CHAPTER 2

THE THREE STATES OF MATTER
GASES, LIQUIDS AND SOLIDS

INTRODUCTION

Matter is defined as any thing which occupies space and has mass. The three states of matter-gas, liquid and solid are easily recognized through their properties.

A gas has no shape of its own; rather it takes the shape of its container. It has no definite volume but can be compressed or squeezed into smaller volume. A gas diffuses, that is, it distributes itself through out a vessel.

A liquid has no definite shape; it takes the shape of its container. It however occupies a definite volume of its own. Although it is not absolutely in-compressible, it is compressed to a negligible extent even by high pressures.

A solid possesses both definite shape and definite volume. Like a liquid, it is not compressed or squeezed appreciably by high pressures.

The same substance, say water, can exist in all the three states, for example, as solid ice, liquid water and the gaseous steam. Despite the same chemical composition, water shows different behaviour in three different states. The only conclusion which can be drawn,is that the particles which make up the three states of matter for any given substance differ only in the manner in which they are arranged and they are not different in kind. The theory which deals with the way in which the arrangement of particles of a substance determines the properties that substance will possess, and particularly the state in which it is likely to be found under a given set of conditions is known as the kinetic theory.

2.1 KINETIC THEORY

The Kinetic theory is so named because it deals with that property of the particles which is so crucial in understanding the three states of matter, namely the motion of the particles. Because the particles are in constant motion, they possess kinetic energy which tends to keep the particles well spaced out in any substance. Thus in terms of kinetic theory.

1. The gaseous state is the one in which the molecules are widely separated from one another but having negligible volume. The gases are easy to compress due to large empty spaces.

2. The gas molecules are in continuous motion, travelling in straight paths between collisions, but in random directions (Fig. 2.1). The average distance between collisions is the mean free path.

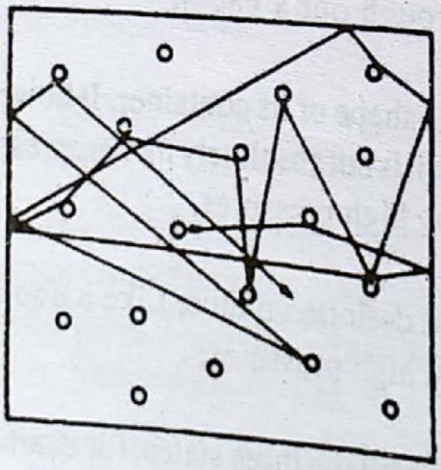


FIGURE 2.1
The random motion of a gas molecule. The average distance between collisions is the mean free path.

3. The molecules collide with one another and with the walls of container, but these collisions are perfectly elastic (result in no loss of energy).

4. Gas pressure is the result of the collision of gas molecules with the walls of the container (fig. 2.2).

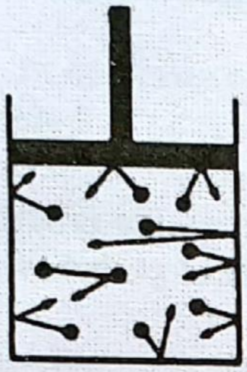


FIGURE 2.2
Gas pressure is the result of the collision of the gas molecules with the walts of the container.

5. In an ideal gas, there are no attractive or repulsive forces between molecules. Thus each molecule acts quite independently of the others.

6. The average kinetic energy of gas molecules depends upon the absolute temperature. At any given temperature, the molecules of all gases have the same average kinetic erergy ($^1/_2\,mv^2$).

The kinetic theory was extended to account for the behaviour of liquids and solids.

For a liquid, the kinetic theory suggests that the particles of a liquid are fairly randomly arranged (as in gas), but consists of clusters' in which they are very close together. This makes liquid have a definite volume, but since the particles are still fairly free to move, it does not have any definite shape of its own.

For a solid, the kinetic theory postulates that the molecules are closely packed, so that the forces of attraction between the molecules are very strong and free movement of particles can not take place. Thus in a solid, the particles are arranged in a fixed pattern and they form a lattice of vibrating masses. This makes a solid have a definite shape.

2.2 GASEOUS STATE

We select to begin our study of matter with the gaseous state because (i) it is the simplest to study and (ii) it can be readily converted into the condensed states – liquid and solid; hence many of our conclusions concerning gases apply directly to these states.

BEHAVIOUR OF GASES

The assumptions of kinetic theory account for most of the properties associated with gases and we can obtain better understanding of gas behaviour.

1. Diffusibility

The distribution or spreading of gas molecules through out the vessel is known as diffusion. Unlike liquids or solids, the gases diffuse very rapidly. A drop of perfume for instance, slowly evaporates out the fragrant gas announces the presence of wearer. It is due to the diffusions of perfume through the air.

In terms of kinetic theory, diffusion is explained as follows. The molecules of a gas are widely separated and there are large empty spaces due to which they are free to move. Due to this free movement the molecules of gases intermingle and spread out easily throughout the vessel. The opposite of diffusions is effusion in which a gas passes through the pores or tiny holes in the vessel (fig. 2.3), for example, the air effuses from the tire as a result of which the tire loses pressure gradually.

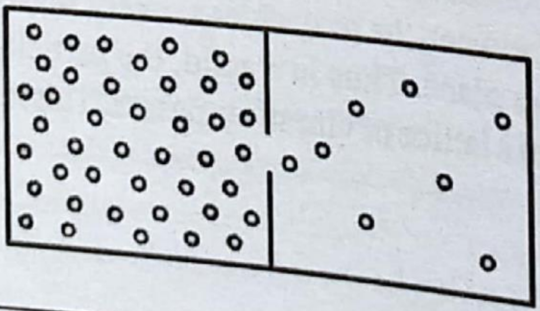


FIGURE 2.3
Effusion of gas through a small hole.

2. Compressibility:

In contrast to liquids or solids, all the gases are easily compressed or squeezed. In terms of 'kinetic theory', gases are easily compressed due to large empty spaces. By applying pressure, the molecules come closer. Air, for instance is squeezed into automobile tires. When the tire is punctured, the air rushes out. It is the reverse of compressibility – the expansibility, thus volumes of gases are highly affected by the changes in pressure. [Volume is measured in cubic decimetres, dm^3 or cubic centimetres, cm^3. [1 dm^3 = 1000 cm^3]

3. Pressure:

All the gases exert pressure. It may appear surprising but it is a fact that we are being pressed upon by an enormously heavy blanket of atmosphere. The mass of atmosphere on our body at seal level and at 0°C is 14.7 psi (pounds per square inch, lb/in^2) or there is about 20 tons total pressure on our bodies.

When a gas is confined in a closed container, it exerts pressure on the walls of the container which is due to the collisions of gas molecules with the walls The tires of automobiles are filled with air until the guage shows the pressure of about 28 psi. This means that the pressure inside the tire is 28 psi greater than the outside pressure; since the external atmospheric pressure is 14.7 psi, hence the total pressure on the tire is 28 + 14.7 = 42.7 psi.

Since pressure is defined as a force pushing on a unit of area, therefore pressure may be measured in psi (pounds per square inch, lb/in^2), Kilograms per square metre (Kg/m^2). The S.I. unit of pressure is pascal (1 pa = $1N/m^2$). Since the units of a newton of force are $Kg/m.s^2$, the S.I. units of pressure are:

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}} = \frac{\text{newton}}{m^2} = \frac{Kg\ m/s^2}{m^2} = Kg/\ m\ s^2$$

Normal atmospheric pressure at sea level at 273 K (0°C) is expressed in several ways.

14.7 psi; 760 mm of mercury = 760 torr (∴ 1 mm = 1 torr);

76 cm = 1 atmosphere (atm). Since they are more useful for our purpose than the S.I. units, therefore we will use these units).

Problem 1 : What is the pressure in torr of 3.5 atm?

Solution: ∴ 1 atm = 760 torr

$$\therefore\ 3.5\ \text{atm} = \frac{760\ \text{torr}}{1\ \text{atm}} \times 3.5\ \text{atm} = 2660\ \text{torr}$$

Problem 2 : How many atmospheres correspond to 950 torr?

Solution : ∴ 760 torr = 1 atm

$$\therefore\ 950\ \text{torr} = \frac{1\ \text{atm} \times 950\ \text{torr}}{760\ \text{torr}} = 1.25\ \text{atm}$$

Temperature may be defined as that property which determines the flow of heat. The common scale of temperature is Celsius (Centigrade). The S.I. unit is however, Kelvin (K).

$$K = {}^{\circ}C + 273$$

e.g. 7°C = 7 + 273 = 280 K.

2.3 GAS LAWS

The mass, volume, temperature and pressure of gases bear a simple mathematical relationship to each other. Further, other variables, such as rates of diffusion of different gases are related mathematically. The precise statements of these relationships are known as the gas laws.

BOYLE'S LAW

Robert Boyle studied the effect of pressure changes on the volume of the gases. For this purpose, he designed a J - shaped glass tube and its shorter arm was closed. (Fig. 2.4). He then poured mercury in the longer arm so as to enclose some amount of air in the shorter arm. The equal level of mercury in the two arms indicated that the confined air was subjected to the atmospheric presure Fig: 2.4(a). Boyle added more mercury as a result of which the mercury level raised above the original level (h) and the increased pressure reduced the volume of the air Fig:. 2.4(b). The total pressure was calculated as:

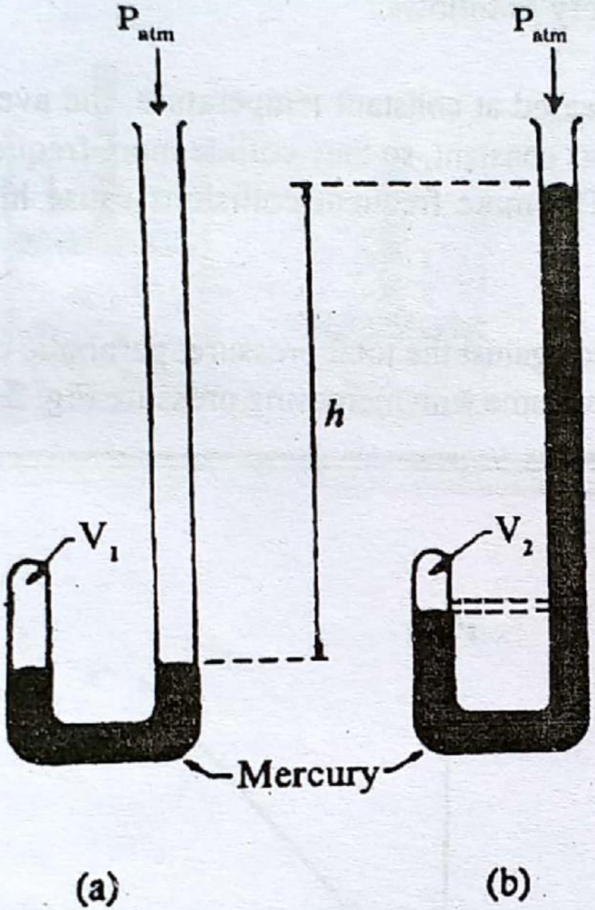


FIGURE 2.4
Boyle's J-tube experiment. Adding mercury to the open arm of the tube compresses the trapped sample of air to a smaller volume.

$P_{total} = P_{atm} + P_h$, where P_h = excess pressure.

Boyle compared the volume of the air with the total pressure and observed the inverse or non-linerar relationship:

"The volume of a given mass of a gas is inversely proportional to the pressure exerted on it at a given temperature".

This relationship reported in 1660 is known as Boyle's law. According to this law, at constant temperature, if pressure is doubled, volume will be reduced to half, if pressure is reduced to half, the volume will be doubled. Boyle's law can be explained in terms of kinetic theory as follows :

If the volume of a gas is decreased at constant temperature, the average velocity of the gas molecules remains constant, so they collide more frequently with the walls of smaller vessel. The more frequent collisions cause higher pressure.

If the volume of a gas is plotted against the total pressure, parabolic curve is obtained showing the decrease in volume with increasing pressure Fig: 2.5(a).

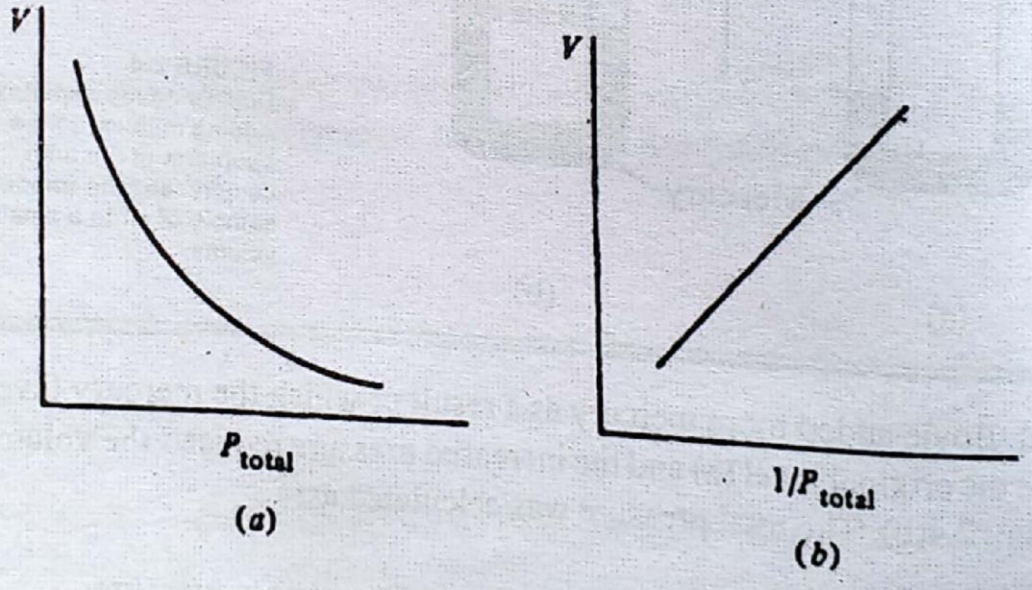


FIGURE 2.5
Graphical representation of Boyle's law. (a) A plot of volume versus total pressure is nonlinear showing that the relationship is inverse. (b) A plot of volume versus the reciprocal of the pressure is a straight line.

A graph of volume with I/P gives a straight line Fig. 2.5(b) The Boyle's law i interpreted in terms of molecules in the Fig. 2.6. Applying the pressure brings th molecules closer together causing the reduction in volume. Mathematically Boyle's law can be represented as :

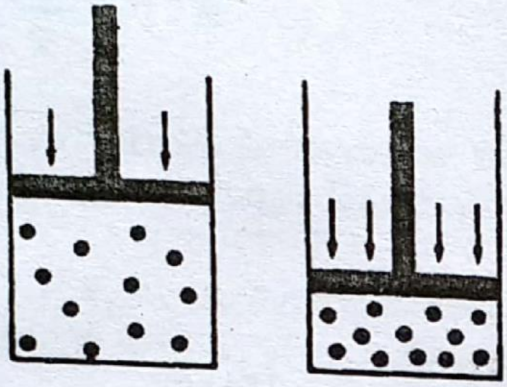


FIGURE 2.6
Increasing the pressure on a mass of gas forces the molecules closer together.

$$V \propto \frac{1}{P} \quad \text{(Temperature constant)}$$

$$V = K \times \frac{1}{P}$$

PV = K i.e. the product of pressure and volume of a given gas at constant temperature is always constant. If P_1 and V_1 are the initial pressure and volume, P_2 and V_2 are the changed pressure and volume, then according to Boyle's law:

$$P_1 V_1 = K; \; P_2 V_2 = K \text{ i.e.}$$

$$P_1 V_1 = P_2 V_2$$

This is the mathematical expression of Boyle's law which is used while solving the problems on Boyle's law.

Problem : What volume does 400 cm³ sample of a gas at 700 torr occupy when the pressure is changed to 2 atm?

Solution: Tabulate the data. (700 torr = 700/760 = 0.92 atm)

	P	V	T
Original	0.92 atm	400 cm^3	K
Changed to	2.0 atm	V	K

According to the Boyle's law,

$$P_1 V_1 = P_2 V_2$$

$$V_2 = \frac{P_1 V_1}{P_2}$$

$$= \frac{0.92 \text{ atm} \times 400 \text{ cm}^3}{2 \text{ atm}} = 184 \text{ cm}^3$$

CHARLES S LAW:

It is a matter of common observation that substances expand on heating and contract on cooling. Change in volume due to expansion and contraction is very small in case of solids and liquids because of their compact structure and lack of inter molecular spaces. Gases exhibit enormous changes due to the presence of large intermolecular spaces. Changes of volumes of gases due to changes of temperature at constant pressure were studied by charles (1746-1823).

The volume of a given mass of a gas increases or decreases by $\frac{1}{273}$ times of its original volume at $0^{\circ}C$ for every degree rise and fall o temperature at a given pressure. It is also stated as the volume of a given mass of a gas is directly proportional to the absolute temperature at given pressure.

In terms of kinetic theory, Charles's law can be explained as follows: A decrease in temperature, decreases the average kinetic energy of the ga molecules, that is, average molecular velocity decreases. At constant pressure the decreased velocity causes the sample of gas to shrink and occupy a smalle volume.

Mathematically this law can be represented as :

$$V \propto T \text{ (absolute)} \quad \text{(P is constant)}$$

$$V = KT$$

$$\frac{V}{T} = K$$

In other words the ratio of volume of a given mass of a gas to its absolute temperature is always constant provided the pressure is kept constant.

If V_1, V_2 are the original and changed volumes; T_1, T_2 are the original and changed temperatures, then according to the Charles's law.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\therefore \quad \frac{V_1}{T_1} = \frac{V_2}{T_2} = K$$

This is the mathematical representation of Charles 's law which is used while solving the problems on Charles's law.

The graphical representation of the observation of Charles is shown in Fig. 2.7. When the volume of the gas is plotted against the temperature, a straight line is obtained. If this straight line is extra plotted, it will be seen that it intercepts the temperature axis at –273.16°C (Fig. 2.7)

It suggests that the volume of a gas would theoretically be zero at –273.16°C. In fact this temperature has never been achieved for any gas. In fact all the gases condense to liquid at a temperature above this point, so that gas behaviour can no longer be observed. The temperature of –273.16°C is referred to as the *Absolute zero* or zero degree of the *absolute scale* (Kelvin Scale). At absolute zero, the volume of gases is considered to be zero and all the motion ceases to exist. As already pointed out that centigrade scale is changed into Kelvin scale by adding 273.

°C + 273 = K
K − 273 = °C

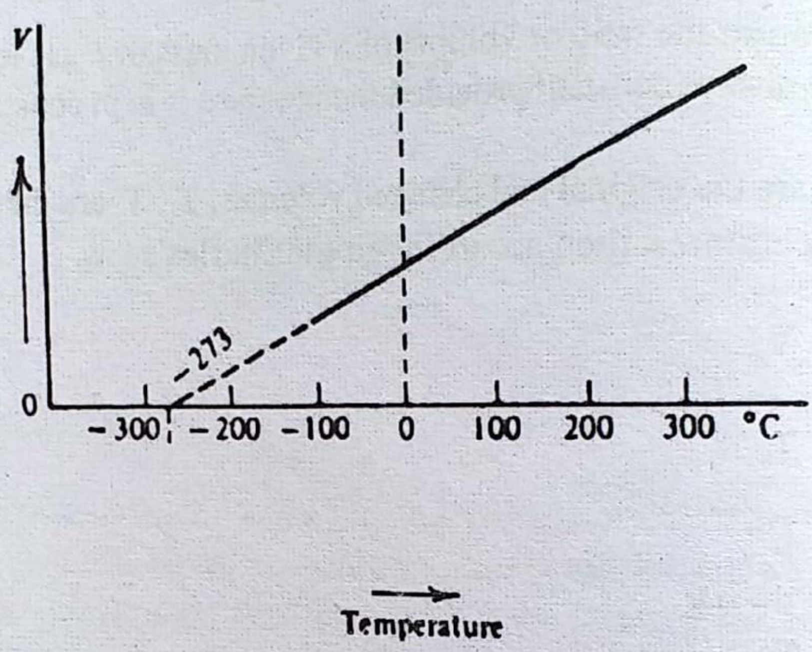


FIGURE 2.7
Graphical representation of the volume temperature relationship of a sample of gas. The gas volume decreases with decreasing temperature.

Problem.1 A child's balloon has a volume of 3.80 dm^3 when the temperature is 35°C. What is the volume, if the balloon is put into a refrigerator and cooled to 5°C? Assume that pressure inside the balloon is equal to atmospheric pressure at all times.

Solution : Tabulate the data.

Original

$V_1 = 3.80\ dm^3$, $T_1 = 35+273 = 308\ K$, $P = K$

Changed to $V_2 = ?$, $T_2 = 5+273 = 278\ K$, $P = K$

According to Charles's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1 T_2}{T_1}$$

$$= \frac{3.8\ \text{dm}^3 \times 278\ \text{K}}{308\ \text{K}} = 3.42\ \text{dm}^3$$

AVOGADRO'S LAW :

In 1811, Amadeo Avogadro advanced a brilliant hypothesis regarding the relationship between the volume and the number of molecules of a gas. This hypothesis, now called Avogadro's law which states:

"Equal volumes of all gases at the same temperature and pressure contain the same number of molecules". Thus the volume of a gas is directly proportional to the number of molecules of the gas at constant temperature and pressure.

i.e $V \propto n$ (at constant temperature and pressure) where 'n' is the number of molecules of a gas. Thus 1 dm^3 (or cm^3, m^3 or other unit of volume) of oxygen contains the same number of molecules as 1dm^3 (or cm^3 etc.) of hydrogen or of any other gas, provided the volumes are measured under the same conditions.

Based on Avogadro's law, 22.4 dm^3 of any gas at S.T.P. (standard temperature is 273 K and standard pressure is 1 atmosphere) constitutes 1 mole of that gas; 22.4 dm^3 at S.T.P. is called the molar gas volume or the volume of one mole of a gas at S.T.P.

The mass in grams of 22.4 dm^3 of a gas at S.T.P. is the mass of one mole of that gas, e.g. 22.4 dm^3 of SO_3 at S.T.P. is equal to 80 g (one mole). It is shown in the fig. 2.8.

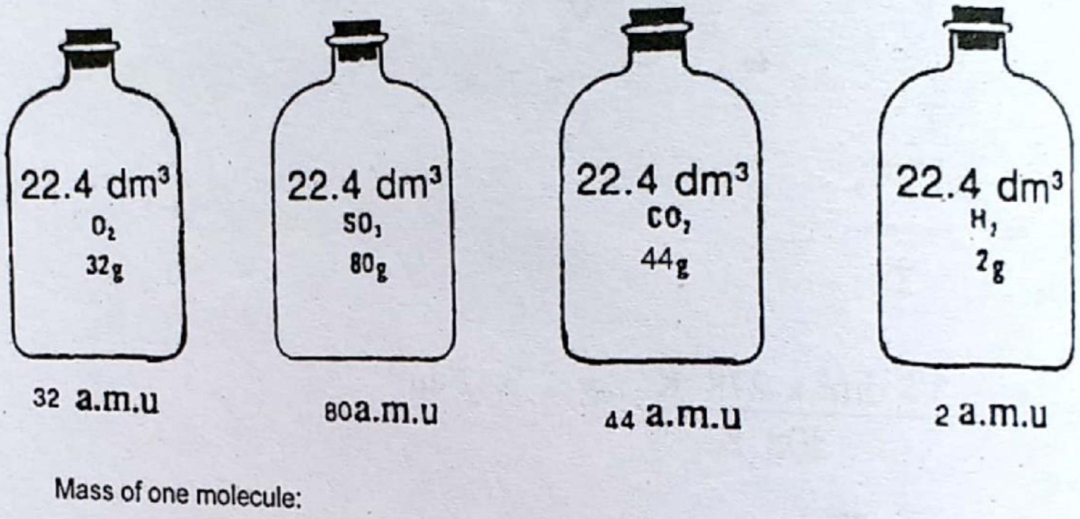


FIGURE 2.8
Measured at STP, 22.4 dm³ of a gas equals a mole.

Avogadro's law enables us to determine the relative molecular mass of gases as illustrated in the following example:

Example :

At standard conditions, 1 dm^3 of Oxygen weighs 1.43 g and 1 dm^3 of Carbon monoxide weighs 1.25 g. According to the Avogadro's law, 1dm^3 of CO at S.T.P. contains the same number of molecules as 1dm^3 of O_2 under similar conditions. Hence a molecule of CO has mass 1.25/1.43 times as much as a molecule of O_2. Accordingly, if we take the molecular mass of O_2 as 32, then the molecular mass of CO would be :

$$\frac{1.25}{1.43} \times 32 \quad = \quad 28 \text{ g/mol}$$

The General Gas equation (Ideal gas equation).

Equation of state: Boyle's law, Charles's law and Avogadro's law may be combined together to give a general relation between the volume, temperature, pressure and number of moles of a gas.

According to Boyle's law: $V \propto$

According to Charles's law $V \propto T$ (absolute) (P is constant)
According to Avogadro's law $V \propto n$ (P and T constant)

Combining these laws we get :

$$V \propto \frac{1}{P} \times T \times n$$

$$V = R \times \frac{1}{P} \times T \times n$$

$\therefore$ $PV = nRT$. This is known as the equation of state because when we specify the four variables – pressure, temperature, volume and number of moles, we define the state for a gas.

Where R is a constant and is called gas constant. The above equation is known as the equation of state or ideal gas equation.

For 1 mole of a gas , $n = 1$ therefore $PV = RT$.

$$\frac{PV}{T} = R$$

If pressure changes to P_1 and temperature to T_1 then volume also changes to V_1

Then $\dfrac{P_1V_1}{T_1} = R$

Similarly $\dfrac{P_2V_2}{T_2} = R$

$$\therefore \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \cdots\cdots\cdots \frac{P_nV_n}{T_n} = R$$

This relationship is used to solve problems regarding changes of volumes of gases, due to the changes in the pressure and temperature.

The value of R (The gas constant):-

(i) When P is expressed in atmospheres, V in litres or dm^3 .

$R = \frac{PV}{nT}$ P = 1 atmosphere
V = 22.4 dm^3 (Molar volume)

$= \frac{1 \text{ mole} \times 22.4 dm^3}{1 \text{ mole} \times 273 \text{ K}}$ Since n = 1
T = 273K

= 0.0821 dm^3 atmosphere K^{-1} $mole^{-1}$

(ii) When P is expressed in newtons per square metre

$N m^{-2}$, Vin cubic metres;

P = 101300 Nm^{-2}
V = 0.0224 m^3 (since 1 $dm^3 = 10^{-3} m^3$)
n = 1
T = 273 K

$$R = \frac{PV}{nT} = \frac{101300 (Nm^{-2}) \times 0.0224 (m^3)}{1 (mole) \times 273 (K)}$$

$= 8.3143 \text{ N m } K^{-1} mol^{-1} = 8.3143 \text{ J } K^{-1} mol^{-1}$
$= 8.3143 \text{ J/(mol K)}$

Problem – 1

What will be the volume occupied by 14 g of nitrogen at 20°C and 740 torr pressure?

Data:

$P = 740 \text{ torr} = \frac{740}{760} = 0.974 \text{ atm.}$

$V = ?$

$n = \frac{14g}{28} = 0.5$ mole

$R = 0.0821\ dm^3 - atm\ K^{-1}\ mole^{-1}$

$T = 20 + 273 = 293\ K.$

$PV = nRT$

$$V = \frac{nRT}{P}$$

$$= \frac{.5 \times .0821 \times 293}{0.974} = 12.345\ dm^3$$

Problem - 2

A certain mass of nitrogen gas at 20°C and at 740 torr pressure occupies 12.345 dm³. Calculate the volume that it will occupy at S.T.P.

Data:
$V_1 = 12.345\ dm^3$
$V_2 = ?$
$P_1 = 740$ torr
$P_2 = 760$ torr
$T_1 = 20 + 273 = 293\ K$
$T_2 = 273\ K$

We know that:

$$\frac{P_2 V_2}{T_2} = \frac{P_1 V_1}{T_1}$$

$$V_2 = \frac{P_1 V_1}{T_1} \times \frac{T_2}{P_2}$$

$$= \frac{740\ torr \times 12.345 dm^3}{293\ K} \times \frac{273\ K}{760\ torr} = 11.12\ dm^3$$

Problem - 3

Calculate the volume that will be occupied by 0.8 moles of oxygen gas taken at 30°C and at 800 torr of pressure?

Data:

$V = ?$

$n = 0.8$ mole

$R = 0.0821\ dm^3\ atm/K\ mole$

$T = 30 + 273 = 303\ K$

$$P = \frac{800\ torr}{760} = 1.053\ \text{Atmosphere}$$

Solution:

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$V = \frac{0.8 \times 0.0821 \times 303}{1.053}$$

$$= 18.899\ dm^3$$

GRAHAM'S LAW OF DIFFUSION:

Diffusion means mixing up of molecules of different gases moving in opposite directions. We know that gas molecules are constantly moving in haphazard direction. Therefore when two gases are placed separated by a porous membrane, they diffuse through the membrane and intermingle with each other. Light gases diffuse more rapidly than the heavier one. Graham (1881) established a quantitative relationship between the rates of diffusion of gases and their densities and is termed as Graham's law of diffusion.

This law states "Rate of diffusion of a gas is inversely proportional to the square root of the density of that gas provided the pressure and temperature are the same for the two gases". This law is stated mathematically.

$$r \propto \frac{1}{\sqrt{d}}$$

where 'r' is the rate of diffusion and 'd' the density of the gas.

If we consider two gases, having densities d_1 and d_2 respectively ; then;

$$r_1 \propto \frac{1}{\sqrt{d_1}}$$

$$\therefore \quad r_1 = \frac{K}{\sqrt{d_1}}$$

Similarly :

$$r_2 \propto \frac{1}{\sqrt{d_2}}$$

$$r_2 = \frac{K}{\sqrt{d_2}}$$

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$$

Since the density of a gas is proportional to its molecular mass, we can write the above equation as : $\left(d_1 = \frac{M_1}{V} \text{ and } d_2 = \frac{M_2}{V}\right)$

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Problem - 1

The ratio of the rates of diffusion of two gases A and B is 1.5 : 1. If the relative molecular mass of gas A is 16, find out the relative molecular mass of gas B?

Solution:

According to Graham's law :

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$$

$$\frac{1.5}{1} = \sqrt{\frac{M_B}{16}}$$

By taking square of both sides

$$\left(\frac{1.5}{1}\right)^2 = \left(\sqrt{\frac{M_B}{16}}\right)^2$$

$$
\begin{aligned}
\therefore M_B &= (1.5)^2 \times 16 \\
&= 2.25 \times 16 \\
&= 36
\end{aligned}
$$

Problem -2

Compare the rates of Helium and sulphur dioxide.

Solution :

The molecular masses of Helium and sulphur dioxide are 4 and 64 respectively. Hence,

$$\frac{r_{He}}{r_{SO_2}} = \sqrt{\frac{M_{SO_2}}{M_{He}}} = \sqrt{\frac{64}{4}} = \frac{8}{2} = 4$$

Helium diffuses four times as fast as SO_2.

DALTON'S LAW OF PARTIAL PRESSURES:

When two or more gases which do not react chemically, are mixed in the same container, each gas will exert the same pressure as it would exert if it alone occupied the volume containing the mixed gases. This portion of the total pressure of a mixture is known as the *partial pressure*. Based on this behaviour of gases, John Dalton formulated a law in 1801 which is known as the *Dalton's law of partial pressures*. It is stated as :

"The total pressure of a mixture of gases is the sum of the partial pressures of the gases in the mixture". This law is stated mathematically.

$$P = p_1 + p_2 + p_3 \ldots\ldots$$

P is the total pressure; the small p's refer to the part of the pressure (partial pressure) exerted by each gas in the mixture. If the gas 'A' exerts a pressure of 0.3 atm., and gas 'B' 0.2 atm., the total pressure of the mixture would be 0.5 atm (fig 2.9). Partial pressure of a gas is calculated by using the equation:

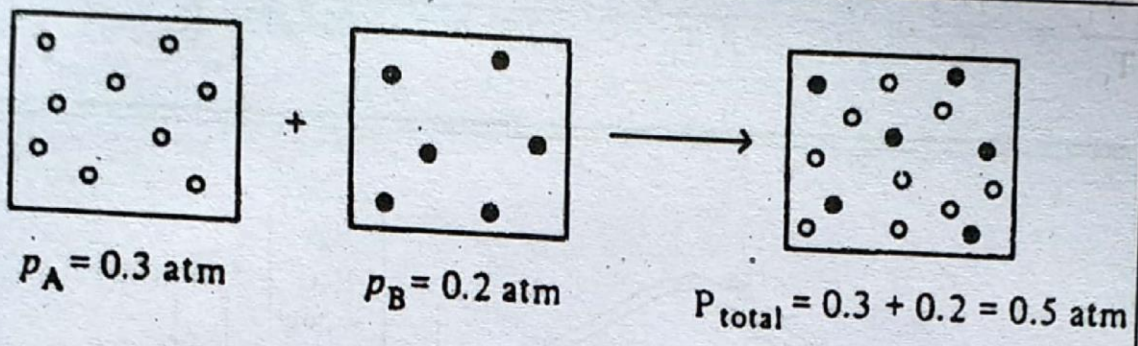


FIGURE 2.9

$$\frac{\text{Partial pressure of gas}}{\text{Total pressure}} = \frac{\text{moles of the gas}}{\text{Total moles}}$$

$$p_{gas} = \frac{P \times n_{gas}}{n_{(total)}}$$

In terms of *kinetic theory*, Dalton's law can be explained as follows: In a non-reactive mixture, each gas exerts a separate pressure on the container because of collisions of its molecules with the walls of container. Thus the total pressure in the container is caused by the sum of all the collisions.

Application: This law applies most commonly to the case of a gas collected over water (See fig. 2.10). For example, suppose 100 cm^3 of an insoluble gas are collected over water at a pressure of 745 torr and at 15°C. If it is saturated with water vapour, the pressure of the dry gas is (745-13) torr, since the vapour pressure of water at 15°C is 13 torr. Thus

The pressure of moist gas = Pressure of dry gas + pressure of water vapour
The pressure of dry gas = Pressure of moist gas - pressure of water vapour
745 torr – 13 torr
∴ P dry gas = 732 torr

Now using gas equation, we can calculate the volume of the dry gas at S.T.P.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \text{ or } V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{732 \text{ torr} \times 100 \text{cm}^3 \times 273 \text{ K}}{760 \text{ torr} \times (15+273) \text{ K}} = 91.3 \text{cm}^3$$

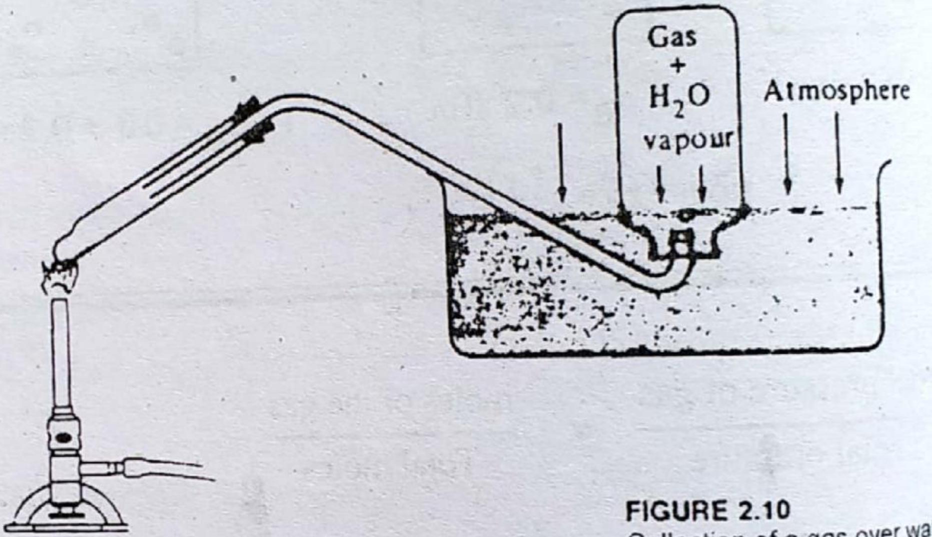


FIGURE 2.10
Collection of a gas over water.

Problem - 1

In the preparation of oxygen in the laboratory, 500 cm^3 of the moist ga was collected over water at 25°C and 724 torr. What volume of dry oxygen a S.T.P. was produced? (Pressure of water vapours at 25°C = 24 torr).

Solution:

Pressure of dry O_2 is calculated by using Dalton's law:

$$p_{\text{dry } O_2} = p_{\text{moist gas}} - p_{H_2O}$$

$$= 724 \text{ torr} - 24 \text{ torr}$$

$$= 700 \text{ torr}$$

Now calculate the volume of dry gas at S.T.P.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \text{(Gas equation)}$$

$$V_{\text{dry gas}} = \frac{P_1 V_1 T_2}{P_2 T_1}$$

$$= \frac{700 \text{ torr} \times 500 \text{ cm}^3 \times 273 \text{ K}}{760 \text{ torr} \times 298 \text{ K}} = 421.9 \text{ cm}^3$$

Problem - 2

A mixture of gases at 760 torr contains 2 moles nitrogen and 4 moles carbon dioxide. What is the partial pressure of each gas in torr?

Solution:

Total moles (n) = 6 moles; total pressure = 760 torr

$$P_{\text{gas}} = \frac{P_{(\text{total})} \times n_{(\text{gas})}}{n_{(\text{total})}}$$

$$P_{N_2} = \frac{760 \text{ torr} \times 2 \text{ mol}}{6 \text{ mol}} = 253.33 \text{ torr}$$

$$P_{CO_2} = \frac{760 \text{ torr} \times 4 \text{ mol}}{6 \text{ mol}} = 506.67 \text{ torr}$$

Check: Total pressure = Sum of partial pressures = 760 torr

2.4 IDEAL GASES

An ideal gas is one whose behaviour can be predicted precisely on the basis of kinetic molecular theory and gas laws. Two significant properties of an ideal gas are that the molecules do not attract each other and they occupy negligible space. Actually, no such gas exists, although gases such as hydrogen and oxygen do not deviate greatly from the ideal behaviour at moderate temperatures and pressures. The main cause of deviation of the real gases from the ideality, is the

At nominal pressure, however the brake liquids can transmit the applied force to perform certain machenical function such as braking the motions in a motor car.

In terms of molecular model, the compressibility of liquids can be explained as follows: The liquid molecules due to their closeness roll over one another. Owing to very little space the liquid molecules cannot be pushed close by pressure. Very high pressures are required for squeezing a liquid.

3. EXPANSION AND CONTRACTION: A liquid normally expands on heating and contracts on cooling (as does mercury or alcohol in thermometers). The molecular model explains this behaviour as follows: On heating, the kinetic energy and so the velocity of molecules increases. As a result, the volume taken by each molecule increases on heating. It results in the expansion. Contraction may similarly be explained.

The distance between liquid molecules are so small that properties of liquids are largely determined by the intermolecular attractions. Viscosity, surface tension and vapour pressure, for intrance, are not only dependent on the intermolecular forces but they are also modified with the temperature changes.

2.7 - VISCOSITY

Like gases, liquids have the ability to flow (fluids) but unlike gases they have definite surface. Different liquids have different rates of flow. Some liquids like honey, mobil oil etc. flow slowly and are called viscous liquids, while gasoline, ether etc. which flow quickly are called less viscous. The intermolecular attractions account for this behaivour of liquids.

Inter molecular attractions between the molecules within the bulk of liquid occur in all directions (Fig. 2.11) i.e. the molecules are attracted mutually in all

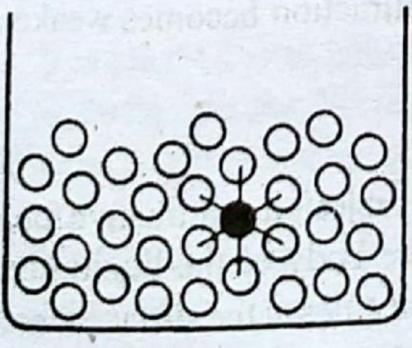


FIGURE 2.11
Intermolecular attractions between molecules in the bulk of a liquid occur in all directions.

directions. These intermolecular attractions which draw the liquid molecules together are responsible for the *viscosity*, that is, *'its resistance to flow'*. The liquids that have stronger intermolecular attraction have greater viscosities, their molecules are less easily moved about with respect to each other. Ethyl alcohol (C_2H_5OH), for instance, is more viscous than ether ($C_2H_5.O.C_2H_5$) because there are hydrogen bonds among alcohol molecules. The grading of motor oils is done on the basis of viscosity for example, sea grade 30 oil is less viscous than the grade 40 oil.

The viscosity is expressed in poises (1 poise = 1 gm/cm.S) or its multiples - centipoises and millipoises. The S.I. unit is newtonx seconds per square metre $N.s.m^{-2}$

1 centipoise = 10^{-3} $N.s.m^{-2}$

Viscosity of certain liquids at different temperatures is given in the table 2.1.

Table 2.1 – Viscosity (centipoise) at different temperatures.

LIQUID	VISCOSITY		
	0°C	20°C	50°C
Water	1.789	1.005	0.55
Ethyl alcohol	1.78	1.19	0.701
Glycerine	12100	1499	—
Acetone	0.395	0.322	0.246

From the above table, you must have noticed that the viscosity of a liquid decreases with the rise in temperature. In fact, with the rise of temperature, the volume of liquid increases and the intermolecular attraction becomes weaker.

2.8 - SURFACE TENSION (γ)

As pointed out earlier, all molecules in a liquid are subjected to attractive forces from neighbouring molecules. For molecules in the body of the liquid, these forces balance each other (See fig. 2.11), but for molecules at the surface there is

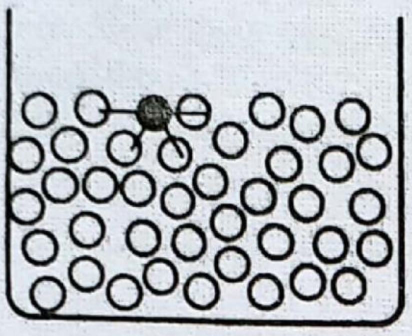


FIGURE 2.12
Intermolecular attractions at a liquid surface occur only within the surface and below.

resultant attractive force acting downwards. (Fig. 2.12). These forces which tend to pull the surface molecules downwards are responsible for the surface tension. It is designated by γ. The dimensions of surface tension are dynes/cm or ergs/cm^2 (numerically the values in both the units are equal). *Hence the surface tension may be regarded as the force per unit length or energy per unit area of the surface of liquid.* Table 2.2 gives the surface tension of some liquids at 20°C.

Table 2.2 — Surface tension (dynes/cm or erg/cm^2) at 20°C.

Liquid	γ	Liquid	γ
Water	72.583	Acetone	23.7
Ethyl alcohol	22.03	Benzene	28.88
Chloroform	27.1	Mercury	471.6

This table shows that the surface tension of water is higher than most of the organic liquids because there are stronger intermolecular attraction in water (Hydrogen bonds).

It is interesting to note that the surface tension tends to reduce the surface area of the liquids. For this reason, a freely falling drop of liquid is spherical (Fig. 2.13), since sphere has the smallest ratio of surface area to volume.

It is our common observation that a liquid always rises in a capillary tube (fig. 2.14), the effect is referred to as the *'capillary action'*. Capillary

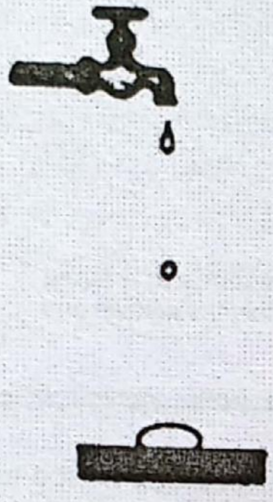


FIGURE 2.13
A free-falling drop of water is nearly spherical. A droplet of water on a waxed surface forms a "bead" as it attempts to become spherical.

action of liquids is also due to their surface tension. The surface area to volume ratio of a capillary of small bore is much smaller than the container in which liquid is placed.

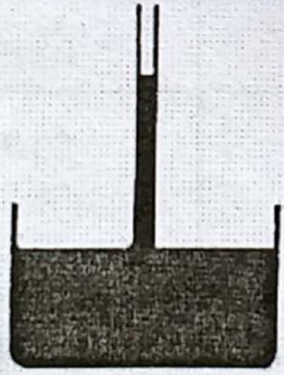


FIGURE 2.14
Capillary action

Like viscosity, the surface tension of a liquid also decreases with the rise of temperature owing to the weakening of the inter molecular attractions.

2.9 - VAPOUR PRESSURE

Water slowly evaporates at room temperature. Gasoline evaporates quite readily without boiling. A liquid that evaporates readily is said to be *volatile*.

If a quantity of a liquid is contained in a tightly covered container at a certain temperature, molecules will evaporate from the liquid surface into the enclosed space. Since these molecules move about above the liquid and can not escape, some of them will collide with the liquid surface and re-join it (condensation) Fig. 2.15). As a result of the two opposite processes-evaporation and condensation, a state of equilibrium is finally established. This equilibrium is marked by the attainment of a *vapour pressure* which is defined as *the pressure exerted by the vapours when they are in equilibrium with the liquid phase*. It is constant at constant temperature. The vapour pressure is independent of the amount of liquid present. For example the vapour pressure of water at 20°C is 17.5 torr.

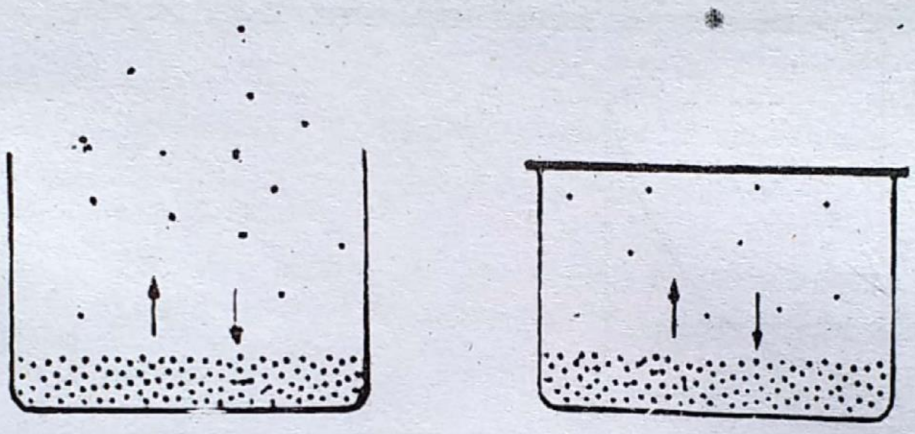


Figure 2.15. In an open container a liquid will evaporate away but in a closed container the rate of condensation becomes equal to the rate of evaporation, and the liquid-vapour system is in equilibrium.

The evaporation and the vapour pressure can be explained in terms of kinetic theory. Evaporation occurs when high energy of molecules at the surface of liquid break away from their neighbours and escape into the gas phase. During evaporation, the escape of high energy molecules lowers the average kinetic energy of the remaining molecules, consequently the temperature of the remaining liquid falls down. Thus it is said that evaporation is a cooling process.

2.10 - CHANGE OF STATE – BOILING

When a liquid is heated, the more molecules acquire increased kinetic energy and escape into the air and the evaporation becomes more rapid. As a consequence, the vapour pressure of the liquid increases. For example at 40°C the vapour pressure of water is 55.3 torr and at 95°C it is 634 torr. It is illustrated by the vapour pressure-temperature curve in fig. 2.16. When the vapour pressure of liquid equals the atmosphere pressure, the bubbles of vapour form easily and rise to the surface. The liquid is said to *boil*. Thus the *boiling point* of a liquid is the temperature at which the vapour pressure of the liquid equals the outside pressure. When the outside pressure is the atmospheric pressure (760 torr), the boiling point is referred to as the *normal boiling point*.

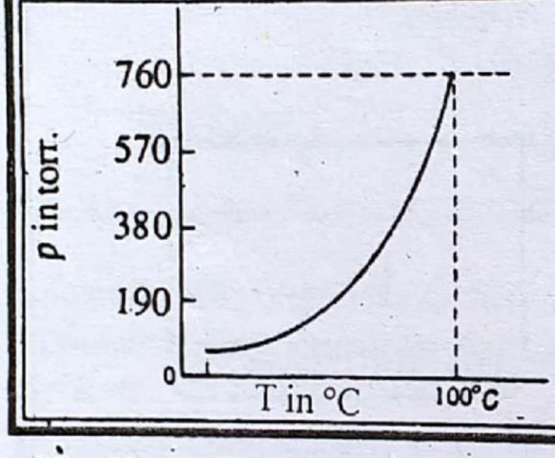


FIGURE 2.16
The vapour pressure-temperature curve for water.

The boiling point of liquids varies with the outside pressure. In mountainous regions, the boiling point of water will be below 100°C because the atmospheric pressure is below 760 torr. Thus in such regions of high altitude, foods take longer time to cook because of the low boiling temperatures. In the pressure cookers, on the contrary, food cook more rapidly because of the higher boiling temperatures.

It may be pointed out that certain liquids tend to decompose at their boiling point temperatures. Such liquids are distilled at low temperatures under reduced *pressure*. By lowering the pressure, boiling point of the liquid is reduced and their decomposition is prevented. For example, glycerine boils at 290°C at 760 torr but it decomposes at this temperature. If it is heated under reduced pressure (50 torr), it boils at 210°C and can easily be distilled without decomposition.

2.11 SOLID STATE

A solid state is that state of matter in which atoms, molecules or ions are held together by strong attractive forces. Due to strong attractive interatomic, molecular or ionic forces, the particles constituting the solid can not move, however, they possess some vibrational energy at even near absolute zero.

Progress in science and technology leads to advancement in knowledge about solid state also. Hence we must endeavour to understand the nature and behaviour of solids.

1– Behaviour of Solids

Compared with the other two states of matter i.e. gases and liquids, the solids have highest degree of order. This orderlines accounts for many behaviour of solids. Some of these are:

(i) Compressibility:– The compressibility of solids is nearly zero. In terms of kinetic theory, the particles of solids are so tightly bound together that only slight unfilled space is left, hence density of solids is much higher than gases and liquids, so they can withstand considerable external stress.

(ii) Deformity:– Solids are deformed or shattered by high pressure. This is because when some particles are dislocated the force of attraction is so strong that the rearranged atoms are held equally well to their new neighbours as shown in Fig.2.17.

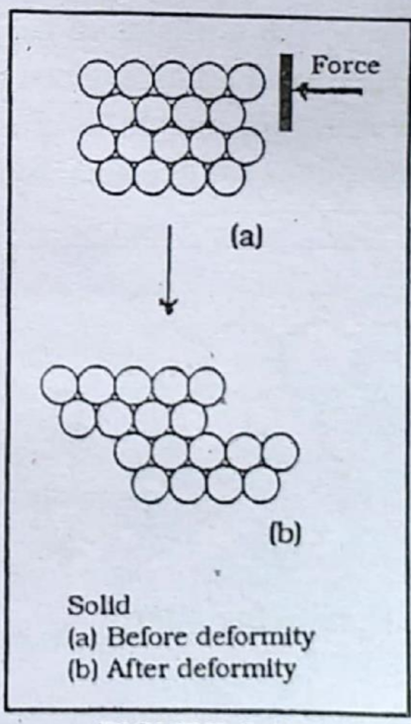


FIGURE:2.17

(iii) Diffusibility:– In terms of kinetic theory, there is no translational movement of molecules in solid, but nevertheless particles are vibrating at their mean positions. This vibrational motion is responsible for diffusion in solids.

(iv) Melting:– Solids on heating melt at a particular temperature.

In terms of kinetic theory when solids are heated vibrational energy of their particles increases; until at melting point some particles are vibrating with sufficient energy to overcome the forces holding them, hence they become mobile i.e. solids fuse.

(v) Sublimation:– There are some solids substances, as camphor, iodine, solid CO_2, naphthalene etc. which change directly to vapour on heating without passing through liquid phase. This phenomenon is called sublimation. In terms of kinetic theory, the inter–molecular force in such solids is less than ordinary solids, hence high–energy molecules at solid surface overcome the attractive forces and directly pass into vapours.

2– Classification of solids:– Solids have been classified mainly as:

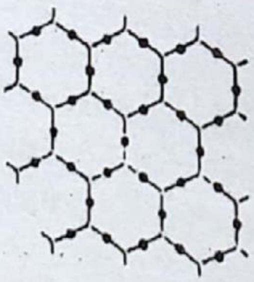
(i) Crystalline (ii) Amorphous

(i) Crystalline solids:– Crystalline solids have characteristic geometrical shape due to highly ordered three dimensional arrangement of particles. They are bounded by plane surfaces called faces; which intersect each other at particular angles.

The reason for the formation of regular pattern or crystal is that atoms or ions or molecules tend to arrange themselves in positions of maximum attraction or in positions of minimum energy. Examples are: sugar, alum, metals, diamonds etc.

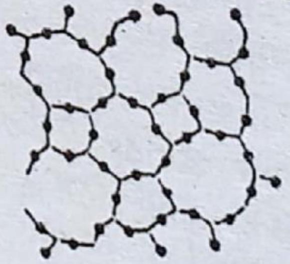
(ii) Amorphous solids:– These are the solids which do not have a definite shape. The word amorphous means "without form".

The particles of such solids have a random i.e. non–repetitive three dimensional arrangements. If a substance in liquid state is cooled rapidly, the particles are unable to arrange themselves in an orderly fashion, hence an amorphous solid results. Examples are: glass, plastic, rubber etc.



Crystalline quartz

FIGURE. 2.18

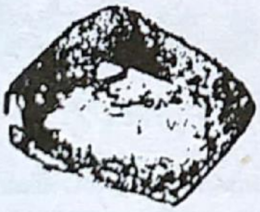


Amorphous quartz
It is formed when crystalline quartz is melted and cooled rapidly.

FIGURE. 2.19

The difference between crystalline and amorphous solids is summarized below:

Crystalline Solids	Amorphous Solids
(i) Geometry	
Particles of crystalline solids are arranged in an orderly three dimensional network called crystal lattice, hence they have definite shape.	Particles of amorphous solids are not arranged in a definite pattern, hence they do not have a definite shape.
(ii) Melting point	
Crystalline solids have sharp melting point, this is because attractive forces between particles are long range and uniform. These forces breakdown at the same instant, at melting point.	Amorphous solids melt over a wide range of temperature i.e. they do not have sharp melting point, because the inter-molecular forces vary from place to place.
(iii) Cleavage and cleavage plane	
The breakage of a big crystal into smaller crystals of identical shape is called cleavage. Crystals cleavage along particular direction, the plane which contains the direction of cleavage is called cleavage plane.	Amorphous solids do not break down at fixed cleavage planes.
(iv) Anisotropy and isotropy	
Physical properties of crystals such as electrical conductivity, refrac-	Amorphous solids are isotropic i.e. their physical properties are



Crystal before cleavage

FIGURE.2.20

Crystal after cleavage

Note: Smaller parts are of same shape as big crystal.

FIGURE.2.21

tive index, etc. are different in different direction. This property is called anisotropy. For example graphite can conduct electricity parallel to its plane of layers but not perpendicular to plane.	same in all directions.

(v) Symmetry

When crystalline solids are rotated about an axis, their appearance does not change i.e. they possess symmetry.	Amorphous solids are not symmetrical.

2.12 TYPES OF CRYSTALS

Crystals are classified in the following four types:

(i) Atomic crystals (ii) Ionic crystals (iii) Covalent crystals
(iv) Molecular crystals

This classification is based on the nature of forces or bonding between the atoms, ions or molecules constituting the crystal.

(i) Atomic crystals:– Examples of such crystals are metals, hence they are also called metallic crystals. They consist of atoms packed and held together by metallic bond.

In metals, the valence electrons move freely in crystal lattice and are said to form an electron gas, in which positive ions are immersed. Such crystals are characterised by: (a) Lustre (b) High melting point (c) Electrical and thermal conductivity (d) Malleability (drawing into sheets) (e) Ductility (drawn into wires).

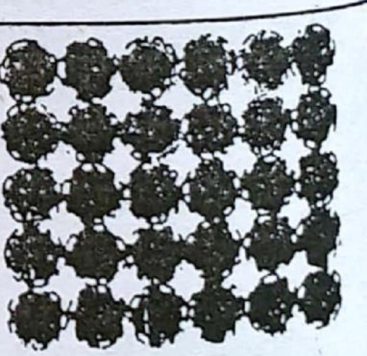
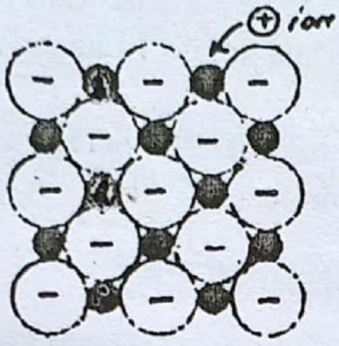


FIGURE:2.22

Two dimensional view of atomic crystals. Atoms vibrate at their mean position.

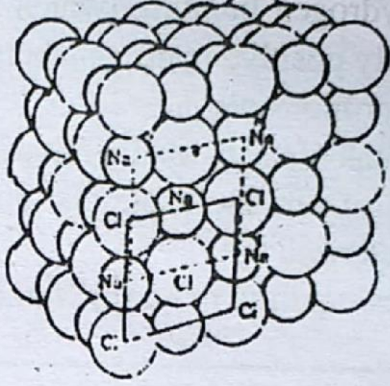
(ii) Ionic crystals:– Such crystals consist of positively and negatively charg ions, held together by electrostatic forces of attraction. Individual ions have identity. They are characterised by:

(a) High melting point (b) Conduction of electricity in fused state; as well as solution form. (c) Brittleness and hardness (d) Indefinite growth of crystal, u it is broken or is stopped from growing.



Two dimensional view of ions in ionic crystal.

FIGURE.2.23



Three dimensional view of ions packing in Na Cl

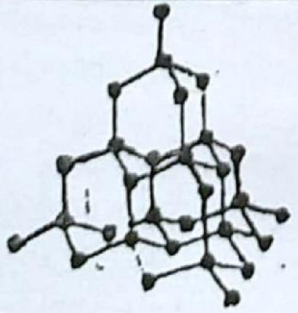
FIGURE.2.24

(iii) Covalent crystals:– Such crystals consist of atoms held together by covalent bonds. Covalent bonds are very strong. They require large amount of energy to break them. Examples of covalent crystals are: graphite, diamond, silica and carborundum.

In diamond each carbon atom is joined to four other carbon atoms at an angle of 109°, called tetrahedral angle. Due to close packing of atoms and large number of covalent bonds, diamond is very hard and has high melting point.

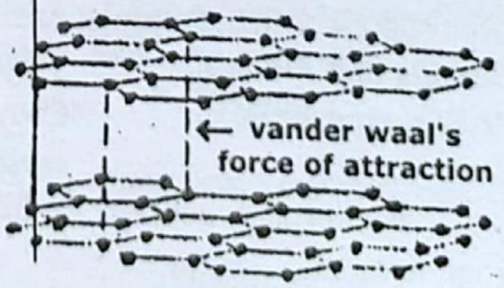
In graphite each carbon atom is joined to three other carbon atoms at an angle of 120°, forming layers of hexagons. Adjacent layers are held together by weak Vander waal's forces. Due to space between layers graphite is soft and as layers easily slide over one another so graphite has greasy texture. Graphite conducts electricity in the plane of layers, as each carbon has one free electron. Covalent crystals are characterised by:

(a) High melting point (b) Low density (c) High refractive index



Space lattice of diamond

FIGURE.2.25



Space lattice of graphite

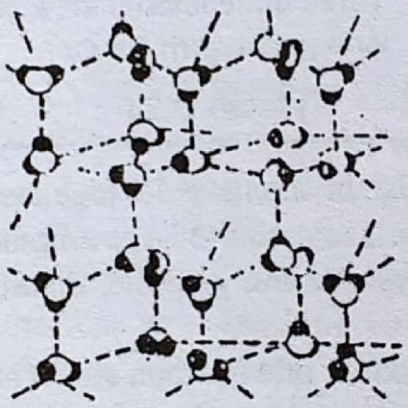
FIGURE.2.26

(iv) Molecular crystals:– Such crystals are composed of molecules, such as ice, I_2, CO_2. The molecules are held together either by the following forces:

(a) Hydrogen bonding (which is an electrostatic force of attraction between partially positive hydrogen atom of one molecule and an electronegative atom of other molecules).

(b) Weak Vander waal's forces (which arise due to attraction between atomic nuclei of one molecule and electrons of other molecule). Such crystals are characterised by:

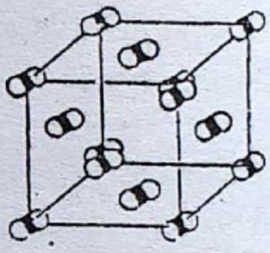
(a) Low melting point (b) Non-conduction of heat and electricity.



Structure of ice

Each water molecule is attached to four others through hydrogen bonding.

FIGURE:2.27



Unit cell of solid CO_2

FIGURE:2.28

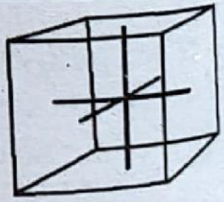
2.13 ISOMORPHISM

When two substances have same crystal structure, they are said to be isomorphous and phenomenon is called isomorphism. For example NaF and MgO both form cubic crystals hence are isomorphous. Other examples are:

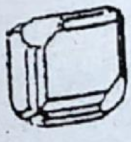
$CaCO_3$ and $NaNO_3$	Trigonal crystals
$ZnSO_4$ and $NiSO_4$	Orthorhombic crystals

Isomorphic substances have following properties:

(i) They have different physical and chemical properties.



A cube



A cube crystal

FIGURE:2.29

(ii) They have same empirical formula.

Pairs	Ratio in atoms
NaF and MgO	1:1
$NaNO_3$ and $CaCO_3$	1:1:3
$ZnSO_4$ and $NiSO_4$	1:1:4

(iii) When their solutions are mixed, they form mixed crystals.

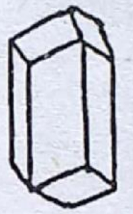
(iv) They show property of over growth; i.e. if crystal of one substance is suspended in saturated solution of other substance, the crystals of latter begin to grow over the former in similar pattern.

2.14 POLYMORPHISM

A substance under different conditions can form more than one type of crystal.

The substance which can exist in more than one crystalline form is called Polymorphous and the phenomenon is known as Polymorphism. For example $CaCO_3$ exists in nature in two crystalline form (i) Calcite, which is trigonal (ii) Aragonite, which is orthorhombic

FIGURE:2.30



Trigonal crystal

Orthorhombic crystal

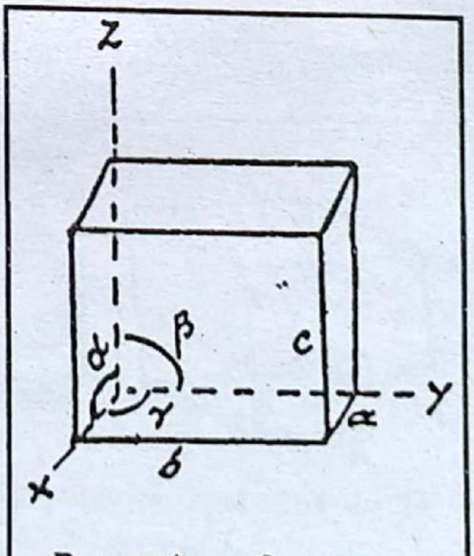
FIGURE:2.30

UNIT CELL: A crystal is formed by a large number of repetition of basic pattern of atoms, ions or molecules in space.

The basic structural unit which when repeated in three dimensions generates the crystal structure is called a unit cell.

A unit cell has a definite geometric shape and hence is identified on the basis of (i) Length of its edges, denoted by letters, a , b and c. (ii) Angle between the edges, represented by letters, α, β and γ.

The length and angles of a unit cell are collectively known as Cell Dimensions or Cell Parameters.

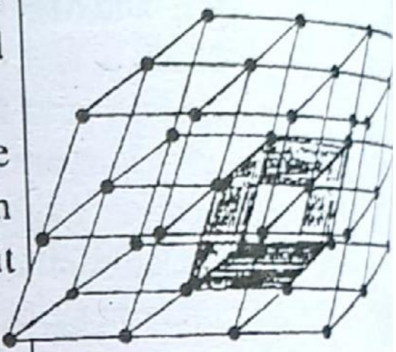


Parameters of a unit cell.

FIGURE:2.31

Space or crystal lattice:– If atoms, ions or molecules constituting a crystal are replaced by points and placed at the same places as in a unit cell, then the three dimensional array of points is called crystal lattice or space lattice.

All points in a crystal lattice have same environment and can not be distinguished from each other. Three dimensional lattice with a unit cell is shown in the diagram.



Crystal lattice
with a unit cell
FIGURE:2.32

The crystal systems:–Based on the difference in shapes and dimensions of unit cells, all the known crystals have been classified into seven groups called crystal systems. The seven crystal systems are as follows:

(i) Cubic system:– In cubic system, all axes are of equal length and all angles are of 90° i.e.

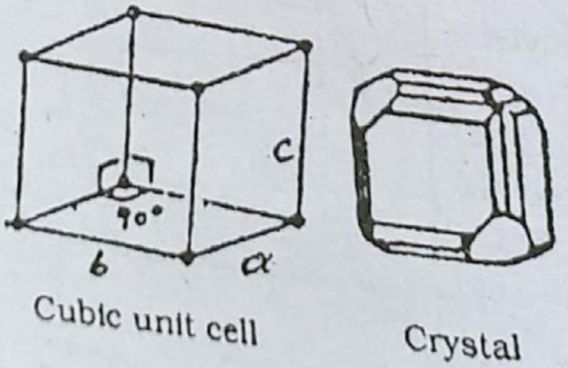
$$a = b = c$$
$$\alpha = \beta = \gamma = 90°$$

Examples: NaCl, ZnS and diamond.

(ii) Tetragonal system:– It has two axes of equal length, the third is differen
All angles are of 90° i.e.

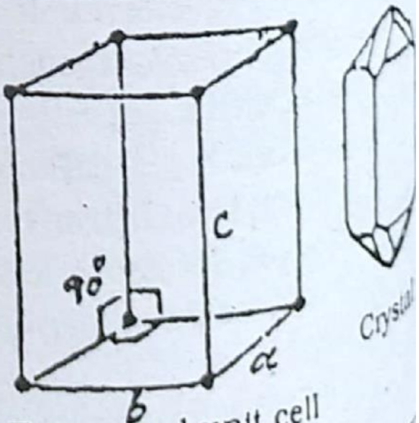
$$a = b \neq c$$
$$\alpha = \beta = \gamma = 90°$$

Examples: SnO_2(white), $BaSO_4 \cdot 4H_2O$



Cubic unit cell Crystal

FIGURE:2.33

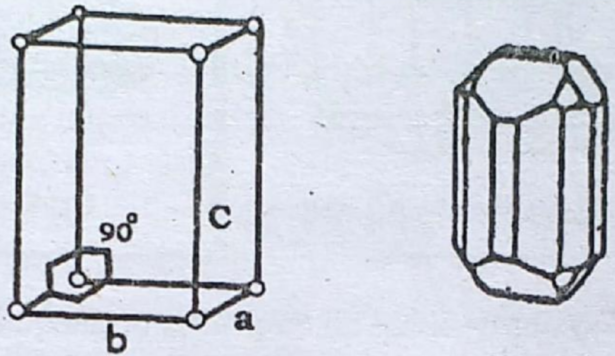


Tetragonal unit cell
FIGURE:2.34

(iii) Orthorhombic system:– In this system, all three axes are of dif length; but all angles are of 90° i.e.

$$a \neq b \neq c$$
$$\alpha = \beta = \gamma = 90^\circ$$

Examples are: $FeSO_4 \ 7H_2O$, $ZnSO_4 \ 7H_2O$, KNO_3



Orthorhombic unit cell

Crystal

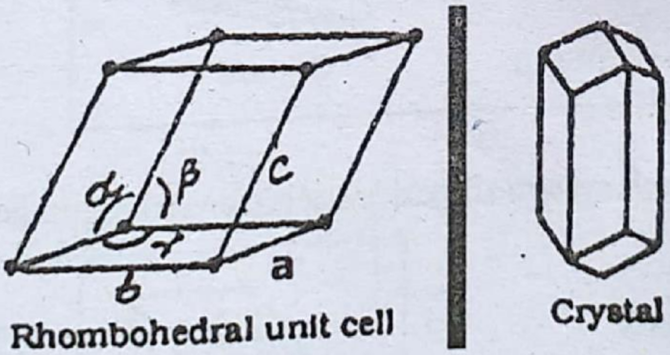
FIGURE:2.35

(iv) Trigonal or rhombohedral system:– It has three axes of equal length. All angles are equal but more than 90° and less than 120° i.e.

$$a = b = c$$
$$\alpha = \beta = \gamma \neq 90^\circ$$

(Angles >90°<120°)

Examples: Calcite, KNO_3, $AgNO_3$, and ice.



Rhombohedral unit cell

Crystal

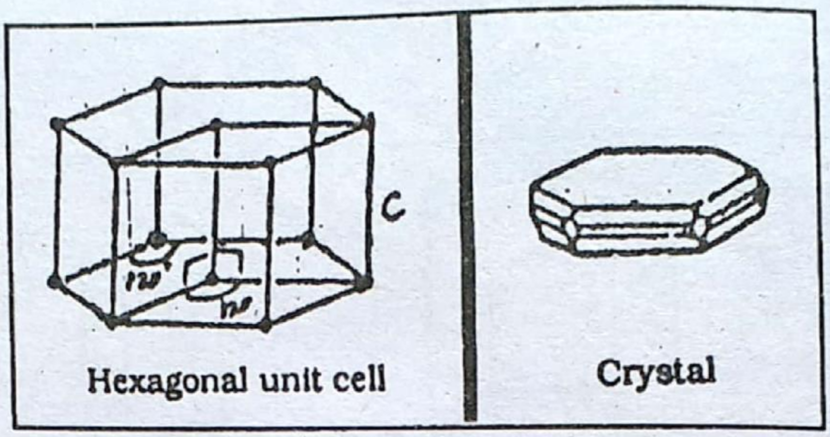
FIGURE:2.36

(v) Hexagonal system:– In this system, two axes are of equal length, but the third is different. Two angles are of 90° and third is of 120° i.e.

$$a = b \neq c$$
$$\alpha = \beta = 90^\circ$$
$$\gamma = 120^\circ$$

Examples: Graphite, snowflakes and quartz (SiO_2)

FIGURE:2.37



Hexagonal unit cell

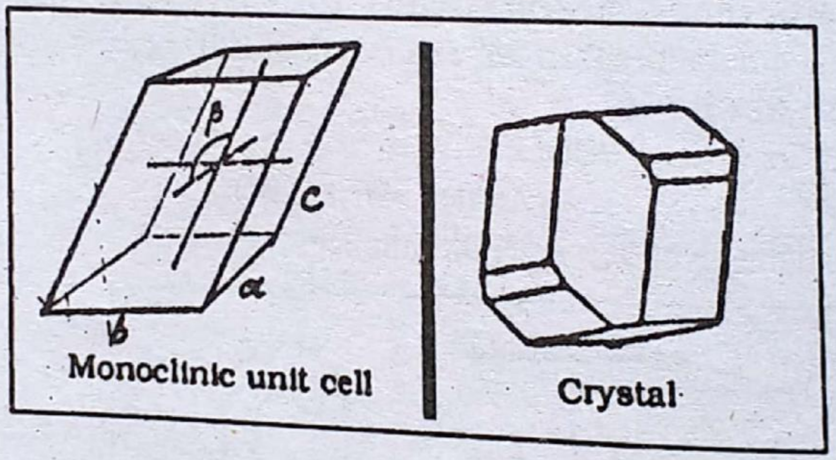
Crystal

(vi) **Mono clinic system:–** It has all three axes of unequal lengths. One of the axis is at right angle to the other two i.e.

$$a \neq b \neq c$$
$$\alpha = \gamma = 90^\circ$$
$$\beta \neq 90^\circ$$

Examples: $CuSO_4$ $5H_2O$, Na_2CO_3 $10H_2O$, cane sugar.

FIGURE:2.38



Monoclinic unit cell

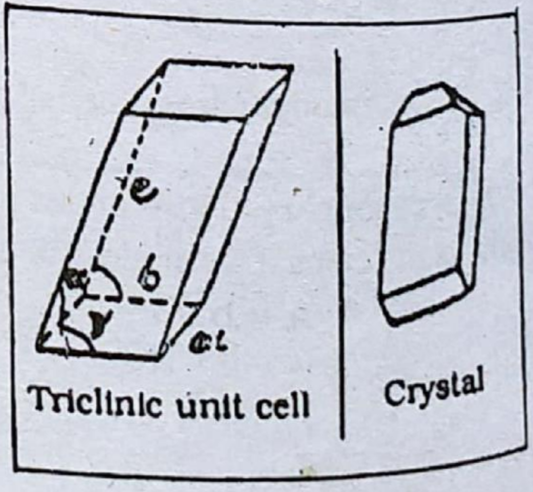
Crystal

(vii) **Triclinic system:–** In this system all axes are of different lengths and angles are also different.

$$a \neq b \neq c$$
$$\alpha \neq \beta \neq \gamma \neq 90^\circ$$

Examples: $CuSO_4 \cdot 5H_2O$, $K_2Cr_2O_7$

FIGURE:2.39



Triclinic unit cell

Crystal

2.15 CHANGE OF STATE (MELTING OR FUSION)

(i) The melting point:– When a pure crystalline solid is heated, a temperature is reached where it changes sharply into liquid. This is melting point of solid. Melting point is defined as that temperature at which there is equilibrium between solid and liquid phases.

Melting point is affected by: **(a) Impurity:–** Pure substances have sharp melting point. Moreover impurities lower the melting point. Hence melting point is used as criteria of purity of solids.

(b) Pressure:– Those substances which expand on melting, have a rise in melting point when pressure is increased.

Melting point of ice decreases on applying pressure because ice is the solid whose volume decreases on melting.

(ii) The latent heat of fusion:– It is defined as the heat energy required to change 1 g of a solid into liquid at its melting point.

Even though heat is being supplied during melting but temperature remains constant because all of the energy is used up by the molecules to over come the intermolecular forces.

PROGRESS TEST 2

1. In terms of the atoms, ions, and molecules that comprise substances, why are some materials solids, some liquids and still others gases at 25°C.?

2. The process of diffusion occurs most rapidly in gases, less rapidly in liquids, and very slowly in solids. Why?

3. Why does water spilled on a floor evaporate more rapidly than the same amount of water in a glass?

4. Would a pressure cooker be of any value on the Mount Everest?

5. No liquid ionic compounds are known, but many of the known covalent compounds are liquids and some are gases. Account for these differences.

6. What is kinetic theory and how does it account for the following properties of gases, liquids and solids?
   (i) Diffusibility (ii) Compression (iii) Expansion

7. 1.40 dm^3 volume of a gas measured at a temperature of 27°C and a pressure of 900 torr was found to have mass 2.273 g. Calculate the molecular mass of the gas [Hint: Calculate "n" by the formula PV=nRT and then molecular mass] (33.72 amu) Ans.

8. A 12.5 dm^3 vessel contains 4.0 g CH_4, 1.8 g N_2 and 10.0g Xe. What is the pressure in the vessel at 0°C?(0.698 atm) Ans. (Hint: Find out the total moles of the gases and then apply PV = nRT).

9. (a) State and explain the Graham's law of Diffusion.
   (b) Compare the rates of diffusion of the following pairs of gases:
   (i) H_2 and D_2 (ii) CH_4 and He (iii) SF_6 and SO_2 [(i) 1.41/1, (ii) 2/1, (iii) 1.5/1. Ans.

10. State the following gas laws and explain in terms of Kinetic theory.
(a) Boyle's law (b) Charles's law (c) Dalton's law of partial pressure.

11. Define the following :(i) Melting point (ii) Boiling point (iii) Diffusion (iv) Latent heat of fusion.

12. Four containers of equal volume are filled as follows:
(i) 2.0g H_2 at 0°C (ii) 1.0g H_2 at 273°C (iii) 24 g O_2 at 0°C
(iv) 16 g CH_4 at 273°C.
Which container (a) is at the greatest pressure (b) is at the lowest pressure.
[a = (iv) b = (iii)] Ans.

13. Explain the relationship between intermolecular attractions and the kinetic energy of molecules in determining the physical state of a substance.

14. Describe gases, liquids and solids on a molecular basis and explain their behaviour and change of state.

15. Co-relate the energy changes accompanying changes of state.

16. Explain in terms of kinetic molecular model:
(a) Why liquids and solids cannot be compressed as gases can?
(b) Why solids do not flow as liquids and gases do?
(c) Why food is cooked more quickly in a pressure cooker than in a covered pot?

17. Explain (i) Viscosity (ii) Surface tension (iii) Vapour pressure.

18. What are crystalline and amorphous solids ? Explain different crystal systems.

19. State and explain Avogadro's law. How it helps in determining relative molecular mass?

20. What are solids? Describe Atomic, Ionic and Molecular solids.

21. Explain the following:
    (i) A falling drop of a liquid is spherical.
    (ii) A drop of ink spreads on a blotting paper.
    (iii) Evaporation is a cooling process.
    (iv) Boiling point of a liquid remains constant although heat is continuously supplied to the liquid.
    (v) Honey is more viscous than water.
    (vi) Mercury has its meniscus upward.

22. 40 dm^3 of hydrogen gas was collected over water at 831 torr Hg pressure at 23°C. What would be the volume of dry hydrogen gas at standard conditions ? The vapour pressure of water at 23°C is 21 torr of Hg.

(Ans :– 39.23 dm^3).

23. What is the density of methane gas (CH_4) at 127°C and 3.50 atmosphere.

(Ans:– density = 1.70 g/dm^3).