STATE OF MATTER I: GAS

Chapter 4

Teaching Periods	Assessment	Weightage
10	1	11

Students will be able to:

- **List** the postulates of Kinetic Molecular Theory.
- **Describe** the motion of particles of a gas according to kinetic theory.
- **State** the values of standard temperature and pressure (STP).
- **Relate** temperature to the average kinetic energy of the particles in a substance.
- **Use** kinetic theory to explain gas pressure.
- **Describe** the effect of change in atmospheric pressure on the weather.
- **Describe** the significance of absolute zero, giving its value in degree Celsius and Kelvin.
- **State** and explain the significance of Avogadro's Law.
- **Derive** Ideal Gas Equation using Boyle's, Charles' and Avogadro's law.
- **Explain** the significance and different units of ideal gas constant.
- **Distinguish** between real and ideal gases.
- **Explain** why real gases deviate from the gas laws.
- **Define** and describe the properties of Plasma.

INTRODUCTION

In previous classes we studied that anything which has mass and occupies space is called matter. In universe, there are four fundamental states of matters solid, liquid, gas and plasma. Among all the four states of matter, gas has quite different behaviour. The particles of gases are not orderly arranged and have very weak attractive forces as compared to solids and liquids.

The word gas is derived from Greek Khaos which means emptiness or gap. Due to the empty spaces, gas particles move in all directions. They move in straight lines until they collide with each other or the walls of the container. Based on elemental composition gases are classified into mono atomic, diatomic and polyatomic. Helium, Neon, Argon etc are mono atomic gases where as Nitrogen, Hydrogen, Oxygen, Fluorine, Chlorine are diatomic gases. There are many poly atomic gases found in the earth such as methane, ethane, carbon dioxide etc. Air is mainly the mixture of nitrogen and oxygen gases.

Gases play significant role in our daily life. The living things (animals and plants) need oxygen for breathing, plants make food using carbon dioxide, certain industries and power stations use natural gas for their production, different vehicles use compressed natural gas as an alternative to gasoline etc. Moreover, our kitchens also rely on natural gas to a greater extent.

4.1 KINETIC MOLECULAR THEORY OF GASES

The general behaviour and properties of all gases are similar and well explained by Kinetic molecular theory. This theory was given by Swiss mathematician Daniel Bernoulli (1738) and then extended by Maxwell and Boltzmann. Thereafter, Rudolf Clausius in 1857 derived an equation on the basis of kinetic theory of gases which provided a base for gas laws. Basically, this theory was developed in reference to ideal gases but can reasonably be applied to real gases too.

This theory explains the macroscopic properties like temperature, pressure, volume etc and transport properties like diffusion, viscosity as well.

4.1.1 Postulates of Kinetic Molecular Theory

The basic postulates of kinetic theory of gases are as follows:

(i) Gases consist of a large number of tiny particles called molecules. Molecules may be mono-atomic (He, Ne, Ar), diatomic (O_2, N_2) or poly atomic (CH_4, C_4H_{10}).

(ii) Gas molecules are far away from each other and occupy negligible volume as compared to the total volume of the container

(iii) Gas molecules are in continuous random motion, traveling in a straight line until they collide with each other or with the walls of the container. The average distance covered by the gas between successive collisions is known as mean free path.

(iv) Gas particles undergo elastic collisions (a collision in which no loss or gain of energy takes place).

(v) Gas exerts pressure when molecules collide with the walls of container.

(vi) There is no force of attraction or repulsion found among ideal gas molecules, each molecule acts as quite independently.

(vii) The average kinetic energy of gas molecules depends upon absolute temperature. Thus when absolute temperature increases kinetic energy of the molecules increases.

$$KE \propto \text{Absolue temperature}$$

On the basis of above postulates, R. J. Clausius derived a kinetic equation.

$$PV = \frac{1}{3} mN\bar{C}^2$$

Where, P = pressure

V = volume

m = mass of a single molecule of a gas

N = number of moles of gas molecules

$\bar{C}^2$ = mean square velocity of the gas molecules

Under the given conditions, the gas molecules do not possess the same velocities, instead mean square velocity is taken for the molecules in the above equation. If n_1 molecules possess c_1, velocity, n_2 molecules with velocity c_2, and so on,

then mean square velocity can be calculated as.

$$\bar{C}^2 = \frac{n_1C_1^2 + n_2C_2^2 + n_3C_3^2 + \ldots}{n_1 + n_2 + n_3 + \ldots}$$

Where, $\bar{C}^2$ is the average of the square of all the possible velocities.

4.1.2 Pressure and its units

The atmosphere of our planet is mixture of gases. It exerts a measurable pressure on all living and non-living objects which have existence on the earth's crust. At 0 °C on sea level our body bears a pressure of 760 torr.

The pressure of air is measured by Barometer this device consists of long glass tube filling with mercury with upper end closed. A manometer is a type of Barometer that can be used to measure the pressure of a gas enclosed in a container.

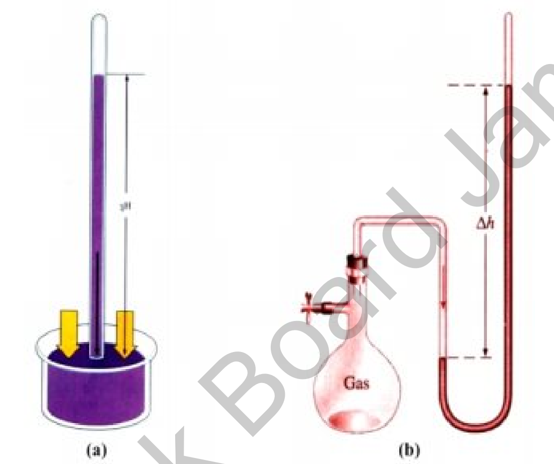


Fig 4.1: (a) Barometer (b) Manometer

We are quite familiar that when a gas is pumped into an automobile tire, it becomes hard which is due to the pressure of gas inside the tire. This pressure is due to the collision of gas molecules on the wall of tire. Blowing up a balloon provide clear evidence that gas exert pressure on the wall of its container. Thus, pressure is defined as, **"The magnitude of force that is applied on the surface of an object per unit area"**.

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

According to the S.I system the unit of force is Newton and area is square meter.

Thus, the unit of pressure in SI measurement is N/m^2 which is known as pascal (Pa). Pascal is defined as, **"One Newton force that is distributed at an area of one square meter"**.

Although Pascal (Pa) is used in many scientific works but it is an in convenient size for most chemical measurements. Certain alternative units of pressure and their inter relation are given as.

101325 Pa = 1 atm = 760 mm Hg = 14.7 psi

Do You Know?

The air pressure decreases as altitude increases due to the following two facts.
(i) Gas particles of air are pulled down due to gravity and becomes more denser near earth surface.
(ii) Atmospheric depth is biggest at sea level and decreases at higher altitude.

Example 4.1

The pressure of gas filled in automobile tire is generally measured in psi convert 32.8 psi into (i) atmosphere (ii) Kpa (iii) torr

Solution:

(i) Psi to atmosphere

Since 14.7 psi = 1 atm

$$32.8 \text{ psi} = \frac{32.8}{14.7} = 2.23 \text{ atm}$$

(ii) Psi to Kilo Pascal

Since 14.7 psi = 101.325 Kpa

$$32.8 = \frac{101.325}{14.7} \times 32.8 = 226.085 \text{ KPa}$$

(iii) Psi to torr

Since 14.7 psi = 760 torr

$$32.8 \text{ psi} = \frac{760}{14.7} \times 32.8 = 1695.8 \text{ torr}$$

4.1.3 Atmospheric pressure and its effect on weather

Gases such as oxygen, water vapors, nitrogen, ozone etc are very important for animals and plants and affect on Earth's atmosphere. Oxygen is used for breathing, nitrogen controls the oxidation reactions and ozone saves our atmosphere from the dangerous radiations of the sun. Similarly, the formation of clouds, temperature conditions, melting, boiling and vaporization processes are also directly related with weather conditions.

The atmospheric pressure is a thermodynamic property which is created by the mass of air molecules on the surface of earth. The size, number and nature of the molecules determine the density and temperature of the air. As the number and motion of the particles increases, atmospheric pressure increases and vice versa.

Low atmospheric pressure system is also called a depression. An area which has low pressure as compared to the surrounding is generally warmer where moist warm air rises up and cools down. This area has suffocating weather and invites winds, clouds and precipitation. Here the wind blows in anticlockwise direction. During day time there is comparatively moderate temperature while nights are warmer due to the trapping of solar radiations in cloudy weather.

The area which experiences a high air pressure as compared to surrounding is said to be in high pressure system. In such area, dry, cold and dense air moves downward to the ground. Here the wind blows in clockwise direction. Thus, clear sky and calm weather are developed.

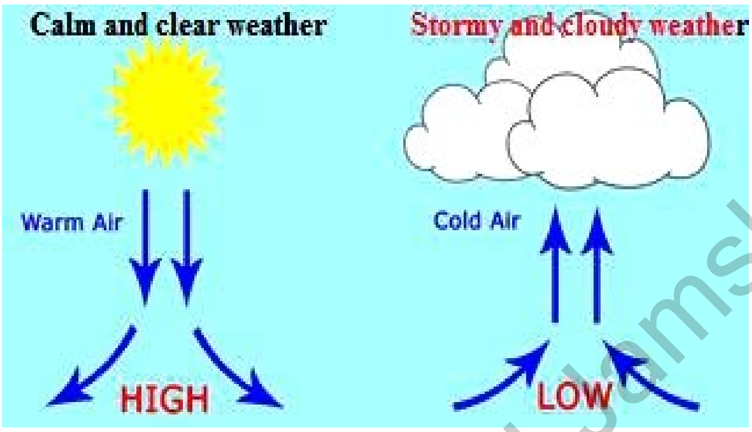


Fig 4.2: Effect of atmospheric pressure on weather

4.2 ABSOLUTE TEMPERATURE SCALE ON THE BASIS OF CHARLES LAW

Thermal energy is related to every form of matter, either matter is in condensed form or gaseous state. The measure of the hotness or coldness of a body is called temperature. Generally, temperature is measured with Celsius and Fahrenheit scales. However, there is also third scale which is known as Kelvin or Absolute scale which was introduced by Lord Kelvin (1824-1907). The temperature which is measured on this scale is known as Absolute temperature.

J. Charles showed a relationship between volume of a gas and absolute temperature while maintaining the pressure constant. According to him, when absolute temperature of a gas increases, its volume also increases because a rise of temperature increase the kinetic energy of gas molecules. This makes the molecules to move freely. As a result, the volume of a gas is increased.

4.2.1 Brief recall of Boyle's Law and Charles Law

Robert Boyle (1660 A. D) studied the effect of pressure changes on the volumes of ideal gases.

This law states as **"The volume of the given mass of a gas is inversely proportional to its pressure at constant temperature"**.

$V \propto \frac{1}{P}$ (At constant temperature)

This statement shows that if pressure of gas increases, its volume decreases with the same proportion provided that the temperature remains constant.

It is common observation that matter expands on heating and contracts on cooling. The change in volume due to expansion and contraction is very small in case of solids and liquids while gases exhibit enormous change due to the presence of large intermolecular spaces. The changes in volumes of gases due to the changes in temperature at constant pressure were studied by French Scientist Jacques Charles (1746-1823). He explained that the volume of given mass of a gas increases or decreases by 1/273 times of its original volume at 0°C for every degree rise and fall of temperature at given pressure. In nutshell it states that, **"The volume of given mass of a gas is directly proportional to the absolute temperature at a given pressure"**.

$V \propto T$ (At constant pressure)

4.2.2 Graphical Explanation of Absolute zero

When the graph is plotted between temperature (T) on x-axis and volume (V) on y-axis, a straight line is obtained. This straight line in upward direction shows that volume increases with the increase of temperature. If this straight line is further extended downward, it will intercept the temperature axis at –273.15ºC.

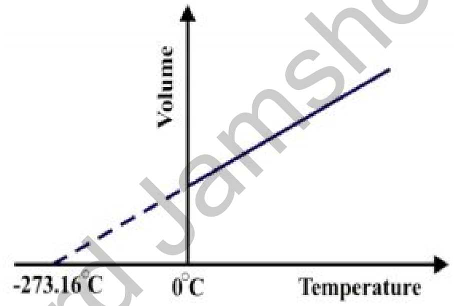


Fig 4.3: Graphical representation of the absolute zero.

According to Charles, at 0 K (-273.15°C) the volume of a gas should be zero. Actually no real gas can achieve this lowest possible temperature and before –273.15°C all gases are condensed in to liquids. Zero Kelvin is also known as Absolute zero. Absolute zero is a theoretical temperature and defines as, **"The temperature at which volume of an ideal gas becomes equal to zero"**. The concept of absolute zero cannot be applied to real gases. The inter-conversion between Celsius and Kelvin scale is given below:

$$K = {}^{o}C + 273$$

Table 4.1: Conversion scale of Kelvin and Celsius temperature

Kelvin Scale	Celsius Scale
Kelvin = Celsius +273	
0	- 273.16
25	- 248.16
50	- 223.16
100	- 173.16
150	- 123.16
200	- 73.16
250	- 23.16
273 K	0°C

Self Assessment

Convert the following celsius temperatures into Kelvin temperatures.
(i) -12 °C (ii) 27 °C (iii) 43 °C (iv) 110 °C (v) 786 °C

4.3 AVOGADRO'S LAW

In 1811, Amedeo Avogadro, an Italian scientist, gave a relationship between volume and number of moles of gases at fixed temperature and pressure. This relationship is known as "Avogadro's Law" and defines as, **"Under the similar conditions of temperature and pressure equal volumes of all gases contain equal number of moles"**.

Thus under similar conditions of temperature and pressure 1dm^3 any gas contains same number of molecules. We know that 1 mole of any gas at standard temperature (0 ^oC) and pressure (1 atm) occupies 22.4 dm^3 (**Molar Volume**) and contains 6.02 x 10^{23} molecules. It means that two or more gases having same volume must have same number of molecules and moles but different masses.

According to Avogadro's Law, **"the volume of a gas is directly proportional to the number of moles if the pressure and temperature are kept constant"**.

$V \propto \boldsymbol{n}$ (At constant temperature and pressure)

Or $V = K\boldsymbol{n}$

$$\frac{V}{n} = k$$

It means that the ratio of volume to number of moles of a gas remains constant. Suppose a gas of $\mathbf{n_1}$ moles is enclosed in a vessel of $\mathbf{v_1}$ volume. If we insert more gas, the volume increases to $\mathbf{v_2}$ hence.

$$\frac{V_1}{n_1} = k \; ; \; \frac{V_2}{n_2} = k$$

And

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Example 4.2

A cylinder contains 2.2 moles of oxygen gas at S.T.P. When more oxygen gas is pumped into the cylinder, the volume of a gas is changed from 2.0 dm^3 to 3.4 dm^3. Calculate how many moles of the oxygen gas are added to the cylinder?

Solution:

$V_1 = 2.0$ dm^3

$n_1 = 2.2$ moles

$V_2 = 3.4$ dm^3

$n_2 = ?$

Applying Avogadro's Law

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Substituting the values

$$\frac{2.0}{2.2} = \frac{3.4}{n_2}$$

$$n_2 = \frac{3.4 \times 2.2}{2.0}$$

$n_2 = 3.74$ moles

$\Delta n = n_2 - n_1$

$\Delta n = 3.74 - 2.2 = 1.54$ moles

Self Assessment

At standard temperature and pressure 26.4 dm^3 of a gas contains 1.26 moles. If 0.25 moles are added to the gas, what will be the new volume of the gas?

4.4 IDEAL GAS EQUATION

Boyle's law, Charles's law and Avogardo's law can be combined into a single statement known as ideal gas laws or ideal gas equation This equation describes the behavior of a gas on the basis of relationship between volume and other variables like pressure, temperature and number of moles. When the values of any three of the variables **P,V,T & n** are known, the value of fourth can be calculated by using ideal gas equation.

4.4.1 Derivation of an ideal gas equation

According to Boyle's Law $V \propto \frac{1}{p}$ (At constant temperature)

According to Charle's Law $V \propto T$ (At constant pressure)

According to Avogadro's Law $V \propto n$ (At constant temperature and pressure)

By combining these laws we get:

$$V \propto \frac{n\,T}{P}$$

$$\text{or } V = \frac{RnT}{P}$$

And $PV = nRT$ - - - - (i)

Where, R is general gas constant.

For one mole of a gas this equation can be written as:

$PV = RT$

$$R = \frac{PV}{T} \text{ - - - - (ii)}$$

If we change the pressure from P_1 to P_2 and temperature from T_1 to T_2, then this equation can be written as:

$$R = \frac{P_1V_1}{T_1} \text{ and } R = \frac{P_2V_2}{T_2}$$

Therefore,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{P_nV_n}{T_n} --- \text{(iii)}$$

Example 4.3

Laughing gas (N_2O) at 30°C and 820 torr pressure occupies a volume of 10.32 dm^3. Calculate the volume that it will occupy at standard temperature and pressure.

Solution:

$T_1 = 30°C = 30 + 273 = 303k$

$P_1 = 820$ torr

$V_1 = 10.32\ dm^3$

$V_2 = ?$

$T_2 = 273k$

$P_2 = 760$ torr

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$V_2 = \frac{820 \times 10.32 \times 273}{303 \times 760}$$

$V_2 = 10.03\ dm^3$

Self Assessment

An steal gas cylinder has a capacity of 15.8 dm^3 and filled with 785g of Helium gas at a temperature of 20°C. Calculate the pressure of Helium in the cylinder (mol. mass of Helium = 4 g/mol).

4.4.2 Gas constant and its units

Equation (ii) reveals that gas constant is the ratio of the product of pressure and volume to the absolute temperature. Its numerical value can be determined at STP.

(a) Value of R when pressure is in atmosphere and volume in dm^3.

We know that one mole of an ideal gas at S.T.P (one atmosphere pressure and 273 K) occupies 22.4 dm^3.

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

$$R = \frac{1 \text{ atmosphere} \times 22.4\ dm^3}{1 \text{ mole } 273k}$$

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

(b) Value of R when pressure is in N/m^2 and volume in m^3 (S.I unit):

According to the S.I system, pressure is measured in N/m^2 and volume in m^3:
since 1 atm = $101325 N/m^2$
And 1 $dm^3 = 0.0224\ m^3$

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

$$R = \frac{101300 \times 0.0224}{1\ mol \times 273}$$

$R = 8.31 Nm\ mol^{-1} k^{-1}$
$R = 8.31\ J\ mol^{-1}\ k^{-1}$ (Because 1 Nm = 1 J)
Since 1 cal = 4.18 J, therefore the value of R may also be written as
$R = 1.99\ cal\ mol^{-1}\ K^{-1}$

Application of Ideal gas equation

Ideal gas equation is used to determine the molecular mass and density of gases.

(a) Molecular mass of the gas

As we know $\quad \text{Moles} = \dfrac{\text{Given mass (m)}}{\text{molar mass (M)}}$

Now put this value of number of moles in equation (i)

$$PV = \frac{m}{M}RT \ldots\ldots\ldots\ldots(iv)$$

$$MPV = mRT$$

$$M = \frac{mRT}{PV} \ldots\ldots\ldots\ldots(v)$$

This equation is used to calculate the molecular mass of a gas.

(b) Density of the gas

By manipulating eq. (iv) we get

$$PM = \frac{m}{V}RT$$

Since d= m/V, therefore PM = dRT

Thus, $\quad d = \dfrac{PM}{RT} \ldots\ldots\ldots\ldots(vi)$

This equation is used for the determination of density of a gas. This equation indicates that density is directly proportional to pressure and molecular mass inversely proportional to absolute temperature.

Example 4.4

Calculate the density of oxygen gas at 45 °C when the gas is confined in cylinder at 1.54 atmospheric pressure.

Solution:

d = ?

T = 45°C = 45 + 273 = 318 K

P = 1.54 atm

R = 0.0821 atm dm^3 mol^{-1} K^{-1}

M_{O2} = 32 g mol^{-1}

The formula of density derived from general gas equation is given as

$$d = \frac{PM}{RT}$$

By substituting the values, we get

$$d = \frac{1.54 \times 32}{0.0821 \times 318}$$

d = 1.889 g/dm^3

Self Assessment

A chemist has synthesized a gas and find that its density is 1.88 g/dm^3 at 27°C and 1 atm pressure. Calculate its molar mass.

4.5 DEVIATION FROM IDEAL GAS BEHAVIOUR

A gas which strictly obeys gas laws (Boyle's law, Charles's law etc) or general gas equation (PV = nRT) under all conditions of temperature and pressure and behaves according to the kinetic molecular theory is known as an ideal gas. Actually none of the known gases exactly follow the ideal gas laws and called as real gases.

Real gas shows deviation from gas laws particularly at high pressure and low temperature. Therefore when the temperature of gas is lowered the attractive forces become significant, eventually the gas liquefies. This property of real gas causes deviate them from ideal behavior.

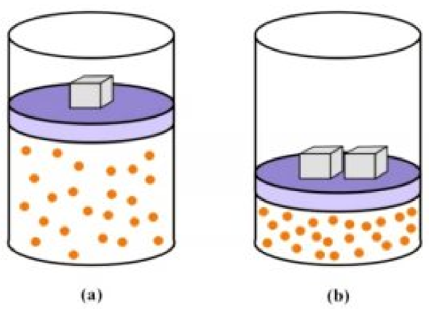


Fig 4.4: Effect of pressure on the volume of gas.

4.5.1 Graphical explanation of deviation of real gases

To understand the attitude of real gases graphically, the general gas equation should be amended as.

$PV = Z(nRT)$ or $Z = \frac{PV}{nRT}$

Where Z is called as compressibility factor, its value is unity for an ideal gas, less than 1 for real gas showing negative deviation and more than unity for real gases showing positive deviation.

The graph plotted between Z and P for an ideal gas and various real gases provides the following information.

(i) All gases reaches a value z = 1, when the pressure approaches to zero. This reveals that all gases tend to act like ideal gas at very low pressure.

(ii) The extent of deviation of real gases from ideality is based on pressure, temperature and the nature of gases.

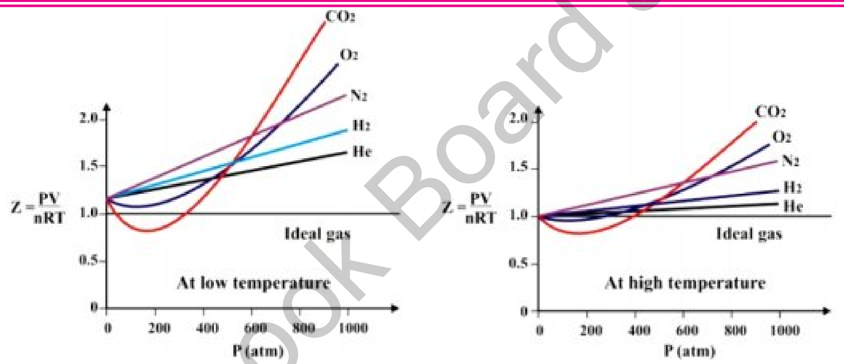


Fig 4.5: Deviation of real gases from ideality (a) at low temperature (b) at high temperature

4.5.2 Causes of deviation of real gases from ideal behaviour

To analyze why real gases deviate from ideal behavior, we should know the basics on which the ideal gas equation was formulated. The ideal gas equation was obtained from certain assumptions of kinetic molecular theory; two main assumptions are given below.

(i) The actual volume of the gas molecules is negligibly small as compared to the total space of the container.

This assumption remains valid at low pressure where the volume of gas is much larger due to large intermolecular spaces but at high pressure gas is in the compressed state and the volume of gas molecules becomes significant as compared to the volume of gas enclosed in the container.

(ii) Gas molecules have neither attractive nor repulsive forces.

This assumption is valid at low pressure and high temperature at which molecules tend to be far apart, because these forces diminish rapidly as the distance between molecules increases. But at high pressure and low temperature intermolecular forces becomes significant because molecules tend to be close together.

4.6 VAN DER WAAL'S EQUATION

In order to modify the two conflicting postulates of the Kinetic Molecular theory regarding volumes and intermolecular attraction of gas molecules, a Dutch scientist, J.D. van der Waal (1873), gave a mathematical solution by the correction in molecular volume and intermolecular forces. This is known as van der Waal equation.

Do You Know?

Among all real gases Helium acts most likely as an ideal gas at room temperature.

4.6.1 Volume Correction

When pressure is applied to a gas, the molecules come closer to each other. By continuous increase in pressure a point is reached when molecules cannot further be compressed since repulsive forces are created. This indicates that gas molecules have definite volume. Although this is very small but not negligible. Keeping in view the definite volume of gas molecules, van der Waal calculated the actual volume of a gas as follows:

$V = V_{vessel} - nb$ ……………….. (i)

Where,

V = free volume

V_{vessel} = Volume of the vessel in which gas molecules are present.

n = number of moles

b = excluded volume of gas molecules per mole in highly compressed gaseous sate. It is theoretical volume.

The value of 'b' depends upon the size of gas molecules. Here it should be noted that **'b'** is not actual volume of a gas but it is assumed four times the actual volume of molecules.

$b = 4V_{molecule}$

Where, $V_{molecule}$ is actual volume of one mole of a gas.

4.6.2 Pressure Correction

van der Waal also corrected the pressure produced by the molecules of the real gases. As we have already discussed that the intermolecular attraction comes into play when the molecules are brought close together by squeezing the gas. Consider a molecule **A** in the interior of the gas, which is completely surrounded and attracted by the other gas molecules **B**. The resultant attractive force on the molecule **A** due to all the surrounding **B** molecules is zero as shown in figure 4.5b. However, as this molecule **A** approaches to strike the wall of vessel, it experience a net inward pull due to the attractive forces of molecules of **B**. It means that pressure produced on the wall would be little bit lesser than pressure of an ideal gas molecule.

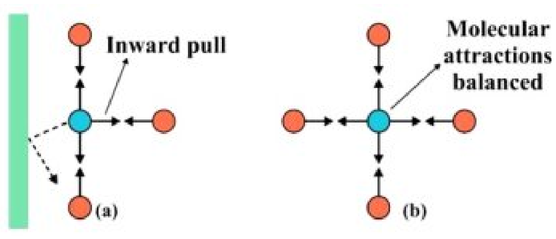


Fig 4.6: pressure of molecule which strike on the wall is less than ideal gas pressure

Therefore,

$$P_{observed} = P_{ideal} - P_{less}$$

If observed pressure is simply indicated by P, ideal pressure P_i and less pressure P_L

Then equation will be:

$P = P_i - P_L$

Or

$P_i = P + P_L$ ……………….. (ii)

Since P_L is the pressure drop due to backward pull of striking molecules, it depends upon number of particles (A and B) per unit volume.

$$P_L \propto [A] . [B]$$

$P_L \propto \frac{n}{V}\frac{n}{V}$ or $P_L = \frac{an^2}{V^2}$

Insert the value of P_L in equation (ii).

$P_i = P + \frac{an^2}{V^2}$ ……………… (iii)

The corrected pressure and volume is now put in general gas equation, we get.

$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$

If n = 1 then equation will be:

$\left(P + \frac{a}{V^2}\right)(V - b) = RT$ ………….. (v)

This is van der Waal's equation. Here 'a' and 'b' are van der Waal's constants.

Table 4.2: van der Waal's constants for some common real gases

Gas	'a' (atm dm^6 $mole^{-2}$)	'b' (dm^3 $mole^{-1}$)
Hydrogen	0.0247	0.0266
Nitrogen	1.390	0.0391
Oxygen	1.360	0.0318
Methane	2.253	0.0428
Carbon dioxide	3.590	0.0428
Ammonia	4.170	0.0371
Sulphur dioxide	6.170	0.0564
Chlorine	6.493	0.0562

Units for van der Waal's constants 'a' and 'b':

Since $P = \frac{an^2}{V^2}$, hence $a = \frac{PV^2}{n^2}$

By substituting the units of P, V and n.

a = atm dm^6 $mole^{-2}$

Since "b", represents the volume per mol of gas, its unit dm^3/mol.

Example 4.5

One mole of ammonia gas is kept in a cylinder of 5.5dm^3 at 27°C. Assuming ammonia gas as a real gas determine its pressure. The ven der Waal constant for ammonia are a = 4.17 atm dm^6 mol^{-2} and b = 0.0371 dm^3 mol^{-1}.

Solution:

n = 1 mole

T = 27°C = 27 + 273 = 300K

V = 5.5 dm^3

R = 0.0821 atm dm^3 mol^{-1} K^{-1})

P = ?

According to van der waal equation.

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

$$\left(P + \frac{4.17 \times 1}{(5.5)^2}\right)(5.5\text{-}0.0371) = (1)(0.082 \times 300)$$

$$(P + 0.138)(5.462) = 24.6$$

$$P = 4.365 \text{ atm}$$

Example 4.6

Two moles of oxygen gas is kept in a vessel of 15.5dm^3 at a temperature of 37°C. Calculate the pressure exerted by the gas if

(a) gas behave as ideal (b) gas behave is non ideal

The ven der waal constant of O_2 gas are given as

a = 1.36dm^6 atm mol^{-2} and b = 0.0318 dm^3 mol^{-1}.

Solution:

n = 2 moles

T = 37 °C = 37+273 = 310 K

V = 15.5 dm^3

R = 0.0821 atm dm^3 mol^{-1} K^{-1})

P = ?

i) General gas equation is used if gas behave ideally,

PV=nRT

$$P = \frac{2 \times 0.0821 \times 310}{15.5} = 3.28 \text{ atm}$$

ii) van der Waal equation is used if it behaves as non ideal gas

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

$$\left(P + \frac{1.36 \times 4}{(15.5)^2}\right)(15.5\text{-}2 \times 0.0318) = 2 \times 0.082 \times 310$$

$$P = 3.27 \text{ atm}$$

4.7 DALTON'S LAW OF PARTIAL PRESSURE

In 1801, an English chemist John Dalton gave a law dealing with the pressure exerted by non-reacting mixture of gases. This law states as **"The total pressure exerted by a mixture of non reacting gases in a closed vessel (fixed volume) is always equal to the sum of their individual pressures at a constant temperature"**.

If P_A, P_B, P_C etc are the individual pressures of gases in the mixture, the total pressure of the mixture is written as.

$$P_t = P_A + P_B + P_C + \text{------}$$

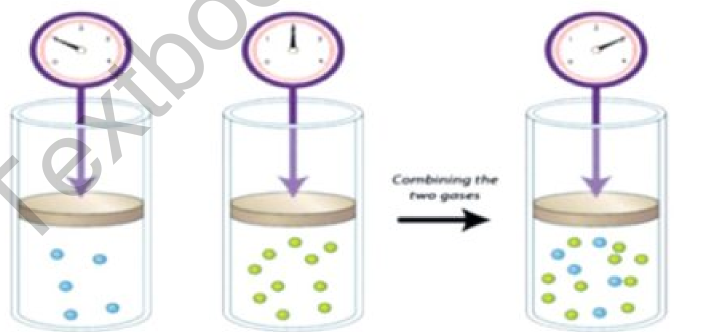
The pressure exerted by each individual gas in a mixture is called the partial pressure of that gas; this is the pressure that individual gas would exert if it were alone in the same container at the same temperature.

> **Do You Know?**
>
> Atmospheric pressure is due to weight of the blanket of air surrounding earth. With greater depth of atmosphere, more air is pressing down from above. Therefore, the atmospheric pressure is greatest at sea level and it decreases with increasing altitude. On the top of mount Everest atmospheric pressure falls to about one-third of the pressure at sea level.

Suppose we have three empty cylinders of equal capacity (1 dm^3), O_2 and N_2 gases are filled separately in first two cylinders at constant temperature. Let the pressure exerted by the O_2 gas and N_2 gas are 500 torr and 700 torr respectively. If now these gases are transferred into third empty cylinder under the condition of same temperature; the pressure exerted by the mixture gases is found to be 1200 torr which is exactly equal to the sum of partial pressure of O_2 and N_2.

According to Dalton's law

$Pt = P_{O_2} + P_{N_2} = 500 + 700 = 1200$ torr



$P_{O_2} = 500$ torr $\quad P_{N_2} = 700$ torr $\quad P_{O_2} + P_{N_2} = 1200$ torr

Fig 4.7: Pressure of individual gases and of their mixture

Since in the mixture, all non reacting gases behave independently, the molecules of each gas has equal opportunity to strike with the walls of cylinder and exert its own pressure without the involvement of pressure of other gases. Therefore general gas equation can be applied to individual gases in mixture and may be written as.

$$P_A V = n_A RT \qquad \text{or} \qquad P_A = \frac{RT}{V} n_A \; . \; . \; . \; . \; . \; \text{(i)}$$

$$P_BV = n_BRT \qquad \text{or} \qquad P_B = \frac{RT}{V} n_B \; ..... (ii)$$

Since R, V and T are constant, therefore

$$P_A \propto n_A \quad \text{and} \quad P_B \propto n_B$$

Thus, at constant temperature and volume, the partial pressure of each individual gas in a mixture is proportional to its number of moles.

For the mixture of gases, the general gas equation may be written as

$$P_tV = n_tRT \qquad \text{or} \qquad P_t = \frac{RT}{V} n_t \; ..... (iii)$$

Dividing equation (i) by (iii). we get

$$\frac{P_A}{P_t} = \frac{n_A}{n_t} \; .... (iv)$$

Similarly by dividing equation (ii) with (iii), we get

$$\frac{P_B}{P_t} = \frac{n_B}{n_t} \; .... (v)$$

Since mole fraction is the ratio of number of moles of individual gas and the total number of moles of all gases present in the mixture.

$$\frac{n_A}{n_t} = X_A \qquad \text{and} \qquad \frac{n_B}{n_t} = X_B, \text{ hence}$$

Equation (iv) and (v) are now reduced as

$$P_A = X_A P_t \; .... (vi)$$

$$P_B = X_B P_t \; .... (vii)$$

From the equation (vi) and (vii), it is concluded that partial pressure of any gas of the mixture is equal to the product of moles fraction of that gas and total pressure of the mixture. It should be noted that mole fraction of any gas in the mixture is less than one but sum of the mole fractions is always equal to one.

Applications of Daltons Law of partial pressure:

(i) Pressure of gases collected over water: When a gas is collected over water by downward displacement of water in a gas jar, the pressure of dry gas can be calculated by Daltons Law. When gas passes through water, it becomes moist and the total pressure will be equal to the sum of partial pressure of dry gas and water vapours (aqueous tension).

$$\mathbf{P_{total} = P_{dry\ gas} + P_{water\ vapours}}$$

$$\mathbf{P_{dry\ gas} = P_{total} - P_{water\ vapours}}$$

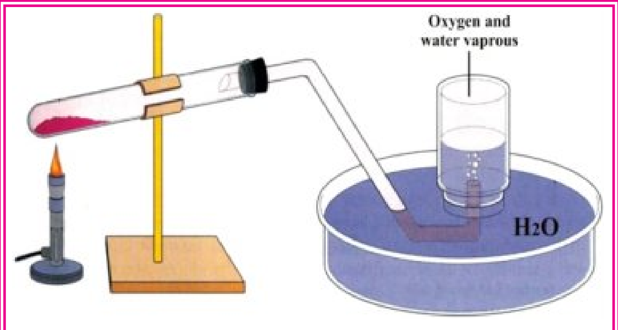


Fig 4.8: Pressure of gas collected over water is equal to the sum of pressure of dry gas and vapours

(ii) Maintenance of oxygen pressure at high altitudes: The pressure of air on high altitudes is lower than sea level due to decrease in number of molecules of gases. Normally, respiration depends upon the difference between the partial pressure of oxygen in the air (159 torr) and in the lungs (116 torr). At higher altitudes, due to low partial pressure of oxygen causes the problem in the process of respiration.

(iii) Maintenance of oxygen pressure for deep sea divers: Opposite to altitude, as distance increases downward in the sea, partial pressure of oxygen increases. At the depth of 40 meters, pressure increases to five times. This increased pressure also causes problem in respiration. Therefore, deep sea divers use the SCUBA (Self Contained Underwater Breathing Apparatus), breathing tank for respiration. Scuba contains 96% helium gas and 4% oxygen gas.

Example 4.7

A 20 dm^3 cylinder is filled with 4.25 moles of oxygen gas and 12 moles of helium gas at 25 °C. Calculate the total pressure of gas mixture and partial pressures of oxygen and helium gases in the cylinder?

Solution:

$V = 20\ dm^3$

$n_{Oxygen} = 4.25$ moles

$n_{Helium} = 12$ moles

$n_t = n_{Oxygen} + n_{Helium}$

$n_t = 4.25 + 12 = 16.25$ moles

$T = 25\ °C = 25 + 273 = 298\ K$

$P_t = ?$

$P_{Oxygen} = ?$

$P_{helium} = ?$

According to general gas equation,

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

$$P_t = \frac{16.25 \times 0.0821 \times 298}{20}$$

$P_t = 19.88$ atm

Now the partial pressure of each gas in the mixture can be calculated by substituting the moles and total pressure in the given formula of Dalton's law.

$$P_{(oxygen)} = \frac{n_{O2}}{n_t} \times P_t$$

$$P_{oxygen} = \frac{4.25}{16.25} \times 19.88 = 5.19 atm$$

Similarly,

$$P_{helium} = \frac{12}{16.25} \times 19.88$$

$P_{helium} = 14.68$ atm

Example 4.8

Oxygen gas is produced by heating potassium nitrate

$$2KNO_3 \longrightarrow 2KNO_2 + O_2$$

The gas is collected over water. If 225cm³ of gas is collected at 25°C and 785mm Hg total pressure, what is the mass of O_2 gas collected? (Pressure of vapours at 25°C is 23.8 mm Hg)

Solution:

According to Dalton's law, the total pressure of mixture is equal to the sum of partial pressure of O_2 and vapours thus

$P_t = P_{(O2)} + P(\text{vapours})$

Hence

$P_{(O2)} = 785 - 23.8 = 761.2$ mmHg

Convert the units of pressure from mm Hg to atm & unit of volume from cm^3 to dm^3

$$P = \frac{761.2}{760} = 1.001 \text{ atm;}$$

$$V = 225cm^3 = \frac{225}{1000} = 0.225dm^3$$

T=25°C=25+273=298K

We can use general gas equation to determine the number of moles of oxygen gas

$PV = nRT$

$$n = \frac{1.001 \times 0.225}{0.082 \times 298} = 9.2 \times 10^{-3} \text{moles}$$

Now, mass of $O_2 = 9.2 \times 10^{-3} \times 32 = 0.294g$

Self Assessment

The mole fraction of oxygen in air is 0.2093, determine the partial pressure of oxygen in air if the atmospheric pressure is 760 torr.

4.8 GRAHAM'S LAW OF DIFFUSION AND EFFUSION

In the kinetic Molecular theory, we studied that gas molecules possess kinetic energy and move randomly. The movement of gas molecules through tiny hole into the region of low pressure is termed as **Effusion**. The diameter of tiny hole is considerably smaller than mean free path. Slow escaping of air from a tyre pinhole is an example of effusion. On the other hand, the process in which molecules of different gases homogenously intermix with each other is known as **Diffusion**. For example the spreading of fragrance of a rose flower or a perfume is because of diffusion. Lighter gases diffuse more rapidly than heavier gases.

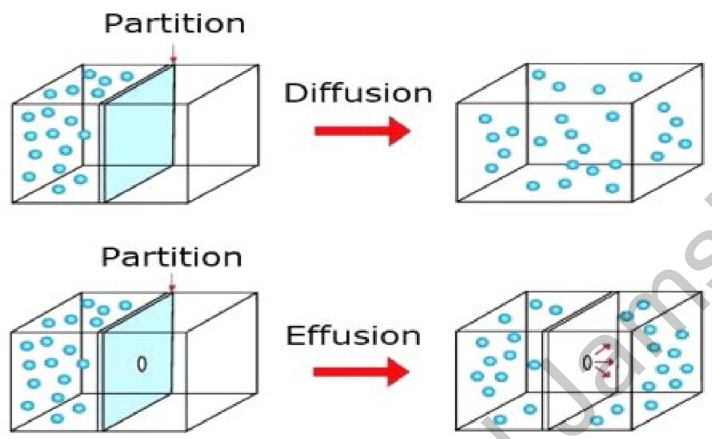


Fig 4.9: Diffusion and Effusion of gases

In 1831, Thomas Graham introduced a quantitative relationship between the rate of effusion or diffusion of gases and their densities by maintaining the temperature and pressure. After seventeen years in 1848, he modified his law by establishing the relation between rates of effusion or diffusion with molar masses in addition to densities of gases. This is known as Graham's law of diffusion which is stated as, **"At constant temperature and pressure, the rate of effusion or diffusion of a gas is inversely proportional to the square root of its density or molar mass"**.

If 'r' is rate of diffusion and 'd' is density of a gas then according to Graham's law:

$$r \propto \frac{1}{\sqrt{d}} \quad \text{(At constant temperature and pressure)}$$

$$r = \frac{k}{\sqrt{d}} \quad \text{Where, k is constant for proportionality.}$$

If we consider two gases 1 and 2, having rates of diffusion r_1 and r_2 and densities d_1 and d_2 respectively then:

$$r_1 = \frac{k}{\sqrt{d_1}} \text{-----(i)}$$

$$r_2 = \frac{k}{\sqrt{d_2}} \text{-----(ii)}$$

Now divide equation (i) with equation (ii) and we get:

$$\frac{r_1}{r_2} = \frac{\sqrt{d_2}}{\sqrt{d_1}} \text{-----(iii)}$$

The density of a given gas is directly proportional to its molecular mass, therefore,

$$d_1 = \frac{M_1}{V} \qquad \text{and} \qquad d_2 = \frac{M_2}{V}$$

We can write the equation (iii) as:

$$\frac{r_1}{r_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}} \text{-----(iv)}$$

Where, M_1 and M_2 are the molar masses of gases.

Applications of Graham's Law:

(i) Densities and molar masses of different gases can be determined by using Graham's law.

(ii) The effect of toxic gases can be reduced by diffusing them into air.

(iii) Various isotopes of gases can be separated by applying Graham's law because isotopes posses different masses and have different rate of diffusion.

Example 4.9

Compare the rates of diffusion of helium (H_e) and methane (CH_4) gases.

Solution:

Mass of He = 4 a.m.u

Molecular mass of CH_4 = 12 + 4 = 16 a.m.u

$$\frac{rHe}{rCH_4} = \sqrt{\frac{MCH_4}{MHe}}$$

$$\frac{rHe}{rCH_4} = \sqrt{\frac{16}{4}} = \frac{4}{2} = \frac{2}{1} \text{ Ans}$$

Thus helium diffuses two times as fast as CH_4.

Example 4.10

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

$$\frac{\gamma_A}{\gamma_B} = \sqrt{\frac{M_B}{M_A}}$$

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

By applying the square root on both sides,

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

$M_B = (1.5)^2 \times 16$

$M_B = 2.25 \times 16$

$M_B = 36$

Example 4.11

At a specific temperature and pressure, it takes 290s for a 1.5dm^3 sample of He to effuse through a porous membrane. Under similar conditions, if 1.5dm^3 of an unknown gas "X" takes 1085s to effuse, calculate the molar mass of gas "X".

Solution:

$$\frac{r_{He}}{r_x} = \sqrt{\frac{M_x}{M_{He}}}$$

$$\frac{V_{He}/t_{He}}{V_x/t_x} = \sqrt{\frac{M_x}{M_{He}}}$$

$$\frac{1.5/290}{1.5/1085} = \sqrt{\frac{M_x}{4}}$$

$$\frac{1085}{290} = \sqrt{\frac{M_x}{4}}$$

Squaring on both sides we get:

$$\left(\frac{1085}{290}\right)^2 = \frac{M_x}{4}$$

M_x=56g/mol

Self Assessment

If it takes 8.5 seconds for 5cm^3 of CO_2 gas to effuse through a porous material at a particular temperature and pressure. How long would it take for 5 cm^3 of SO_2 gas to effuse from the same container at the same temperature and pressure?

4.9 LIQUEFACTION OF GASES

Liquefaction is a physical process or technique in which gases are converted into liquids. Gas molecules at high pressure and low temperature come closer and attractive forces start operating. When the temperature is further lowered the attractive forces draw the molecules together to form a liquid.

Generally, it has been observed that, every gas has a specific temperature above which gas cannot be converted into liquid even at high pressure. The highest temperature, at which a gas can exist as a liquid, is called its **Critical Temperature**. While the pressure which is required at critical temperature to convert a gas into liquid is known as **Critical Pressure**. However, critical volume depends upon critical temperature and pressure. At critical temperature and pressure, the volume of one mole of a gas is called **Critical Volume**.

It is important to note that critical temperature and critical pressure are the parameters which provide us the information about the liquefaction of gases. Each gas has the specific values of critical temperatures and pressures as mentioned in table 4.3. By maintaining these parameters gases are liquefied. There are various methods for liquefaction of gas.

Table 4.3: Critical temperature and critical pressure of some common gases.

Gas	Critical Temperature (°C)	Critical Pressure (atm)
Oxygen , O_2	-118.75	49.7
Ammonia, NH_3	132.44	111.5
Nitrogen, N_2	-147.06	33.5
Argon, Ar	-122.26	48
Freon-12, CCl_2F_2	111.54	39.6

4.9.1 Joule-Thomson effect

When a compressed gas is allowed to expand from a region of higher pressure to a lower pressure through a nozzle causes a fall in temperature. This is known as Joule Thomson effect. The fall in temperature is proportional to the pressure difference between compressed and expanded gas. The molecule of gas in low pressure area absorb energy from environment to reduce their attractive forces. The most common example of this effe ct is applied to liquefy Nitrogen and carbon dioxide (dry ice).

Do You Know?

Liquid natural gas (LNG) and liquid petroleum gas (LPG) are used as a fuel in automobile.
LPG and LNG are chemically different since LNG consists mainly of methane gas.

4.9.2 Linde's Method of Liquefaction of Gases

Hampson-Linde (1842-1934) introduced this method based on the principle of Joule-Thomson effect for the liquefaction of gases. According to this method, "when a gas is allowed to expand in a adiabatic closed system from a region of higher pressure to region of extremely lower pressure, the liquefaction of air or any gas takes place.

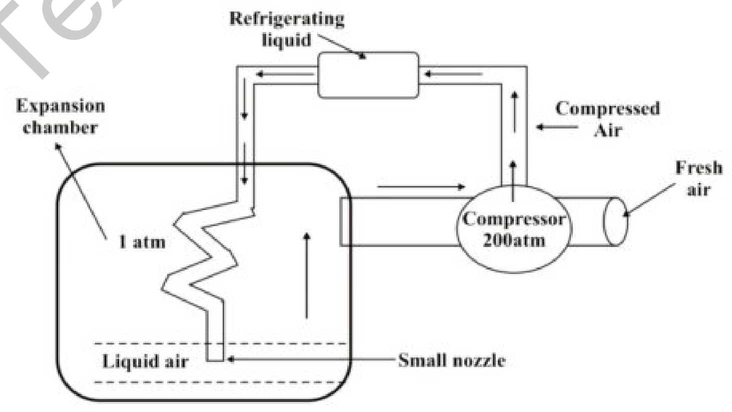


Fig 4.10: Liquification of gas by Linde's method

In this method, first of all fresh air or any other gas is compressed to around 200 atmospheres in compressor. Then compressed air is allowed to pass through pipe which contains water. Here moisture present in air is condensed as water and removed. In this part, heat of compressed air is absorbed by the water and dry air moved through copper spiral coil having jet at the end. When dry air passes through it, expansion occurs pressure is reduced to one atmosphere. This causes the decrease in temperature of air. As a result, cooled air is formed in the expansion chamber which moves up and cools the incoming compressed air. Then uncondensed air from expansion chamber returns to the compression pump for recycling. This process of cooling and compression followed by expansion is repeated many times. Thus, liquid air is collected at the bottom of the expansion chamber and removed.

4.9.3 Liquid air and its uses

Ordinary air on compression and cooling to extremely low temperature becomes liquefies and known as liquid air. It is mainly used for the cooling purpose. It absorbs the heat very rapidly and helps to convert the gases (filled in containers) into liquids and liquids into the solids. It is the industrial source of various gases such as oxygen, argon, nitrogen etc through a process named as air separation. These liquefied gases have various applications in various fields and common life. Some important uses of liquefied gases are given below:

(i) Liquefied natural gas (L.N.G) is used as a fuel.
(ii) Liquefied petroleum gas (L.P.G) is a mixture of butane and other hydrocarbons, used as a fuel for portable heaters, gas cooktops and ovens. It is also used as fuel for engines.
(iii) Ammonia and liquid sulphur dioxide are used as refrigerants.
(iv) Liquid air is an important source of oxygen in rockets.
(v) Compressed oxygen is used for wielding purpose.
(vi) Nitrogen is used in medical science.

4.10 FOURTH STATE OF MATTER: PLASMA

Except solid, liquid and gas, the fourth state of matter is called "**Plasma**". This word is taken from the Greek "**plassein**" which means moldable substance. This state of matter was identified by William Crooks in 1879 while he was working on "Discharge Tube Experiment". He called it "radiant matter". But later on, in the 1920s, Irving Langmuir used the word "**Plasma**". Before solids, liquids and gases, plasma was existed in the universe. More than 99% of the visible universe is composed of plasma. Plasma occurs naturally in the sun, stars, nebulas and ionosphere, but rarely on our Earth. The shining of stars and the sun is due to the plasma. On Earth it is found in lightning bolts, flames, welding arcs, auroras and fluorescent lights.

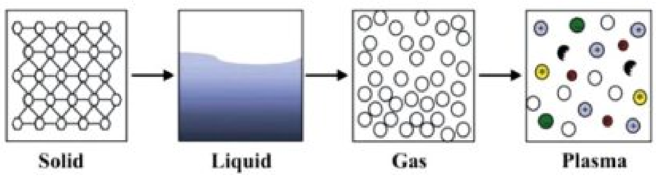


Fig 4.11: solid, liquid, gas and plasma states of matter.

When a gas is heated at high temperature ($10^4 – 10^5$ K), most of the molecules are ionized along with formation of cloud of electrons. Despite high temperature, there still exist some un-ionized molecules and atoms. **"The mixture of positive ions, electrons and un-ionized molecules and atoms is known as plasma"**.

Natural plasma exists only at high temperature or low temperature vacuums. Natural plasma does not break down or react rapidly but it is extremely hot (over 20000°C). Its energy is so high that it can vaporize any matter. Plasma as a whole is neutral because it contains equal number of negatively charged electrons and positively charged ions. Plasma contains significant number of charged particles which affect its properties. The presence of charged particles enables it to respond to electrical and magnetic fields. Like gases, plasma has neither definite shape nor volume. Plasma shows characteristic glow in discharge tube depending upon the nature of the gas. For example hydrogen exhibits green, oxygen red and nitrogen purple glows.

Uses of plasma:

Plasma has enormous uses in various fields which are given below:

(i) When light of fluorescent bulb or neon signs is turned on, gas is excited and creates glowing plasma which lightens the surroundings.

(ii) It is used in television and computer chips.

(iii) Plasma is used in rocket propulsion, cleaning the environment, destroying biological hazards and healing wounds.

(iv) It is used in the processing of semiconductors and sterilization of some medical products.

(v) It is used in lasers, lamps, pulsed power switches and diamond coated films.

Society, Technology and Science

Deep sea diving is a dangerous task but with SCUBA (Self-Contained Underwater Breathing Apparatus) tank it becomes significantly safe and wonderful for experience. This breathing tank mainly contains oxygen along with nitrogen or helium and a small amount of carbon dioxide and it makes it quite possible to dive much deeper than 60 feet.

When anyone breathes through the SCUBA, the pressure of oxygen in lungs must be equal to the pressure exerted on the body. At the surface of water the partial pressure of oxygen in air is 0.21 atm. As we go down in the sea the partial pressure of oxygen increases continuously. At the depth of 33 feet, the partial pressure of oxygen becomes double. Therefore, SCUBA tank is maintained according to the required partial pressure of oxygen.

Activity

In this activity, you can validate Charles' law at your home. Inflate two balloons up to the same size (volume) and then put either of them in your fridge and the other one in your room in normal ambient temperature for 1 hour. Can you guess what will happen an hour later?

The balloon that is kept inside the fridge will shrink as the air inside follows Charles law that at a constant pressure inside fridge, the decrease in the temperature of air (within balloon) causes the air inside to contract and eventually the balloon shrinks. Thus, it appears smaller in size as compared to the one kept in ambient temperature.

SUMMARY with Key Terms

- **Gas** is the state of a matter which has neither definite shape nor definite volume. Gas molecules are far away from each other and have weak attractive forces.
- **Kinetic Molecular Theory of gases** deals with the movement, arrangement, attractive force and pressure of gas particles (atoms or molecules).
- **Mean free path** is the average distance covered by a gas molecule before their collision.
- **Pressure of a gas** is the amount of force that applied to the surface of an object per unit area. The SI unit of pressure is pascal (N/m^2), however it is also measured in atmosphere and torr.
- **Atmospheric pressure** nearby sea level is taken as 1 atmosphere. It is measured by barometer. The pressure of air is lower than 1 atmosphere in hill sides.
- **Boyle's Law** describes the relationship between volume and pressure of an ideal gas and state as the volume of the given mass of a gas is inversely proportional to its pressure keeping the temperature constant.
- **Charles's Law** describes the relationship between volume and absolute temperature of an ideal gas and state as the volume of the given mass of a gas is directly proportional to its temperature keeping the pressure constant.
- **Absolute zero** is the minimum possible temperature at which gas occupies a theoretical zero volume. Actually before reaching this temperature all gases become liquefy.
- **Avogadro's Law** states that equal volume of all gases contains equal number of moles and the volume of a gas is directly proportional to its number of moles keeping the temperature and pressure constant.
- **Molar Volume** is the volume of one mole of any gas at standard temperature (0^oC) and pressure (1 atm). This volume is 22.4 dm^3.
- **Ideal gas** is that which obey general gas equation at all temperature and pressure.

- **Real gas** is that which obey van der Waal equation. It reaches to ideal behavior at high temperature and low pressure.
- **Dalton law of partial pressure** tells that the total pressure of the mixture of two or more non reacting gases is equal to the sum of partial pressure of individual gases in a fixed volume enclosed container.
- **Graham's law of diffusion** tells that the rate of diffusion of a gas is inversely proportion to the square root of its density or molecular mass keeping the temperature and pressure constant.
- **Critical temperature** is the highest temperature at which a gas exists as liquid.
- **Critical pressure** is the pressure which is required at critical temperature to convert a gas into liquid.
- **Liquid air** Ordinary air on compression and cooling to extremely low temperature becomes liquefies and known as liquid air. It is mainly used for the cooling purpose.
- **Plasma** is the fourth state of matter. It is a mixture of ions and un ionized atoms and molecules.

EXERCISE

Multiple Choice Questions

1. Choose the correct answer

(i) According to Graham's Law of diffusion, the ratio of diffusion of H_2 and O_2 are respectively:
(a) 1:2 (b) 2:1 (c) 1:4 (d) 4:1

(ii) Collection of gas over water is an example of:
(a) Graham's law (b) Dalton's law
(c) Avogadro's law (d) Gaylussac law

(iii) The molar volume of oxygen gas is maximum at:
(a) 0°C and 1 atm (b) 0°C and 2 atm
(c) 25°C and 1 atm (d) 25°C and 2 atm

(iv) The diffusion rate of C_3H_8 and CO_2 are same because:
(a) Both are poly atomic gases (b) Both are denser than air
(c) Both have same molar mass (d) Both contains carbon atoms

(v) The volume of gas would be theoretically zero at:
(a) 0°C (b) 0 K (c) 273 K (d) 273°C

(vi) Real gas reaches the ideal behavior at:
(a) Low temperature and low pressure (b) High temperature and high pressure
(c) Low temperature and high pressure (d) High temperature and low pressure

(vii) Which one of the following statement is incorrect about the gas molecules?
(a) They have large spaces
(b) They possess kinetic energy
(c) Their collision is elastic
(d) Their molar mass depends upon temperature

(viii) If the Kelvin temperature of ideal gas is increase to double and pressure is reduce to one half, the volume of gas will:
(a) Remains same
(b) Double
(c) Reduced to half
(d) Four time

(ix) The molar volume of oxygen gas is 22.4 dm^3 at:
(a) 0°C and 1 atm
(b) 25°C and 0.5 atm
(c) 0 K and 1 atm
(d) 25 K and 0.5 atm

(x) Under similar condition CH_4 gas diffuses........ times faster than SO_2 gas:
(a) 1.5 time
(b) 2 times
(c) 4 times
(d) 16 times

Short Questions

1. State the following gas laws.
(i) Avogadro's law
(ii) Charles law
2. State main postulates of kinetic molecular theory of gas.
3. Explain the following.
(i) Pressure and its various units
(ii) Absolute zero
4. What is liquid air? Mention its three uses.
5. What is plasma? Give its significance in daily life.

Descriptive Questions

1. State and explain Dalton's law of partial pressure. Give practical applications of Dalton's law.
2. Derive general gas equation. Also deduce the value of R in atm dm^3/mol.K and J/mol.K.
3. How an ideal gas is differentiated from real gas? What are the causes of deviation of real gas from ideal behavior? Explain this deviation at low temperature and high pressure.
4. What is meant by diffusion and effusion? Explain Graham's law of diffusion.

Numerical Questions

1. Gorakh hill station is coldest area of Sindh province and has 83Kpa barometer pressure. What will be the pressure of this area in psi and atm units?

2. Calculate the volume occupied by 8g of methane gas at 40°C and 842 torr pressure? [Ans: V = 11.58 dm^3]

3. At 35°C, oxygen gas in a cylinder has 456 cm^3 volume and 0.85 atm pressure. Calculate the pressure when this oxygen gas is transferred to 10 dm^3 cylinder and cooled to 20°C. [Ans: 0.037 atm]

4. Compare the rates of diffusion of the following pairs of gases:
(1) H_2 and D_2 (2) He and SO_2 (3) SF_6 and SO_2 [Ans: 1.41/1,4/1, 1.51/1]

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

6. A 500cm^3 vessel contains H_2 gas at 400 torr pressure and another 1dm^3 vessel contains O_2 gas at 600 torr pressure. If under the similar condition of temperature these gases are transferred to 2dm^3 empty vessel, calculate the pressure of the mixture of gases in new vessel.

7. If 16 cm^3 of hydrogen effuses in 30 sec, from a porous material, what volume of SO_2 will effuse in the same time (30 sec.) under similar conditions?

8. 40dm^3 of hydrogen gas was collected over water at 831 torr 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.