bonded with each other?
    iii.   Give the general formula of amino acids.
    iv.   What do you mean by non-essential amino acids?
    v.   Why are the ten amino acids essential for us?
    vi.   How are proteins formed?
    vii.   How is gelatin obtained?
    viii.   Where are the proteins found?

Extensive Questions.

1. Explain the sources and uses of proteins.
2. Explain that amino acids are building blocks of proteins.

Chapter # 10

Carbohydrate

STUDENTS LEARNING OUTCOMES:

- Distinguish between mono-, di-and trisaccharides
- Explain the sources and uses of carbohydrates.

Most biologically important substances are organic compounds built up from the skeleton of carbon atoms. Many of them are very large molecules and most of these are polymers. The four major classes of organic compounds in living cells are carbohydrates, lipids, proteins, and nucleic acid.

The term carbohydrate is applied to a large number of relatively heterogeneous compounds. They are the most abundant biomolecules on earth.

Carbohydrates are macromolecules defined as polyhydroxy aldehydes or ketones. They have general formula $C_x(H_2O)_y$. They are commonly called 'sugars'. Carbohydrates are synthesized by plants through photosynthesis process from carbon dioxide and water in the presence of sunlight and green pigment chlorophyll.

$$6CO_2 \ + \ 6H_2O \xrightarrow[\text{chlorophyll}]{\text{sunlight}} C_6H_{12}O_6 \ + 6O_2$$
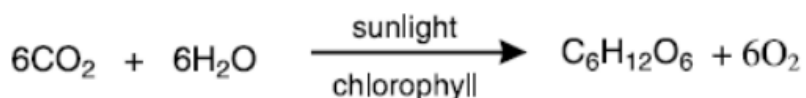
The glucose is further polymerized to form starch and cellulose.

10.1 Classification of Carbohydrates

i.Monosaccharides:

These are simple sugars that cannot be hydrolyzed. They have an empirical formula $(CH_2O)_n$ where n = 3 or a large number. Monosaccharides are either aldoses (aldehydic group) or ketoses (ketonic group). Common examples are glyceraldehyde, glucose, fructose, etc. Sugars with five carbon atoms (pentoses) or six carbon atoms (hexoses) are more stable as cyclic structures than as open-chain structures. Glucose and fructose are very common examples of hexoses, both of which have molecular formula, $C_6H_{12}O_6$.

Glucose

Glucose also called dextrose, grape sugar, or blood sugar, occurs naturally in both combined and free states. In the free state, it is present in most sweet fruits and honey. Small quantities of glucose are also present in human blood and urine. In the combined state it forms a major component of many disaccharides and polysaccharides. It is the source of energy in our body. Fructose is also found in combined and free states. It is used as a sweetening agent in confectionery and as a substitute for cane sugar. Other examples of monosaccharides are galactose and mannose.

Glucose (open chain form)

Glucose (cyclic form)

α-D-Glucose

ß-D-Glucose

Fructose (open chain form)

Fructose (cyclic form)
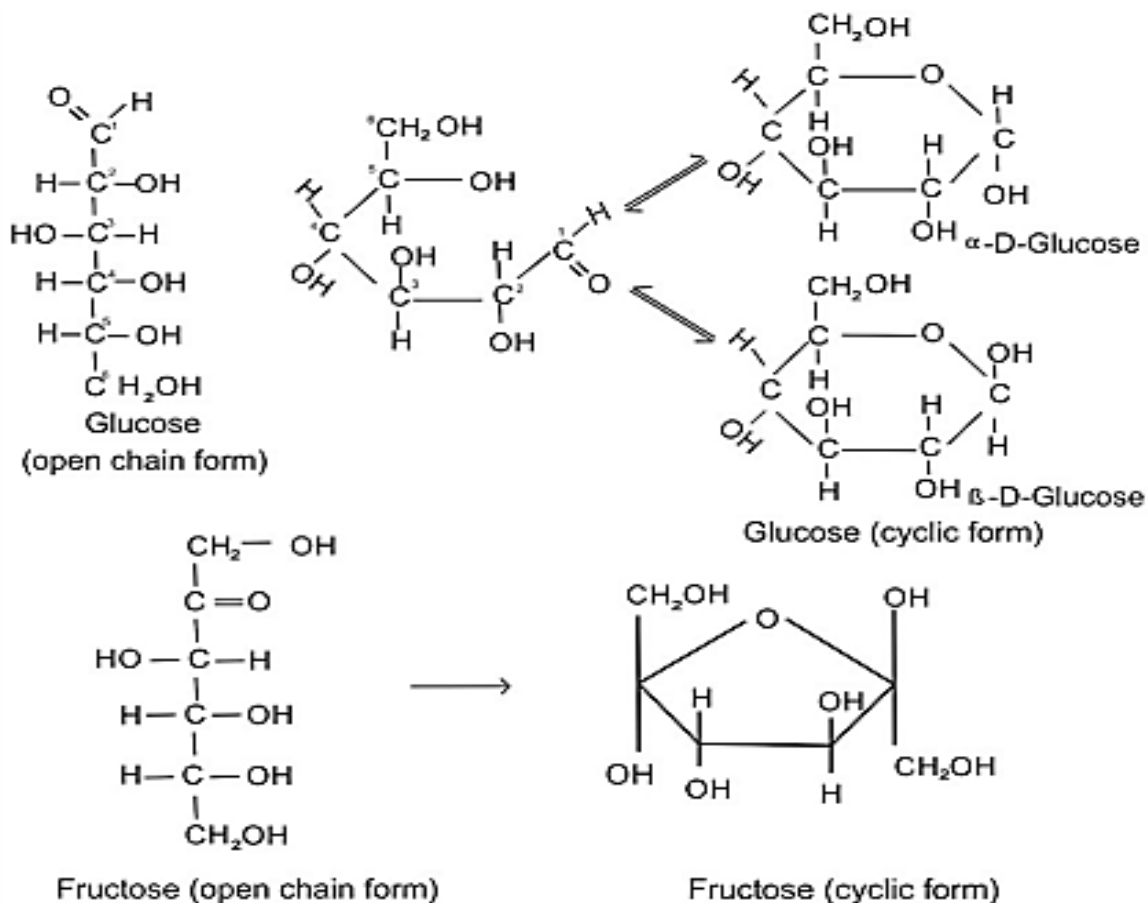
ii.Disaccharides or Oligosaccharides

The oligosaccharides are formed when two to nine monosaccharide units combine by the loss of water molecules. This results in the formation of a glycosidic linkage. For example; sucrose which is a common table sugar, is a disaccharide of glucose and fructose.

α-D-Glucose   Fructose

Glycosidic linkage
Sucrose
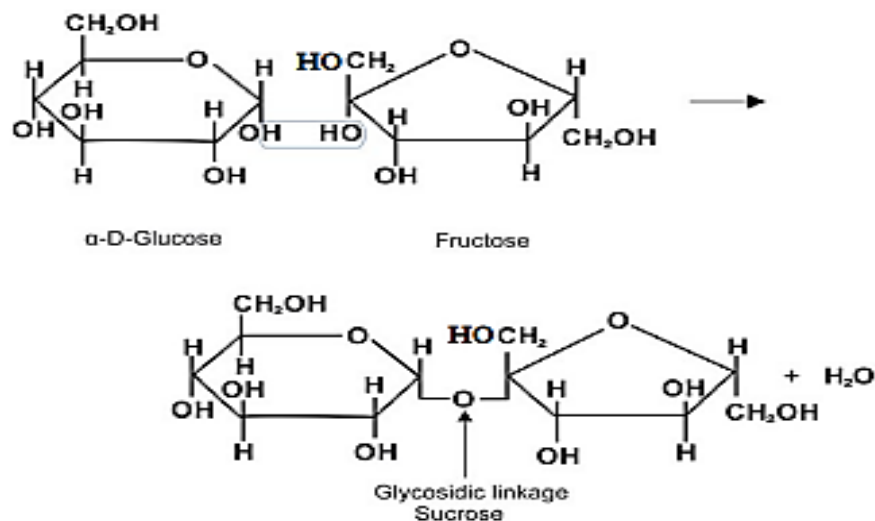
Conversely, hydrolysis of an oligosaccharide by water in the presence of an acid or by enzymes yields two or more monosaccharide units. Among the most common disaccharides are sucrose, lactose and maltose. Of these, sucrose occurs in sugar cane, sugar beet, pineapple, apricot, mango, almond, coffee and honey. Lactose (milk sugar) occurs in the milk of all animals. It does not occur in plants.

Trisaccharides, which yield three monosaccharide molecules on hydrolysis, have molecular formula, $C_{18}H_{32}O_{16}$, for example, raffinose. In general, monosaccharides and oligosaccharides are crystalline solids soluble in water and sweet to taste. They are collectively known as 'sugars'.

iii.Polysaccharides

The polysaccharides are carbohydrates of high molecular mass which yield many monosaccharide molecules on hydrolysis. Examples are starch and cellulose, both of which have molecular formula, $(C_6H_{10}O_5)_n$ .The polysaccharides are amorphous solids, insoluble in water and tasteless and are called 'non-sugars'. Polysaccharides perform two principal functions in animals and plants. They are used as energy storage compounds and for building structural elements of cells. Plants store glucose as starch and animals store glucose in the form of a highly branched polymer known as glycogen. Glycogen is stored in the liver and muscles.

a.Starch

Starch is the most important source of carbohydrates in the human diet. The chief commercial sources of starch are wheat, rice, maize, potatoes, and barley. Starch is a polymer of a-D-glucose. Starch is not a pure compound. It is a mixture of two polysaccharides, amylose and amylopectin which can be separated from one another.

Amylose is soluble in water and gives a deep blue colour with iodine while amylopectin is insoluble and gives no colour. Natural starch consists of 10 to 20% amylose and 80 to 90% amylopectin. It is used in coating and sizing paper to improve the writing qualities. It is also used in laundering and in the manufacture of glucose and ethyl alcohol.

Structure of Amylose (Polymer of α-D-glucose)

Structure of Amylopectin
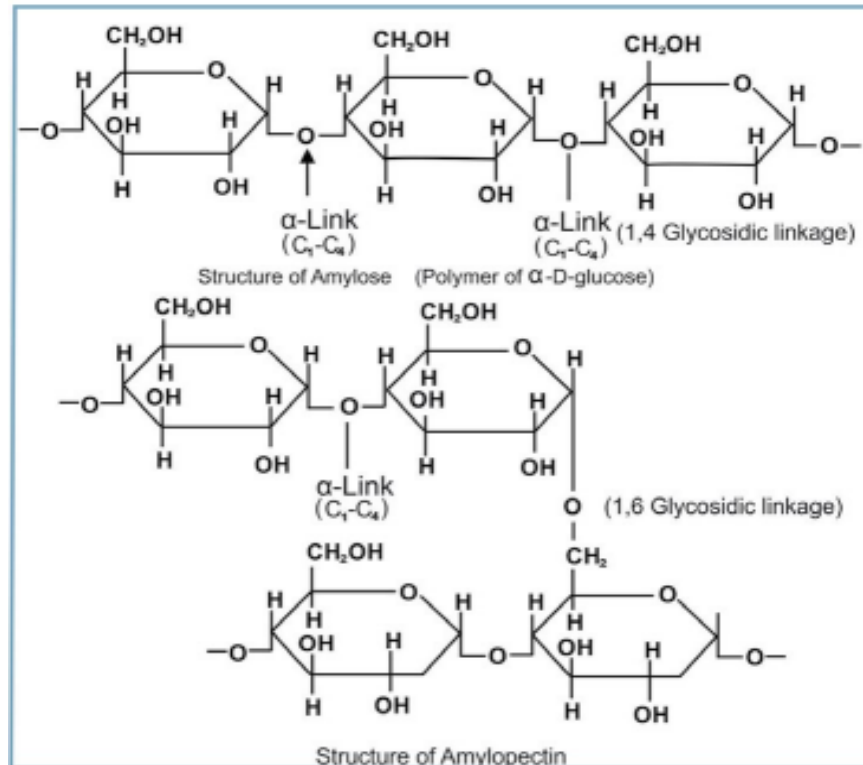
b. Cellulose

By far, the most abundant structural polysaccharide is cellulose. Some 100 billion tons of cellulose are produced each year by plants. For example, cotton is 99% cellulose, and the woody parts of trees are generally more than 50% cellulose. It is a polymer of β -D-glucose. It is present mainly in the plant kingdom but also occurs in some marine animals. It is an unbranched polymer consisting of a large number (up to 2500) of glucose residues joined to each other through β -1—>4 linkages.

Cellulose (Polymer of β-D-glucose)
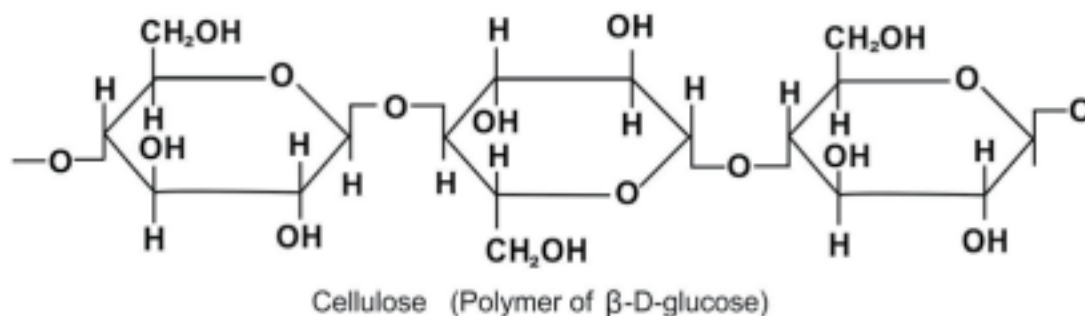
c.Glycogen

It occurs mainly in the liver and muscles where it represents the main storage polysaccharide in the same way as starch functions in plant cells. Glycogen is therefore also called 'animal starch'. Its structure closely resembles that of amylopectin having 1 4 and 1 6 glycosidic linkages. Human glycogen is a much more branched molecule than amylopectin. On hydrolysis it yields glucose units.

10.2 Sources and Uses of Carbohydrates

Carbohydrates are one of the three major macronutrients, along with proteins and fats, and they serve as a primary source of energy for the human body. Carbohydrates can be found in a variety of foods, both from plant and animal sources. Here are some common sources of carbohydrates:

i.Grains:

a) Cereals: Rice, wheat, oats, barley, and corn are staple grains that are rich in carbohydrates.

b) Bread: Various types of bread, including whole wheat, white, and multigrain, contain carbohydrates.

ii.Starchy Vegetables:

Potatoes: Potatoes are a significant source of carbohydrates, particularly when consumed as mashed potatoes, fries, or potato-based dishes.

Sweet Potatoes: Sweet potatoes are another starchy vegetable rich in carbohydrates.

iii.Legumes:

a)Beans: Varieties like black beans, kidney beans, pinto beans, and lentils are high in carbohydrates.

b)Peas: Green peas and split peas are sources of carbohydrates.

iv.Fruits:

a)Bananas:Bananas are known for their carbohydrate content, making them an excellent energy source.

b)Apples: Apples and other fruits like oranges, grapes, and berries provide natural sugars and carbohydrates.

c)Mangoes, Pineapples, and Grapes: These fruits are relatively high in carbohydrates compared to some other fruits.

iv.Dairy Products:

a)Milk: Milk contains lactose, a natural sugar and carbohydrate.

b)Yogurt: Yogurt contains lactose and is a source of carbohydrates.

c)Ice Cream: Ice cream and other frozen dairy desserts have added sugars and carbohydrates.

v.Sugars and Sweets:

a)Table Sugar (Sucrose): Table sugar is a common source of added sugars and simple carbohydrates.

Besides energy-providing materials, carbohydrates also provide the following usage to our body.

1. They regulate the amount of sugar level in our body. Low sugar level in the body results in hypoglycemia.

2. They provide essential nutrients for bacteria in the intestinal tract that help in digestion.

3. Dietary fiber helps to keep the bowel functioning properly.

4. Faber helps in lowering cholesterol level and regulates blood pressure.

5. Carbohydrates protect our muscles from cramping.

EXERCISE

Give short answers

i. Define carbohydrates.
ii. Give the characteristics of disaccharides.
iii. Give the balanced equation for the formation of glucose.
iv. Draw the structure of glucose.
v. Give the balanced equation for the hydrolysis of sucrose.
vi. How plants synthesize carbohydrates?
vii. Give the characteristics of monosaccharides.
viii. What is the difference between glucose and fructose?
ix. Give an example of a disaccharide. How is it hydrolyzed into monosaccharides?
x. Give the characteristics of polysaccharides.
xi. Describe the uses of carbohydrates.
xii. Lactose is disaccharide; which monosaccharides are present in it?

Extensive Questions

1. What are carbohydrates? How are monosaccharides prepared? Give their characteristics.
2. Explain oligosaccharides.
3. What are polysaccharides? Give their properties.
4. What are carbohydrates and how are they classified?
5. Point out one difference between the compounds in each of the following pairs.
(a) Glucose and fructose (b) Sucrose and maltose (c) Cellulose and starch

Chapter # 11

LIPIDS

STUDENTS LEARNING OUTCOMES:

- Structure and general properties of lipids.
- Classify lipids.
- Explain the sources and uses of lipids.
- Differentiate between fats and oils

11.1 Introduction:

Lipids (Greek, lipos means fat) are naturally occurring organic compounds of animals and plants origin which are soluble in organic solvents and belong to a very heterogeneous group of substances.

11.2 Characteristics of lipids:

Lipids have the following characteristics:

i.   They are insoluble in water and soluble in non-polar solvents e.g. ether, chloroform and benzene, etc.
ii. Their primary building blocks are fatty acids, glycerol, and sterols.
iii. They are utilized by living organisms.

11.3 Fats and oils

Fats and oils are the most important lipids found in nature. They are one of the three major "food factors" needed for the human body, the other two being proteins and carbohydrates. Fats and oils are widely distributed in various types of foods and are of great nutritional value. Not only do the edible fats and oils occupy a place of pride in the human diet but they also find use as raw materials for the manufacture of soaps and detergents, paints, varnishes, polishes, cosmetics, printing inks, and pharmaceuticals.

11.3.1 Sources of Fats and Oils

Fats and oils come from a variety of natural sources like animals, plants, and marine organisms. Animal fats are located particularly in adipose tissue cells. Butter and ghee are a special type of animal fats that are made from milk. Vegetable oils are chiefly present in the seeds and nuts of plants. Marine oils are obtained from sea animals like salmon and whales etc.

11.3.2 Structure and Composition of Fats and Oils:

Animal and vegetable fats and oils have similar chemical structures. They are triesters formed from glycerol and long-chain acids called fatty acids.
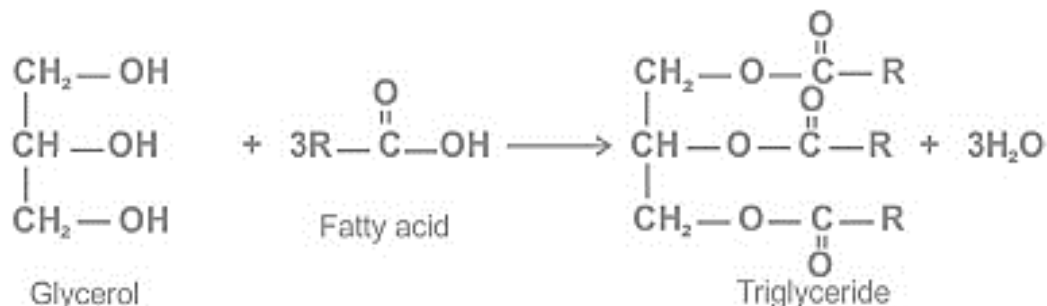
A triester of glycerol is called triglyceride or glyceride. The degree of unsaturation of the constituent fatty acid determines whether a triglyceride will be a solid or a liquid. The glycerides in which long- chain saturated acid components predominate tend to be solid or semisolid and are termed as fats. On the other hand, oils are glycerol esters which contain a higher proportion of unsaturated fatty acid components.

The melting points of mixed glycerides would depend on the extent of unsaturated fatty acid components in the molecule. The polyunsaturated glycerides therefore have very low melting points and are liquids (oils). Chemically common oils and fats are a mixture of saturated and unsaturated triglycerides, present in various ratios.

11.4 Classification:

Lipids are classified as:

a. Simple Lipids:

These are esters of fatty acids with glycerol. For example, common fats and oils.

b. Compound Lipids:

These contain radicals in addition to fatty acids and alcohol and include glycerol phospholipids, sphingolipids, lipoproteins, and lipopolysaccharides. The melting points of mixed glycerides would depend on the extent of unsaturated fatty acid components in the molecule. The polyunsaturated glycerides therefore have very low melting points and are liquids (oils). Chemically common oils and fats are a mixture of saturated and unsaturated triglycerides, present in various ratios.

c. Derived or Associated Lipids:

They are the hydrolytic products of the above-mentioned compounds. Sterols, vitamin D and terpenes belong to this class of lipids.

11.5 Physical Properties:

   i.   Oils and fats may either be liquid or non-crystalline solids at room temperature.

   ii.   When pure they are colourless, odourless and tasteless.

   iii.   They are insoluble in water and readily soluble in organic solvents like diethyl ether, acetone, carbon tetrachloride, and carbon disulphide.

   iv.   They readily form emulsions when agitated with H_2O in the presence of soap or other emulsifiers.

   v.   They are poor conductors of heat and electricity and therefore serve as excellent insulator for the animal body.

11.6 Chemical Properties

a) Hydrolysis

Triglycerides are easily hydrolyzed by enzymes called lipases to fatty acids and glycerol.
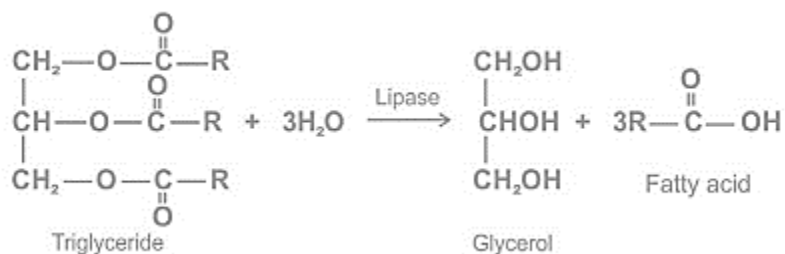
b) Saponification

It is the hydrolysis of a fat or an oil with an alkali to form soap (salt of fatty acid) and glycerol.
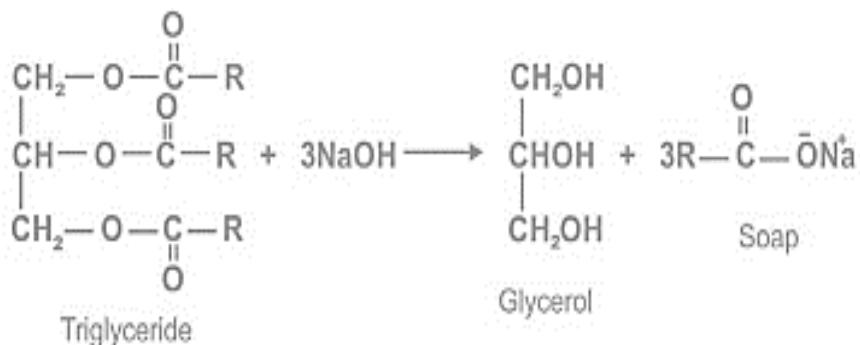
c) Hardening of Oils

Unsaturated glycerides react with hydrogen in the presence of a metal catalyst to give saturated glycerides. The result is the conversion of a liquid glyceride (an oil) into a semi-solid glyceride (a fat).

This reaction is used commercially to harden vegetable oils to produce vegetable ghee or margarine. Hardened oils are also extensively used for making soaps and candles.

d) **Saponification Number**

It is defined as the number of milligrams of potassium hydroxide or sodium hydroxide required to saponify one gram of the fat or oil. For example, one mole of glycerol tripalmitate (mol. wt = 807) requires 168,000 mg of KOH for saponification. Therefore, one gram of fat will require 168000/807 mg of KOH. Hence the saponification number of glycerol tripalmitate is 208.

e) **Rancidity of Fats or Oils**

Fats or oils are liable to spoilage and give off an odour known as rancidity. It is mainly caused by the hydrolytic or oxidative reactions which release foul smelling aldelydes and fatty acids. Oils from sea animals which contain a relatively high proportion of unsaturated acid chains deteriorate rapidly.

f) **Iodine Number**

The extent of unsaturation in a fat or an oil is expressed in terms of its iodine number. It is defined as the number of grams of iodine that will add to 100 grams of fat or an oil The value of iodine number depends on the number of double bonds present in the acid component of the glycerides. The glycerides with no double bonds have zero iodine number.

g) **Acid Number**

The acid number of a fat or an oil tells the amount of free fatty acids present in it. It is expressed as the number of milligrams of potassium hydroxide required to neutralize one gram of fat.

h) **Steroids**

Steroids are naturally occurring lipids. Their parent nucleus has perhydrocyclopentanophenanthrene component which consists of three six-membered rings (A, B, and C) and one five-membered ring (D). These rings are joined or fused and have a total of 17-C atoms Very small variations in the bonding of atoms in the ring and the groups attached to them give rise to compounds that are remarkably diverse in their biological functions. Some of the naturally occurring compounds belonging to steroids are cholesterol, ergosterol, male and female sex hormones, and the hormones of the adrenal cortex.
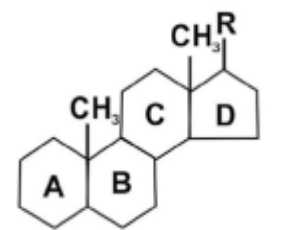
Structure of Steroid Nucleus

11.7 Importance of lipids

i. They are good source of energy and make the food more palatable.
ii. They exert an insulating effect on the nervous tissues.

iii. They are good energy reservoirs in the body.
iv. Lipids are an integral part of cell protoplasm and cell membranes.
v. Some lipids act as precursors of very important physiological compounds. For example, cholesterol is the precursor of steroid hormones.

EXERCISE

Give short answers.

    i.   What is the difference between ghee and oil?
    ii.   Give the characteristics of fats.
   iii.   Give the sources and uses of animal fats.
   iv.   Plants are source of oils, justify.

Extensive Questions

1. Explain the sources and uses of lipids.
2. What are lipids? In what way fats and oils are different?
3. Define saponification number and iodine number. Discuss the term rancidity.

Chemistry Practicals.

Experiment #1 :How fitting up a wash bottle prepared?

Material Required:

Water bottles, Water, Dish soap, Safety goggles, Gloves, Rubber band.

Procedure:

  i. Ensure you have the required safety equipment, such as gloves and safety goggles, if handling chemicals.
 ii. Verify that the wash bottle is clean and dry.
iii. Remove the cap or lid. Insert a short piece of glass or plastic tubing into a rubber stopper.
 iv. Place a funnel into the top of the rubber stopper. Insert the rubber stopper with the tubing into the neck of the wash bottle.
  v. Create a tight seal around the neck of the wash bottle using a grommet or rubber washer.
 vi. Attach one end of the rubber or plastic tubing to the tubing in the rubber stopper. Secure the wash bottle to a lab stand or ring stand to prevent tipping.
vii. If necessary, gently heat the wash bottle using a hot plate or heat source. Use the funnel to pour the desired liquid into the wash bottle.
viii. Dispense the liquid by gently squeezing the wash bottle.
 ix. Test the wash bottle to ensure it dispenses the liquid correctly and adjust the tubing or nozzle if needed.
  x. Label the wash bottle with the contents and safety information. Store the wash bottle in a safe location, away from direct sunlight or heat sources.

Experiment # 2: To pacify the given sample of imposed naphthalene crystallization.

Material Required:

Beaker, Funnel, Filter paper, Burner, Tripod stand, Wire gauze, Stirrer, Thermometer.

Chemicals Required:

Distilled water, Given the mixture of impure naphthalene and phosphoric acid.

Procedure:

i. Ensure safety gear (gloves, goggles) is worn and work in a well-ventilated area.
ii. Weigh the impure naphthalene sample accurately and record the initial mass. Place the impure naphthalene in a beaker.
iii. Add a non-polar solvent (e.g., petroleum ether) to dissolve the impurities while keeping naphthalene in solution.
iv. Heat and stir gently.
v. Set up a filtration apparatus with filter paper or a Büchner funnel and filter flask.
vi. Filter the hot solution to remove insoluble impurities.
vii. Transfer the clear naphthalene solution to a dish. Allow the solution to cool slowly, facilitating crystallization.
viii. Collect the naphthalene crystals and drain off any remaining liquid. Air-dry the crystals completely.
ix. Weigh the dried crystals accurately and record the final mass.
x. Calculate the percentage yield and purity of the purified naphthalene.
xi. Dispose of waste materials properly and clean glassware and equipment.

$$Percentage\ Yield = \frac{Theoretical\ Yield}{Actual\ Yield} \times 100\%$$

$$Purity = \frac{Total\ Mass\ of\ the\ Obtained\ Crystals}{Actual\ Mass\ of\ Purified\ Naphthalene} \times 100\%$$

xii. By following these steps and calculations, you can determine both the percentage yield and the purity of the purified naphthalene obtained in your experiment.
xiii. Weigh the dried naphthalene crystals accurately, and record this value as the "Actual Yield."
xiv. Calculate the theoretical yield by assuming 100% purity of the impure sample.
xv. If you had "X" grams of impure naphthalene initially, then the theoretical yield would be "X" grams. Use the values from step 1 and step 2 to calculate the percentage yield using the formula mentioned earlier.
xvi. To calculate the purity, you need to know the mass of the impurities in the obtained crystals.

xvii. If you don't have this information, you can conduct further testing or analysis to determine the purity more accurately. However, the purity calculation is typically done by analyzing the product using techniques like spectroscopy or chromatography.

xviii. Once you have the mass of the impurities, use the formula for purity to determine the purity of the purified naphthalene.

Experiment # 3: Separate Naphthalene from the given mixture of sand and naphthalene by sublimation.
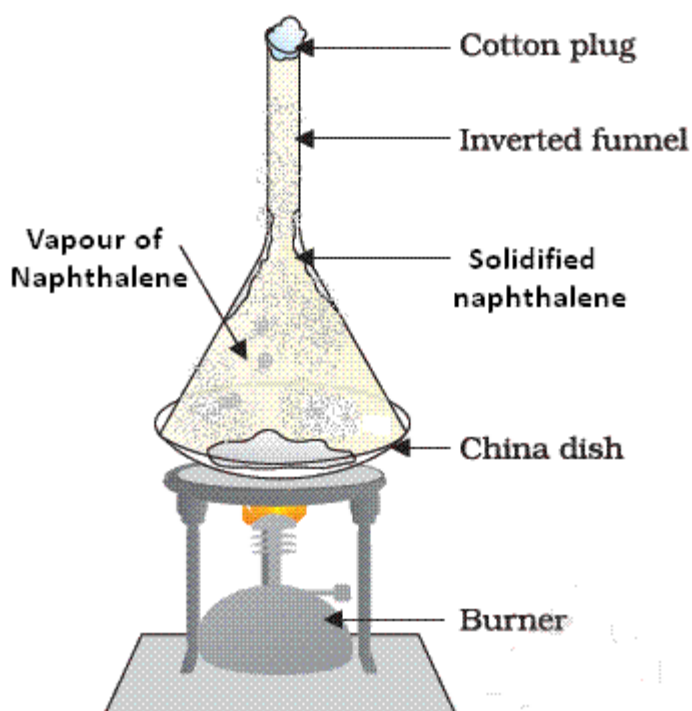
Material Required:

Evaporating dish, burner, funnel, filter paper, cotton, tripod stand, sand bath

Chemicals Required:

Given mixture of sand and naphthalene.

Procedure:

i. Take about 5g of the mixture on a China dish and place it on the sand bath.

ii. Fold filter paper in a cone shape and moisten the inner side of the paper with distilled water. Place the filter paper on the funnel.

iii. Close the open end of the funnel's stem with cotton.

iv. Invert the funnel on the China dish so that the vapors of the given mixture don't escape the system.

v. Heat the sand bath for 5-10 minutes by using the burner.

vi. Now, carefully lift the funnel, and remove the filter paper out of the funnel.

vii. Observe the crystals of pure naphthalene stuck on filter paper.

Result:

Given organic compound (Naphthalene) is separated from the mixture.

Experiment #4: Determine the melting point of the given compound (Naphthalene).
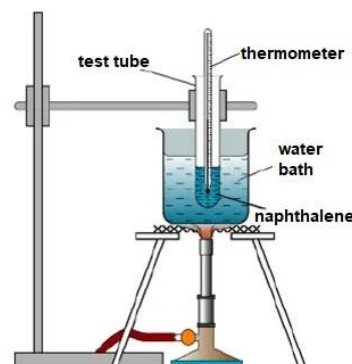
Material Required:

Thermometer, Beaker, Thread, Capillary tube, Tripod stand, wire guaze, Stand with clamp.

Chemicals Required:

Water and given compound (Naphthalene)

Procedure:

i. Take a capillary tube and close its one end by heating the end in the flame for 2-3 minutes while continuously rotating it.
ii. Take naphthalene on a tile and crush it into a fine powder.
iii. Stuff some naphthalene into the capillary tube and strap a thermometer to it.
iv. Immerse the capillary tube in a hot water bath.
v. As the temperature of water rises, at a particular temperature the naphthalene melts out of the capillary tube.
vi. Make note of the temperature (t1) as soon as the compound begins to melt.
vii. Make note of the temperature (t2) as soon as the compound is melted completely.
viii. The average of the two readings gives the correct melting point of the substance.

Observations:

No of observations	Melting point in °C
1	
2	

Mean value=T1+T2/2

Result:

The melting point of the given organic compound is _____°C.

Experiment # 5: Determine the Boiling point of the given compound (Acetone).
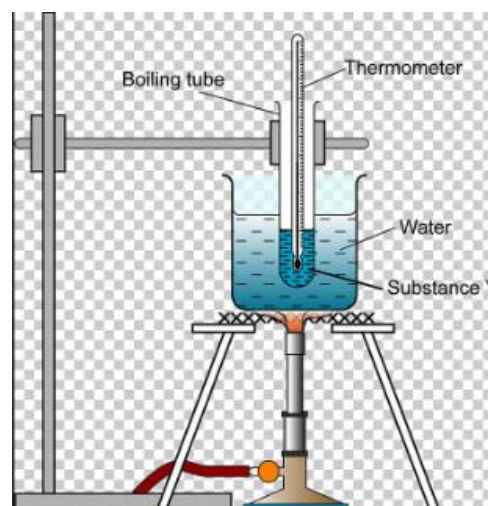
Material Required:

Thermometer, Beaker, Thread, Fusion tube, Tripod stand, wire guaze, Stand with clamp and stirrer.

Chemicals Required:

Water and given compound (Acetone)

Procedure:

i. Close one end of the capillary tube by gently heating on oxidizing flame.

ii. Attach the fusion tube with the thermometer by using thread in such a way that the bulb of the thermometer and base of the fusion tube are at the same level.

iii. Hang the thermometer along with the fusion tube using thread with the clamp of the stand.

iv. Take 250 cm^3 beaker and add approximately 100 cm^3 of water in it, place this beaker on tripod stand over the wire guaze.

v. Now dip the thermometer along with the fusion tube onto the beaker of water.

vi. Fill one-third of the fusion tube with Acetone with the help of dropper.

vii. Heat the beaker with continuous stirrer.

viii. Note the temperature at which continuous steam of bubbles start to come out.

Observations:

No of observations	Boiling point in °C
1	
2	

Mean value=T1+T2/2

Result:

The Boiling point of the given organic compound is _____°C

Experiment# 6: Prepare 100cm3 of 0.1M sodium hydroxide (NaOH) solution.
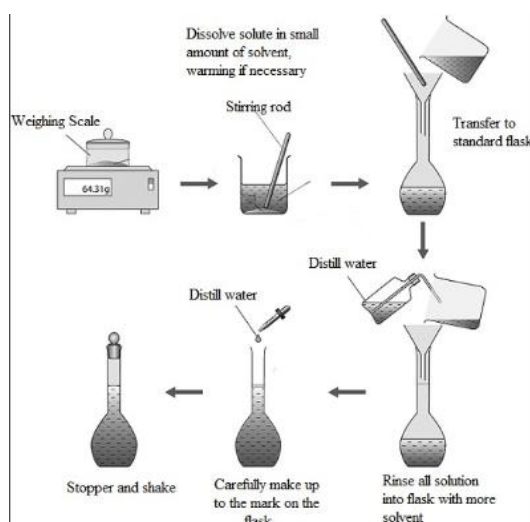
Material Required:

Top loading balance, Beaker, Measuring flask with stopper, Funnel, Spatula, watch glass.

Chemicals Required:

Distill Water and solid NaOH.

Procedure:

First calculate the mass of NaOH required to make 0.1M NaOH solution.

Molar mass of NaOH = 23+16+1 = 40 g/mol

Volume of NaOH required in dm^3 = $\frac{100}{1000}$ x1dm30 = 0.1 dm^3

No pf moles of NaOH required = molarity of NaOH required x Volume of NaOH required in dm^3

= o.1 x 0.1 = 0.01 mol

Mass of NaOH required in g = moles of NaOH x Molar Mass

= 0.01 x 40 = 0.4 g

Set the top loading balance at 0.00g.

Place dry watch glass on top loading balance and tare the reading.

Now weigh 0.4g NaOH on the watch glass by using a spatula.

Transfer NaOH into the beaker, wash the watch glass with distilled water, and transfer it to the beaker.

Now add about 20 cm^3 of water in beaker and stir the solution so NaOH dissolve in distill water completely.

Transfer the solution into measuring flask (100cm^3) with the help of funnel.

Add more water to raise the level of the solution below the mark of measuring flask.

Now add distill water carefully with the help of dropper till the lower meniscus of the solution.

Close the measuring flask and turn it upside down to ensure the solution is thoroughly mixed.

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Experiment # 7: To prepare a standard solution of oxalic acid and with its help standardize a solution of NaOH.

Material Required:

Top loading balance, Beaker, measuring flask with stopper, Funnel, Spatula, watch glass, Burette, pipette, Measuring cylinder.
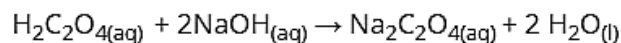
Chemicals Required:

Distill Water and solid Oxalic Acid, 0.1M NaOH solution, 2-3 drops of phenolphthalein indicator.

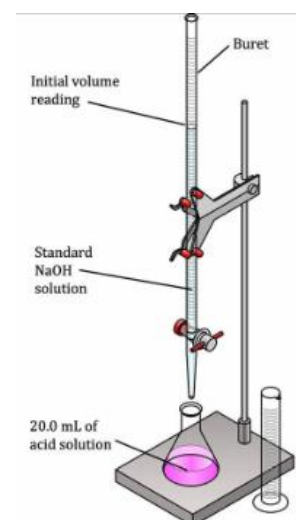**Principle**: Acid-Base Titration.

Indicator: Phenolphthalein

End-point: colorless to light pink

Chemicals Reaction:

$$H_2C_2O_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2C_2O_{4(aq)} + 2\ H_2O_{(l)}$$

Mole Ratio:

1:2

Preparing the Standard Solution of Oxalic Acid:

i. Measure approximately 4 grams of oxalic acid dihydrate (H2C2O4·2H2O) using an analytical balance.
ii. Record the exact mass.
iii. Dissolve the measured oxalic acid in about 300 mL of deionized or distilled water in a clean beaker.
iv. Stir the solution until all the oxalic acid has dissolved.
v. This solution contains approximately 0.1 M (mol/L) of oxalic acid.
vi. Carefully transfer the oxalic acid solution to a clean, dry, 500 mL volumetric flask using a funnel.
vii. Rinse the beaker with deionized water and add the rinsings to the flask.
viii. Fill the flask up to the mark with deionized or distilled water. Mix the solution thoroughly by inverting the flask several times. This is your standard solution of oxalic acid.
ix. Standardizing the Sodium Hydroxide (NaOH) Solution:

x. Rinse a burette with a small amount of the sodium hydroxide solution.
xi. Then, fill the burette with the sodium hydroxide solution.
xii. Use a pipette to measure exactly 10 mL of the prepared standard oxalic acid solution and transfer it to a clean conical flask.
xiii. Add a few drops of phenolphthalein indicator to the flask.
xiv. The solution should turn colorless.
xv. Place the flask under the burette, and titrate the oxalic acid solution by slowly adding sodium hydroxide solution drop by drop while swirling the flask continuously.
xvi. Continue adding the NaOH solution until the color changes from colorless to a pale pink that persists for at least 30 seconds.
xvii. This indicates that all the oxalic acid has reacted with the sodium hydroxide.
xviii. Record the volume of sodium hydroxide solution used from the burette.
xix. This volume represents the amount of NaOH required to neutralize the known amount of oxalic acid.

Observations:

No. of Obs.	Initial reading	Final Reading	Mean
1	0	10	$10 cm^3$
2	10	20	$10 cm^3$
3	20	30	$10 cm^3$

Mean volume= $10 cm^3$

Calculations:

Given   :   Required

Oxalic acid   :   NaOH

M_1V_1 / n_1 = M_2V_2 / n_1

     M_1 = $M_2V_2n_1 / V_1n_2$

$M_1 = 0.1 \times 10 \times 1 / 10 \times 2$

=0.05M

Result:

The   molarity   of   the   given   solution   is   _____

Experiment # 8: To prepare approximated N/10 solution of H2SO4 determine its exact normality by titrating it against standard N/10 NaoH.

Material Required:

Top loading balance, Beaker, measuring flask with stopper, Funnel, Spatula, watch glass, Burette, pipette, Measuring cylinder.

Chemicals Required:

Distill Water and 0.1M H_2SO_4, 0.1M NaOH solution, 2-3 drops of phenolphthalein indicator.

Principle: Acid-Base Titration.

Indicator: Phenolphthalein

End-point: Light pink to colorless.

Chemicals Reaction:

2 NaOH + H_2SO_4 → Na_2SO_4 + 2 H_2O

Mole ratio:

2:1

Procedure:

i. Pipette out 10mL of the 0.1M NaOH and transfer it into a clean conical flask.
ii. Add 1-2 drops of phenolphthalein indicator solutions turns purple.
iii. Take oxalic acid in burette and note the initial reading.
iv. Add oxalic acid dropwise from the burette into titrating flask.
v. Titrate it with 0.1N Sulphuric acid solution until colour changes from purple to colorless.
vi. Difference between the two readings give the volume of acetic acid used.
vii. Repeat the experiment to get three concordant readings.
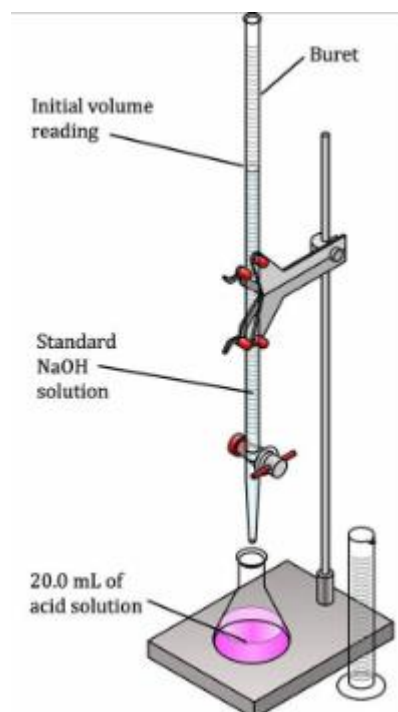
No. of Obs.	Initial reading	Final Reading	Mean
1	0	10	10cm^3
2	10	20	10 cm^3

3	20	30	10 cm^3

Mean volume= 10 cm^3

Calculations:

Given : Required

H_2SO_4 : NaOH

$M_1V_1 / n_1 = M_2V_2 / n_1$

$M_1 = M_2V_2n_1 / V_1n_2$

$M_1 = 0.1 \times 10 / 10 \times 2$

=0.05M

Result:

The normality/molarity of the given solution is _____

Precautions:

Wash the watch glass carefully so that even a single crystal is not left on the watch glass.

Bring the watch glass close to the funnel while transferring the weighed substance and transfer it gently. Wash it repeatedly with distilled water.

Wash the burette with water after titration is over.

The last few drops should be added using a pipette to avoid the extra addition of distilled water above the mark on the neck of the measuring cylinder.

Experiment# 9: Determine the amount of acetic acid in the 100cm3 of vinegar sample.

Principle: Acid-Base Titration.

**Indicator:** Phenolphthalein

End-point: Light pink to colourless.

Standard solution: 0.05 M NaOH

Chemical Equation:

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \longrightarrow CH_3COO^-Na^+{}_{(aq)} + H_2O_{(aq)}$$

Procedure:

1. Take 5cm^3 of vinegar in 100 cm^3 measuring flask and dilute upto mark with distilled water, shake it well.
2. Pipette out 10cm^3 of NaOH solution into titration flask.
3. Add a drop of indicator.
4. Take vinegar (Acetic acid) in burette and note the initial reading.
5. Add acetic acid drop wise from the burette into conical flask.
6. Keep on titration NaOH solution with acetic acid until light pink colour appear.
7. Difference between the two readings give the volume of acetic acid used.
8. Repeat the experiment to get three concordant readings.
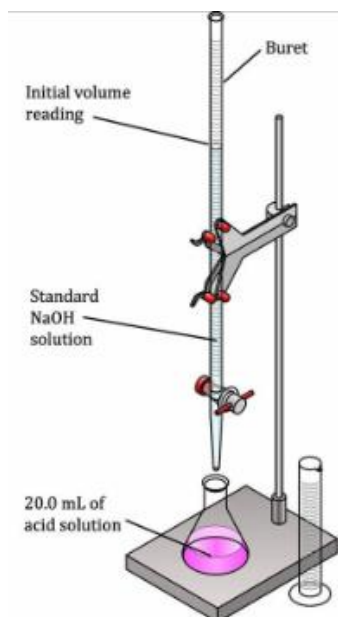
Observations:

No. of Obs.	Initial reading	Final Reading	Mean
1	0	10	10cm^3
2	10	20	10 cm^3
3	20	30	10 cm^3

Mean volume= 10 cm^3

Calculations:

Given : Required

Acetic acid : NaOH

$$M_1V_1 / n_1 = M_2V_2 / n_1$$

$$M_1 = M_2V_2n_1 / V_1n_2$$

$$M_1 = 0.05 \times 10 \times 1 / 10 \times 1$$

=0.05M

Strength of Acetic Acid = molarity x mol. Weight

$= 0.05 \times 60 = 3.0 \text{ g/ dm}^3$

1000cm^3 of vinegar contains Acetic acid = 3.0g

100cm^3 of vinegar contains Acetic acid = 3 x 100 / 1000 = 0.3g

5cm^3 of vinegar contains Acetic acid = 0.3g

100cm^3 of vinegar contains Acetic acid = (0.3 x 100) 5 = 6g

Result:

Experiment # 10: Standardize the given solution of KMNO4 volumetrically. Calculate the volume of KMNO4 required to prepare 1.0 dm3 of 0.01M KMNO4 solution.
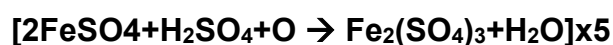
Principle: Oxidation- Reduction Titration.

Indicator: $KMNO_4$ Itself

Endpoint: Light pink to colourless.

Standard solution: 0.1 M $FeSO_4$

Chemical Equation:

$2KMnO_4+3H_2SO_4 \rightarrow K_2SO_4+2MnSO_4+ 3H_2O+5 (O)$

$[2FeSO4+H_2SO_4+O \rightarrow Fe_2(SO_4)_3+H_2O]x5$

$2KMnO_4+ 8H_2SO_4 + 10FeSO_4 \rightarrow 5Fe_2(SO_4)_3+K_2SO_4+2MnSO_4+8H_2O$

Procedure:

1. Pipette out $10cm^3$ of $FeSO_4$ solution into conical flask.
2. Add half test tube of dilute H_2SO_4 into the flask.
3. Take $KMNO_4$ solution in burette and note the initial reading.
4. Add $KMNO_4$ solution drop wise from the burette into conical flask.
5. Keep on titration $FeSO_4$ solution with $KMNO_4$ solution until light pink colour appears.
6. Difference between the two readings give the volume of acetic acid used.
7. Repeat the experiment to get three concordant readings.
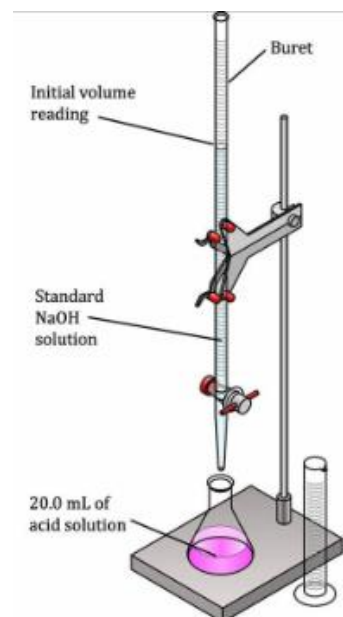
Observations:

No. of Obs.	Initial reading	Final Reading	Mean
1	0	10	$10cm^3$

2	10	20	10 cm^3
3	20	30	10 cm^3

Mean volume= 10 cm^3

Calculations:

Given: Required

KMNO$_4$: FeSO$_4$

M$_1$V$_1$ / n$_1$ = M$_2$V$_2$ / n$_1$

M$_1$ = M$_2$V$_2$n$_1$ / V$_1$n$_2$

M$_1$ = 0.1 X 10 x2 / 10 x10

=0.02M

Given: Required

KMNO$_4$: KMNO$_4$

M$_1$V$_1$ = M$_2$V$_2$

0.02XV$_1$ =0.01X1000

V$_1$ = 0.01 X 1000 / 0.02 = 500cm^3