

GASEOUS STATE

Preface

This chapter deals with a particular physical state of substances called gaseous state. After successful completion of the chapter, you will be able to know about various parameters of gas, gas law & gas equation. Dalton's law, Avagadro's & Graham's law, K.T.G., Van der Waal's equation and critical constants

This book consists of theoretical & practical explanations of all the concepts involved in the chapter. Each article is followed by a ladder of illustration. At the end of the theory part, there are miscellaneous solved examples which involve the application of multiple concepts of this chapter.

Students are advised to go through all these solved examples in order to develop better understanding of the chapter and to have better grasping level in the class.

Total number of questions in Gaseous State are :		
(i) In chapter Examples	18
(ii) Solved Examples	20
Total no. of questions	38

1. GASEOUS STATE ::

When the molecular forces of attraction between the particles of a matter are minimum, the particles exist in a state known as gaseous state or a state of matter in which molecules are far away from each other and free to move in available space is called gaseous state.

1.1 Properties of Gases

- (a) They do not have definite shape and volume.
- (b) They can occupy whole space open to them.
- (c) Gases have unlimited dispersibility and high compressibility.
- (d) They have very low densities because of negligible intermolecular forces.
- (e) Gases exert pressure on the wall of the container with perfectly elastic collisions.
- (f) They diffuse rapidly through each other to form homogeneous mixture against the electric, magnetic and gravitational field.

1.2 Parameters of Gases

The characteristics of gases are described in terms of four measurable parameters and it is also called as measurable properties of gases which are

- (a) Mass
- (b) Volume
- (c) Pressure and
- (d) Temperature

(a) **Mass (m)**- The mass of a gas is denoted by 'm' which is related to the no of moles 'n'. Therefore,

$$n \text{ (no of moles)} = \frac{m(\text{mass in grams})}{M(\text{Molar mass})}$$

$$\text{so, } m = n \times M$$

(b) Volume V-

- (i) Gases occupy whole space available to them. The volume occupied by a gas is simply the volume of container in which it is filled.
- (ii) The volume of a gas is denoted by 'V' and it is measured in units of litre or cubic metre (m^3) or cm^3 or dm^3 .
- (iii) $1 \text{ litre} = 1 \text{ dm}^3 = 1000 \text{ cm}^3 = 1000 \text{ ml}$

(c) Pressure (P)-

- (i) It is force acting per unit area. A confined gas exerts uniform pressure on the walls of its container in all directions.

- (ii) It is denoted by 'P' and specified in pascal (P_a).
- (iii) Other units of pressure are atm, cm hg, mmHg, N/m^2 , bar, torr.
- (iv) $1 \text{ atm} = 76 \text{ cm Hg} = 760 \text{ mm Hg}$
 $\text{Hg} = 1.013 \times 10^5 \text{ N/m}^2 = 1.013 \times 10^5$
 $P_a = 1.013 \text{ bar} = 760 \text{ torr}$.

$$\begin{aligned} \text{(v) } P \text{ (Pressure)} &= \frac{F(\text{Force})}{A(\text{Area})} \\ &= \frac{\text{Mass} \times \text{Acceleration}}{\text{Area}} \end{aligned}$$

(vi) Pressure exerted by a gas is due to kinetic energy (K.E. = $\frac{1}{2} mv^2$) of the gas molecules.

(vii) K.E. of the gas molecules increases, as the temperature is increased so, pressure of a gas is directly proportional to temperature.
 $P \propto T$

(d) Temperature (T)-

- (i) The temperature of a gas is denoted by 'T' and it is measured in the unit of kelvin (K).
- (ii) Other units of temperature are, $^{\circ}\text{C}$, $^{\circ}\text{F}$, $^{\circ}\text{R}$.
- (iii) $K = ^{\circ}\text{C} + 273.15$

$$\text{(iv) } \frac{x^{\circ}\text{C}}{5} = \frac{(y^{\circ}\text{F} - 32)}{9}$$

2. Gas Laws ::

The certain laws which relate the four parameters are called gas laws.

2.1 Boyle's Law

(a) It states that at constant temperature, the volume of a given mass of a gas is inversely proportional to the pressure.

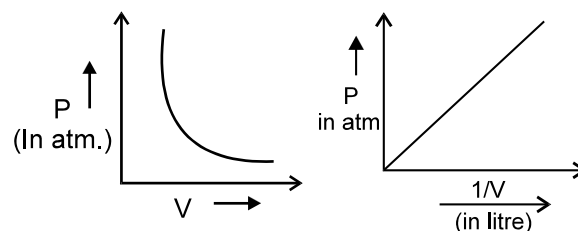
(b) Mathematically -

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

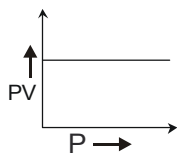
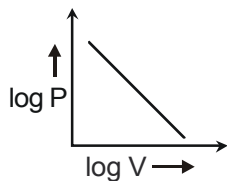
$$\text{or } PV = K \text{ or } P_1V_1 = P_2V_2$$

(c) Graphical representations

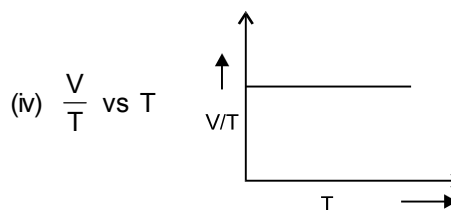
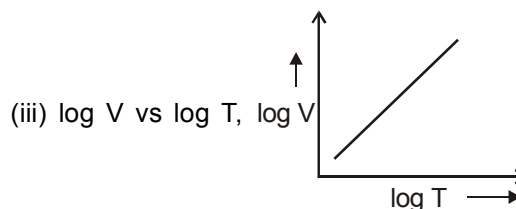
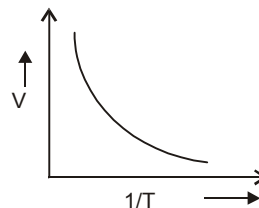
- (i) P vs V
- (ii) P vs $\frac{1}{V}$



(iii) log P vs, log V (iv) PV vs P



(ii) V vs $\frac{1}{T}$ or T vs $\frac{1}{V}$



2.2 Charle's law

(a) This law states that at constant pressure, the volume of a given mass of a gas is directly proportional to its absolute temperature.

(Absolute temperature = $^{\circ}\text{C} + 273.15$)

(b) Mathematically - $V \propto T$ (at constant pressure)

V = volume of gas

T = Absolute temperature

$V = KT$ or

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

Hence, if the volume of a gas of mass is V_1 at temperature T_1 changes to V_2 at T_2 , pressure remaining constant,

$$\text{then } \frac{V_1}{T_1} = \frac{V_2}{T_2} = \text{constant}$$

or $\log V - \log T = \text{constant}$

(c) For each degree change of temperature the volume of sample of a gas changes by the fraction of $\frac{1}{273.15}$ of its volume at 0°C

$$\text{so } V_t = V_0 \left[\frac{273.15 + t}{273.15} \right]$$

This equation is called Charles-gay-lussac equation.

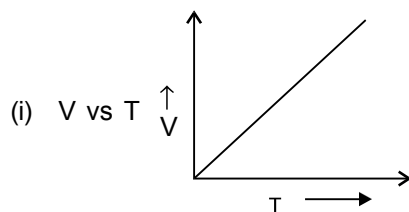
where

V_t = volume of gas at temperature $t^{\circ}\text{C}$

V_0 = volume of gas at 0°C temperature

t = temperature in $^{\circ}\text{C}$.

(d) Graphical representations -



2.3 Gay-Lussac's Law or Amonton's Law

(a) It states that at constant volume, the pressure of a given mass of a gas is directly proportional to its absolute temperature.

(b) Mathematically - $P \propto T$ (at constant volume)

where P = pressure of gas

T = Absolute temperature

$P = KT$

$$\text{or } \frac{P}{T} = K$$

Hence, if the pressure of a gas is P_1 at temperature T_1 changes to P_2 at T_2 , volume remaining constant.

$$\text{then } \frac{P_1}{T_1} = \frac{P_2}{T_2} = \text{constant}$$

$\log P - \log T = \text{constant}$.

$$(c) P_t = P_o \left(1 + \frac{t}{273.15} \right)$$

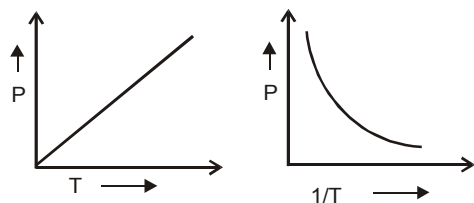
where P_t = Pressure of gas at $t^{\circ}\text{C}$

P_o = Pressure of gas at 0°C

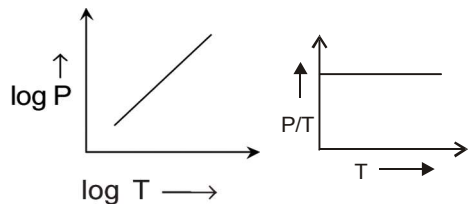
t = Temperature in $^{\circ}\text{C}$.

(d) Graphical representation -

(i) P vs T (ii) P vs $\frac{1}{T}$ or T vs $\frac{1}{P}$



(iii) log P vs log T (iv) $\frac{P}{T}$ vs $V_s T$



Examples based on Gas Laws

Ex.1 A gas occupies a volume of 2.4 litres at a pressure of 740 mm of mercury. Keeping the temperature constant, calculate its volume at standard pressure -

- (A) 2.4 litres (B) 2.34 litres
(C) 2.5 litres (D) None of these

(Ans. B)

Sol. Initial volume (V_1) = 2.4L, Initial pressure (P_1) = 740 mm. Final volume (V_2) = ?
Final pressure (P_2) = 760 mm.

$$P_1 V_1 = P_2 V_2 ; \quad \therefore V_2 = \frac{740 \times 2.4}{760}$$

$$= 2.34 \text{ litres.}$$

Ex.2 A gas occupies a volume of 580 ml at 17 °C. It is heated to 100 °C at constant pressure. Calculate the volume of the gas -

- (A) 746 ml. (B) 760 ml.
(C) 773 ml. (D) 780 ml.

(Ans.A)

Sol. Initial volume (V_1) = 580 ml,
 $T_1 = 17 + 273 = 290 \text{ K}$
Final Volume (V_2) = ?
 $T_2 = 100 + 273 = 373 \text{ K}$

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

$$\therefore V_2 = T_2 \times \frac{V_1}{T_1} = 373 \times \frac{580}{290} = 746 \text{ ml.}$$

Ex.3 A gas occupies 3 litres at 32 °C and one atmospheric pressure. What volume will it occupy if the temperature is changed to 18 °C, the pressure remaining constant -

- (A) 2.91 litres (B) 2.86 litres
(C) 2.30 litres (D) None of these

(Ans. B)

Sol. According to Charle's law,

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

$$\frac{3\text{L}}{(273 + 32)\text{K}} = \frac{V_2}{(273 + 18)\text{K}}$$

$$\text{or } V_2 = \frac{3 \times 291}{305} = 2.86 \text{ litres.}$$

2.4 Ideal gas equation

- (a) It correlate all the four parameters of a gas.
(b) It is the combination of Boyle's and charle's law.

(c) $PV = nRT$

$$PV = \frac{m}{M} RT$$

The equation is called as ideal gas equation.

Where n = no of moles of the gas
 m = mass of the gas
 M = Mol. wt. of the gas
 R = Molar gas constant.

(d) For 1 mole of the gas $n = 1$

$$PV = RT$$

$$\text{So } \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} = R \text{ or } \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Where P_1, V_1, T_1 are the initial pressure, volume and temperature and P_2, V_2, T_2 are final.

The above equation is called as ideal gas equation.

(e) The unit of R is the unit of work or energy per degree per mole as -

$$R = \frac{PV}{nT} = \frac{\text{Pressure} \times \text{volume}}{\text{mole} \times \text{Temperature}}$$

$$R = \frac{\frac{\text{force}}{\text{Area}} \times \text{volume}}{\text{mole} \times \text{Temperature}} = \frac{\text{force} \times \text{length}}{\text{mole} \times \text{temperature}}$$

$$R = \frac{\text{work(energy)}}{\text{mole} \times \text{temperature}}$$

(f) Numerical values of R in different units -

- (i) $R = 0.0821 \text{ litre atm. deg.}^{-1} \text{ mole}^{-1}$
- (ii) $R = 62.4 \text{ litres mm. deg.}^{-1} \text{ mole}^{-1}$
- (iii) $R = 8.314 \times 10^7 \text{ ergs deg.}^{-1} \text{ mole}^{-1}$
- (iv) $R = 82.05 \text{ C.C.atm. deg.}^{-1} \text{ mole}^{-1}$
- (v) $R = 2 \text{ cal. deg.}^{-1} \text{ mole}^{-1}$
- (vi) $R = 8.314 \text{ J K}^{-1} \text{ mole}^{-1}$

Examples based on

The Ideal gas equation

Ex.4 How many litres would 5 moles of H_2 occupy at 25°C and 2 atm pressure -

- (A) 61.20 litres
- (B) 61.09 litres
- (C) 30.50 litres
- (D) 30.60 litres

(Ans.B)

Sol. $PV = nRT$

$$= \frac{(5)(0.082)(298)}{2}$$

$$= 61.09 \text{ litres.}$$

Ex.5 What is the pressure of a mixture of 1 g of hydrogen and 1.4 g of nitrogen stored in a 5 litre vessel at 127°C -

- (A) 5.50 atm.
- (B) 3.61 atm.
- (C) 4.40 atm.
- (D) 4.50 atm.

(Ans.B)

Sol. moles of $\text{H}_2 = \frac{1}{2} = 0.5$

moles of $\text{N}_2 = \frac{1.4}{28} = 0.05$

\therefore Total number of moles of gas (n) = 0.55

Using $PV = nRT$

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

$$= \frac{0.55 \times 0.0821 \times 400}{5}$$

$$= 3.61 \text{ atm.}$$

3. Dalton's Law of Partial Pressure

(a) According to this law, when two or more than two chemically inert gases are kept in a closed container, the total pressure exerted

by the gaseous mixture is equal to the sum of the partial pressures of individual gases - i.e. -

$$P = P_1 + P_2 + P_3 + \dots + P_n$$

(b) Let n_1 & n_2 be the no. of moles of two inert gases A and B which is filled in a container of volume 'V' at temperature T. So the total pressure of container 'P' may be calculated as -

$$PV = (n_1 + n_2) RT \dots\dots (i)$$

Partial pressure of individual gas calculates at -

$$P_A V = n_1 RT \dots (ii)$$

$$P_B V = n_2 RT \dots (iii)$$

On the addition of eq. (ii) & (iii) we get -

$$(P_A + P_B) V = (n_1 + n_2) RT \dots (iv)$$

On the comparison of eq. (i) & (iv)

$$P = P_A + P_B$$

Dividing by equation (ii) by (i), we get

$$\frac{P_A}{P} = \frac{n_1}{n_1 + n_2} = x_A$$

$$P_A = x_A \times P$$

where x_A = mole fraction of 'A'

Similarly dividing (iii) by (i), we get

$$P_B = x_B \times P$$

so

Partial pressure of a component = mole fraction \times total pressure

3.1 Applications of Dalton's Law of Partial pressure

(a) mole fraction of a gas in a mixture of gas

$$= \frac{\text{partial pressure of gas}}{\text{Total pressure of the mixture of gas}}$$

(b) % of a gas in a mixture

$$= \frac{\text{partial pressure of gas}}{\text{Total pressure}} \times 100$$

(c) Pressure of dry gas which is collected over the water is-

$$P_{\text{Total}} = P_{\text{moist air}} = P_{\text{dry gas}} + P_{\text{water vapour}}$$

(Note : $P_{\text{water vapour}}$ is called aqueous tension)

$$\text{so } P_{\text{dry gas}} = P_{\text{Total}} - P_{\text{Water vapour}}$$

(Note: Aqueous tension is directly proportional to absolute temperature)

3.2 Limitations of Dalton's law of partial pressure

- (a) It is applicable only inert gases like N_2 and O_2 , N_2 and Cl_2 etc.
(b) It is not applicable for chemically reactive gases like H_2 and Cl_2 , CO and Cl_2 etc.

Examples based on Dalton's law of partial pressure

Ex.6 0.333 grams of alcohol displaced in a Victor Meyer apparatus 171 c.c. of air measured over water at $15^\circ C$. The barometric pressure was 773 torr. Calculate the molecular weight of alcohol-(Aqueous tension at $15^\circ C = 13$ torr.)

- (A) 33.34 g/ mol. (B) 28.80 g/ mol.
(C) 46.0 g/ mol (D) 13.0 g/mol.

(Ans. C)

Sol. $P_{\text{dry gas}} = 773 - 13 = 760$ torr

$$= \frac{760}{760} = 1 \text{ atm}$$

$$\therefore PV = nRT$$

$$1 \times \frac{171}{1000} = \frac{0.333}{\text{M.wt.}} \times 0.0821 \times 288$$

$$M = 46 \text{ g. per mol.}$$

Ex.7 Atmospheric air contains 20% O_2 and 80% N_2 by volume and exerts a pressure of 760 mm. Calculate the partial pressure of each gas -

- (A) 152 mm, 608 mm
(B) 608 mm, 152 mm
(C) 760 mm both
(D) None of these

(Ans. A)

Sol. Partial Pressure

$$= \text{Mole fraction} \times \text{Total pressure}$$

$$= \text{Vol. fraction} \times \text{Total pressure}$$

$$\therefore P_{O_2} = 0.2 \times 760 = 152 \text{ mm}$$

$$\therefore P_{N_2} = 0.8 \times 760 = 608 \text{ mm}$$

Ex.8 Equal weights of ethane and hydrogen are mixed in an empty container at $25^\circ C$. Determine the fraction of the total pressure exerted by hydrogen -

- (A) 16/30 (B) 15/16
(C) 30/16 (D) 16/15

(Ans. B)

Sol. Let the wt. of ethane and $H_2 = w$ gm

$$\therefore \text{No. of moles of } C_2H_6 = \frac{w}{30}$$

$$\text{No. of moles of } H_2 = \frac{w}{2}$$

$$\text{Total no. of moles} = \frac{w}{30} + \frac{w}{2} = \frac{16w}{30}$$

$$\text{Mole fraction of } H_2 = \frac{w}{2} \times \frac{30}{16w} = \frac{15}{16}$$

Ex.9 2.8 g of N_2 , 2.8 g CO , 4.4 g CO_2 are found to exert a pressure of 700 Torr. Find partial pressure of N_2 gas in the mixture -

- (A) 280.8 Torr (B) 233.3 Torr
(C) 300 Torr (D) None of these

(Ans. B)

Sol. $p_{N_2} = \text{its mole-fraction} \times P_{\text{total}}$

$$= \frac{2.8 / 28}{\frac{2.8}{28} + \frac{2.8}{44} + \frac{4.4}{44}} \times 700$$

$$= \frac{0.1}{0.3} \times 700$$

$$= 233.3 \text{ Torr.}$$

4. Avagadro's Law ::::

(a) According to this law under the same condition of temperature and pressure, equal volumes of all gas contains equal no. of molecules.

$$V \propto n \text{ (At constant temperature \& pressure)}$$

Where V = volume

n = no of molecules

(b) Molar Volume or gram molecular volume - 22.4 litres or 22400 ml of every gas at NTP is the volume occupied by its one gram mole and it is called molar volume or gram molecular volume.

(c) The mole Concept - Mole is defined as the total amount of substance that contains as many basic units as there are atoms in 12 g of the isotopes of carbon -12. Thus a mole will be defined as the Avogadro no of particles which is equal to 6.023×10^{23} .

(d) Loschmidt number - It the no of molecules present in the volume of a gas at S.T.P. Its value is 2.617×10^{19} per c.c.

5. Graham's Law of Diffusion or Effusion ::

Diffusion: It is the ability of a gas to mix spontaneously and to form a homogenous mixture is known as diffusion.

Effusion: It is a process in which a gas is allowed to escape under pressure through a fine orifice from closed container.

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- (a) This law was proposed by Thomas Graham.
 (b) According to this law, at constant temperature and pressure, the rate of diffusion or effusion of a gas is inversely proportional to the square root of its density. Thus and directly proportional to its pressure.

$$\text{Rate of diffusion } (r) \propto \frac{P}{\sqrt{d}}$$

$$\frac{r_1}{r_2} = \frac{p_1}{p_2} \sqrt{\frac{d_2}{d_1}}$$

It pressure is constant

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

where r_1 & r_2 are rates of diffusion of two gases and d_1 & d_2 are densities.

$\therefore 2 \times \text{vapour density} = \text{Molecular mass}$

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

where M_1 & M_2 are the molecular masses of two gases.

(c) Rate of diffusion = $\frac{\text{Volume of gas diffused}}{\text{Time taken for diffusion}}$

5.1 Applications of graham's law of diffusion

- (a) To the Detection of marsh gas in mines.
 (b) Separation of isotopes
 (c) Determination of density and molecular mass of gases.

Examples based on

Graham's Law of diffusion

Ex.10 Calculate the relative rates of diffusion of $^{235}\text{UF}_6$ and $^{238}\text{UF}_6$ in the gaseous state (At. mass F = 19).

- (A) 1.0043 : 1.0000 (B) 1.0000 : 1.0043
 (C) 1.349 : 1.352 (D) 1.352 : 1.349

(Ans. A)

Sol. Mol. mass $^{235}\text{UF}_6 = 235 + 6 \times 19 = 349$
 Mol. mass $^{238}\text{UF}_6 = 238 + 6 \times 19 = 352$

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{352}{349}} = 1.0043$$

$$r_1 : r_2 :: 1.0043 : 1.0000$$

Ex.11 The densities of CH_4 and O_2 are in the ratio 1 : 2. Calculate the ratio of rates of diffusion of oxygen and methane.

- (A) 1.414 : 1 (B) 1: 1.414
 (C) 1.614 : 1 (D) 1.614: 1

(Ans.B)

Sol. $\frac{r_{\text{O}_2}}{r_{\text{CH}_4}} = \sqrt{\left(\frac{d_{\text{CH}_4}}{d_{\text{O}_2}}\right)} = \sqrt{\left(\frac{8}{16}\right)} = \sqrt{\left(\frac{1}{2}\right)}$

$$= 1 : 1.414$$

Ex.12 Calculate the molecular weight of a gas which diffuses through a porous plug at 1/6th of the speed of hydrogen under same conditions.

- (A) 36 (B) 76
 (C) 72 (D) 63

(Ans.C)

Sol. $\sqrt{\left(\frac{M_{\text{gas}}}{M_{\text{H}_2}}\right)} = \frac{\text{Rate of diffusion of H}_2}{\text{Rate of diffusion of gas}}$

$$= \sqrt{\left(\frac{M_{\text{gas}}}{2}\right)} = \left(\frac{1}{1/6}\right)$$

$$\text{or } M_{\text{gas}} = 2 \times 36 = 72$$

Ex.13 The vapour density of gas A is thrice that of the gas B. If the molecular weight of B is M, then calculate the molecular weight of A.

- (1) M (2) 3 M
 (3) M/3 (4) None of these

(Ans. B)

Sol. $\frac{VD_A}{VD_B} = \frac{M_A}{M_B} \therefore \frac{3}{1} = \frac{M_A}{M_B}$, So Mol. wt of A (M_A) = 3 M.

6. Kinetic Theory of Gases ::

- (a) It was a commonly given for ideal gases.
- (b) This theory was proposed by Bernoulli and next to developed by Clausius, Maxwell, Kroning and Boltzman.

6.1 Assumptions of Kinetic Theory of gases

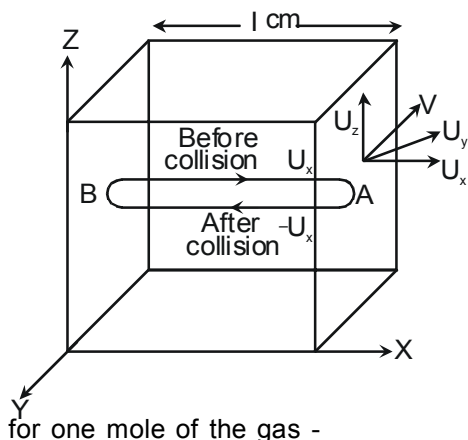
- (a) All the gases are made up of molecules moving randomly in all the directions.
- (b) The volume of the individual molecule is negligible as compared to the total volume of the gas.
- (c) Pressure exerted by gas is due to collision of gas molecules with the wall of the container so.
Pressure \propto No of collisions per unit time per unit area by the molecules on the wall of the container
- (d) All collisions between two molecules or between a molecule and a wall are perfectly elastic.
- (e) All the molecules obey Newton's laws of motion.
- (f) Gas molecules neither attract nor repel each other.
- (g) Kinetic energy of gas molecules depend upon the absolute temperature so.
Kinetic energy \propto absolute temperature
- (h) The force of gravity has no effect on the speed of gas molecules.

6.2 Calculation of Kinetic energy

According to postulates of kinetic gas equation-

$$PV = \frac{1}{3} mnv^2$$

- where P = pressure of the gas
 V = volume of the gas
 n = no of molecules present in the given amount of gas
 v = Root mean square speed.



$$PV = RT \quad \text{and } n = N$$

$$\frac{1}{3} mNv^2 = RT$$

$$\text{or } \frac{2}{3} \cdot \frac{1}{2} mNv^2 = RT$$

$$(\because \frac{1}{2} mNv^2 = \text{K.E. per mole})$$

$$\therefore \frac{2}{3} \text{K.E.} = RT$$

$$\text{K.E.} = \frac{3RT}{2} \Rightarrow \text{K.E. per mole / Avg. K.E. per mole.}$$

$$\therefore \text{ Boltzmann constant } K = \frac{R}{N} \\ = 1.38 \times 10^{-16} \text{ ergs } K^{-1} \text{ molecules}^{-1}$$

$$\text{K.E.} = \frac{3KT}{2} \Rightarrow \text{K.E. per molecule / Avg.}$$

$$\text{K.E. per molecule}$$

6.3 Velocities related to gaseous state

(A) Root mean square velocity (RMS velocity) -

Square root of the mean of the squares of the speed of all the molecules present in the given sample of the gas is called root mean square speed.

$$u = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3P}{d}}$$

(b) Average Speed -

It is the arithmetic mean of the various speeds of the molecules

$$\text{Average speed } (\bar{v}) = \sqrt{\frac{8RT}{\pi M}}$$

$$\text{Average speed} = .9213 \times \text{RMS speed}$$

$$\text{RMS speed} = 1.085 \times \text{Average speed}$$

(c) Most Probable Speed-

Speed possessed by maximum number of molecules of a gas at a given temperature.

$$\text{Most probable speed } (\alpha) = \sqrt{\frac{2RT}{M}}$$

$$\text{MPS} = .816 \times \text{RMS} ; \text{RMS} = 1.224 \text{ MPS}$$

$$\text{MPS} : \text{A.V. speed} : \text{RMS} = 1 : 1.128 : 1.224$$

KTG, K.E. & Different velocities of gaseous molecules

Ex.14 3 moles of a gas are present in a vessel at a temperature of 27 °C. Calculate the value of gas constant (R) in terms of kinetic energy of the molecules of gas.

- (A) 7.4×10^{-4} KE per degree kelvin.
 (B) 9.4×10^{-5} KE per degree kelvin.
 (C) 4.5×10^{-6} KE per degree kelvin.
 (D) None of these **(Ans. A)**

Sol. K.E. for 1 mole = $\frac{3}{2} RT$

$$\text{K.E. for 3 moles} = \frac{9}{2} RT. \quad \text{or} \quad R = \frac{9}{9T}$$

$$\text{KE} = \frac{2}{9(300)} \text{KE} = 7.4 \times 10^{-4} \text{ KE per degree kelvin.}$$

Ex.15 Calculate average kinetic energy, in joules, of the molecules in 8.0 g of methane at 27 °C -

- (A) 8169.75 Joules (B) 1869.75 Joules
 (C) 6189.57 Joules (D) 9186.57 Joules

(Ans.B)

Sol. $K = \frac{3}{2} n RT = \frac{3}{2} \times \frac{8}{16} \times 8.314 \times 300$
 $= 1869.75 \text{ Joules}$

Ex.16 A gaseous mixture contains 4 molecules with a velocity of 6 cm sec⁻¹, 5 molecules with a velocity of 2 cm sec⁻¹ and 10 molecules with a velocity of 3 cm sec⁻¹. What is the RMS velocity of the gas -

- (A) 2.5 cm sec⁻¹ (B) 1.9 cm sec⁻¹
 (C) 3.6 cm sec⁻¹ (D) 4.6 cm sec⁻¹

(Ans.C)

Sol. RMS velocity u

$$= \sqrt{\left(\frac{n_1 v_1^2 + n_2 v_2^2 + n_3 v_3^2 + \dots}{n} \right)}$$

$$= \sqrt{\left(\frac{4 \times (6)^2 + 5 \times (2)^2 + 10(3)^2}{19} \right)}$$

$$= \sqrt{\left(\frac{4 \times 36 + 5 \times 4 + 10 \times 9}{19} \right)} = \sqrt{\left(\frac{254}{19} \right)}$$

$$= 3.6 \text{ cm sec}^{-1}$$

Ex.17 Calculate the RMS velocity of a gas with average velocity 45475 cm/sec -

- (A) 45.475 cm/sec. (B) 92.131 cm/sec.
 (C) 49.330 cm/sec (D) 20.259 cm/sec.

(Ans. C)

Sol. Average velocity (v) = 0.9213 × RMS velocity (u)

$$\text{RMS velocity (u)} = \frac{C}{0.9213}$$

$$= \frac{45475}{0.9213} = 49.330 \text{ cm/sec.}$$

Ex.18 Calculate average velocity and RMS velocity for a group of six particles having speeds 11.2, 9.0, 8.3, 6.5, 3.7 and 1.8 m sec⁻¹ -

- (A) 6.75 ms⁻¹, 7.47 ms⁻¹
 (B) 7.47 ms⁻¹, 6.75 ms⁻¹
 (C) 7.65 ms⁻¹, 8.47 ms⁻¹
 (D) None of these

(Ans.A)

Sol. Average velocity (v)

$$= \frac{11.2 + 9.0 + 8.3 + 6.5 + 3.7 + 1.8}{6}$$

$$= \frac{40.5}{6} = 6.75 \text{ ms}^{-1}$$

$$\text{RMS velocity} = \frac{6.75}{0.9213} = 7.47 \text{ ms}^{-1}$$

7. Difference between real gas & ideal gas ::

S.N.	Ideal gas	Real gas
1.	Obeys gas law under all conditions of P and T.	Obeys only at low P and high T.
2.	Obeys ideal gas equation	Does not obeys ideal gas equation.
3.	Intermolecular interaction between gaseous molecules are negligible.	Intermolecular interaction between gaseous molecules not negligible.
4.	Volume of a particle is negligible as compared to total volume of the gas.	Volume of a particle is not negligible as compared to total volume of the gas.
5.	Exists only at high temperature and low pressure	Exists only at low temperature and high pressure.

7.1 Explanation of deviation

Kinetic theory of gases do not hold good at all condition mostly these two assumptions

- (a) The force of attraction between gaseous molecules are negligible
- (b) The volume occupied by the gaseous molecules is negligible compared to total volume of gas at high pressure. Both the assumptions do not hold good hence deviating from ideal gas.

8. SPECIFIC HEAT OF GASES ::

(a) Specific Heat - It is the amount of heat required to raise the temperature of one gm of the substance through one degree centigrade. It is usually expressed in "Calories".

(b) Calorie - A calorie is defined as the amount of heat required to raise the temperature of one gm of water through 1°C (more accurately from 15.5 °C to 16.5 °C.).

The heat so supplied is used up in increasing the internal energy of the molecule, i.e. in raising the temperature or in raising the average K.E. of the molecules.

(c) Specific Heat of a gas at constant volume (C_V) -

The amount of heat required to raise the temperature of one gm of the gas by 1 °C when the volume is kept constant and pressure is allowed to increase.

(d) Specific Heat of a gas at constant Pressure (C_P) -

The amount of heat required to raise the temperature of one gm of the gas by 1 °C when the pressure is kept constant and volume is allowed to increase.

$$C_V = c_v \times M \quad \text{and} \quad C_P = c_p \times M$$

Where C_V and C_P are molar specific heat or heat capacity per mole at constant volume and at constant pressure.

(e) Ratio for Molar Heats -

(i) For monoatomic gases $\frac{C_P}{C_V} = g$

$$= \frac{5.0}{3.0} = 1.66$$

∴ Molar heat capacity at constant volume (C_V)

$$= \text{K.E.} = \frac{3RT}{2}$$

at 1 °C \Rightarrow K.E = 3/2 R

$$\text{K.E.} = \frac{3}{2} \times 2 = 3.$$

Therefore $\Rightarrow C_V = 3$ calories.

& $C_P =$ Increase in K.E. $\left(\frac{3}{2}R\right) +$ work done

$$= \frac{3}{2} R + R = \frac{5}{2} R \quad (\text{Since } R = 2)$$

$$C_P = \frac{5}{2} \times 2 = 5 \text{ calories.}$$

(ii) For diatomic gases $\frac{C_P}{C_V} = \frac{5+x}{3+x}$

In case of diatomic gases x in many cases is 2.

$$\frac{C_P}{C_V} = \frac{7}{5} = 1.40$$

(iii) For triatomic gases $\frac{C_P}{C_V} = \frac{5+3}{3+3}$

$$= \frac{8}{6} = 1.33$$

9. VANDER WAAL'S EQUATION- A MODIFICATION IN GAS EQUATION ::

According to Vander Waal Maxwell has made two wrong assumptions in his kinetic theory of gases which lead to wrong results.

- (a) The forces of attractions between gaseous molecules are negligible under all conditions.
- (b) The volume occupied by the gaseous molecules is negligible as compared to total volume of gas.
- (c) Volume Correction - The observed volume is greater than ideal volume & correction term 'nb' has to be subtracted from observed volume in order to get ideal volume.

$$V_{\text{ideal}} = V - nb$$

Where 'n' is no. of moles of real gas & 'b' is constant which depends on nature of real gas.

10. CRITICAL CONSTANTS ::

- (a) **Critical temperature (T_c)** - The temperature above which gas cannot be liquified whatever pressure is applied is called critical temperature.

$$T_c = \frac{8a}{27bR}$$

- (b) **Critical pressure (P_c)** - The minimum pressure required to liquefy a gas at its critical temperature is called critical pressure.

$$P_c = \frac{a}{27b^2}$$

- (c) **Critical volume (V_c)** - The volume occupied by one mole of a gas at the critical temperature and critical pressure is called critical volume.

$$V_c = 3b$$

11. LIQUIFICATION OF GASES ::

- (a) On increasing the pressure gas can be easily liquified because intermolecular attraction between gaseous molecules increases.
- (b) On decreasing the temperature K.E. of gaseous molecules decrease That's why gas can be easily liquified.
- (c) Temperature of gas must be lower than critical temperature (T_c).

12. BOYLE'S TEMPERATURE ::

The temperature at which a real gas behaves like an ideal gas over an appreciable pressure range.

$$T_B = \frac{a}{Rb}$$

13. INVERSION TEMPERATURE ::

It is the temperature at which any real gas is expanded adiabatically and no effect of cooling or heating is observed.

$$T_i = \frac{2a}{Rb}$$

SOLVED EXAMPLE

Ex.1 A gas occupies 300 ml at 27°C and 730 mm pressure. What would be its volume at STP-

- (A) 162.2 ml (B) 262.2 ml
(C) 362.2 ml (D) 462.2 ml

Sol. (B)

Given at $T_1 = 300$ K, $T_2 = 273$ K (STP)

$$V_1 = 300 \text{ ml} = \left(\frac{300}{1000}\right) \text{ litre, } P_1 = \left(\frac{730}{760}\right)$$

$$\text{atm. } P_2 = 1 \text{ atm., } V_2 = ?$$

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

$$\frac{730 \times 300}{760 \times 1000 \times 300} = \frac{1 \times V_2}{273}$$

$$\therefore V_2 = 0.2622 \text{ litre} = 262.2 \text{ ml.}$$

Ex.2 A truck carrying oxygen cylinders is filled with oxygen at -23°C and at a pressure of 3 atmosphere in Srinagar, Kashmir. Determine the internal pressure when the truck drives through Madras, Tamilnadu. Where the temperature is 30°C -

- (A) 2.64 atm. (B) 1.64 atm.
(C) 1 atm. (D) 3.64 atm.

Sol. (D)

$$P_1 = 3 \text{ atm., } P_2 = ?$$

$$T_1 = -23 + 273 = 250 \text{ K}$$

$$T_2 = 273 + 30 = 303 \text{ K.}$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2},$$

$$\frac{3}{250} = \frac{P_2}{303}$$

$$P_2 = \frac{3 \times 303}{250} = 3.64 \text{ atm.}$$

Ex.3 The density of a gas at -23°C and 780 torr is 1.40 gram per litre. Which one of the following gases is it -

- (A) CO₂ (B) SO₂
(C) Cl₂ (D) N₂

Sol. (D)

$$PV = nRT$$

$$\frac{780}{760} \times 1 = n \times .082 \times (273 - 23)$$

$$n = \frac{780}{760 \times 0.082 \times 250}$$

$$n = 0.0501 \text{ moles.}$$

$$\therefore .0501 \text{ moles of the gas weigh} = 1.40 \text{ gm.}$$

$$\therefore 1 \text{ mole of gas weigh} = \frac{1.40}{.0501} = 28 \text{ gram.}$$

So gas is N₂.

Ex. 4 Calculate the weight of CH₄ in a 9 litre cylinder at 16 atm and 27°C (R = 0.08 lit. atm/K) -

- (A) 96 gm (B) 86 gm
(C) 80 gm (D) 90 gm

Sol. (A)

$$\text{Given } P = 16 \text{ atm, } V = 9 \text{ litre.}$$

$$T = 300 \text{ K, } m_{\text{CH}_4} = 16, R = 0.08 \text{ litre atm/k.}$$

$$PV = w/m \times R \times T$$

$$16 \times 9 = \frac{w}{16} \times 0.08 \times 300$$

$$w = 96 \text{ gm.}$$

Ex. 5 What is the density of sulphur dioxide (SO₂) at STP -

- (A) 2.86 gm/lit. (B) 1.76 gm/lit
(C) 1.86 gm/lit (D) None of these.

Sol. (A)

The gram molecular weight of SO₂ = 64 gm/mole.

Since 1 mole of SO₂ occupies a volume of 22.4 litres at S.T.P.

$$\text{Density of SO}_2 = \frac{64}{22.4} = 2.86 \text{ gm/lit.}$$

Ex.6 5gm of XeF₄ gas was introduced into a vessel of 6 litre capacity at 80°C. What is the pressure of the gas in atmosphere -

- (A) .21 atm (B) .31 atm
(C) .11 atm (D) .41 atm.

Sol. (C)

Given $V = 6$ litre, $T = 353$ K, $R = 0.082$, $W = 5$ gm. $m = 207.3$

$$PV = \frac{W}{m} \times R \times T$$

$$P \times 6 = \frac{5}{207.3} \times 0.082 \times (273 + 80)$$

$$P = \frac{5 \times 0.082 \times 353}{6 \times 207.3} = .11 \text{ atm.}$$

Ex.7 Calculate the temperature at which 28 gm N_2 occupies a volume of 10 litre at 2.46 atm-

- (A) 300 K (B) 320 K
(C) 340 K (D) 280 K

Sol. (A)

Given $w_{N_2} = 28$ gm, $P = 2.46$ atm, $V = 10$ litre. $m_{N_2} = 28$,

$$\therefore PV = \left(\frac{w}{m}\right) RT$$

$$2.46 \times 10 = \left(\frac{28}{28}\right) \times 0.0821 \times T$$

$$T = 300 \text{ K.}$$

Ex.8 A mixture of gases at 760 mm pressure contains 65% nitrogen, 15% oxygen and 20% Carbon dioxide by volume. What is the partial pressure of each in mm -

- (A) 494, 114, 252 (B) 494, 224, 152
(C) 494, 114, 152 (D) None of these.

Sol. (C)

$$P'_{N_2} = 760 \times \frac{65}{100} = 494 \text{ mm}$$

$$P'_{O_2} = 760 \times \frac{15}{100} = 114 \text{ mm}$$

$$P'_{CO_2} = 760 \times \frac{20}{100} = 152 \text{ mm.}$$

Ex.9 0.45 gm of a gas 1 of molecular weight 60 and 0.22 gm of a gas 2 of molecular weight 44 exert a total pressure of 75 cm of mercury. Calculate the partial pressure of the gas 2 -

- (A) 30 cm of Hg (B) 20 cm of Hg
(C) 10 cm of Hg (D) 40 cm of Hg.

Sol. (A)

$$\begin{aligned} \text{No. of moles of gas 1} = n_1 &= \frac{w_1}{m_1} = \frac{0.45}{60} \\ &= 0.0075 \end{aligned}$$

$$\begin{aligned} \text{No. of moles of gas 2} = n_2 &= \frac{w_2}{m_2} = \frac{0.22}{44} \\ &= 0.0050 \end{aligned}$$

$$\begin{aligned} \text{Total no. of moles} &= n_1 + n_2 \\ &= 0.0075 + 0.0050 = 0.0125 \end{aligned}$$

$$\begin{aligned} P_2 \text{ partial pressure of gas 2} &= \frac{0.0050}{0.0125} \times 75 \\ &= 30 \text{ cm of Hg.} \end{aligned}$$

Ex. 10 The total pressure of a sample of methane collected over water is 735 torr at 29°C. The aqueous tension at 29°C is 30 torr. What is the pressure exerted by dry methane -

- (A) 605 torr (B) 205 torr
(C) 405 torr (D) 705 torr

Sol. (D)

$$\begin{aligned} P_{\text{total}} &= P_{\text{dry methane}} + P_{\text{water}} \\ 735 &= P_{\text{dry methane}} + 30 \end{aligned}$$

$$\therefore P_{\text{dry methane}} = 735 - 30 = 705 \text{ torr.}$$

Ex.11 The odour from a gas A takes six seconds to reach a wall from a given point. If the molecular weight of gas A is 46 grams per mole and the molecular weight of gas B is 64 grams per mole. How long will it take for the odour from gas B to reach the same wall from the same point. Approximately -

- (A) 6 Sec (B) 7 Sec
(C) 8 Sec (D) 9 Sec

Sol. (B)

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

$$\frac{r_A}{r_B} = \sqrt{\frac{64}{46}} = 1.18$$

Time taken for the odour of B to reach the wall = $1.18 \times 6 = 7.08 \text{ sec} \approx 7 \text{ sec.}$

Ex.12 1 litre of oxygen effuses through a small hole in 60 min. and a litre of helium at the same temperature and pressure effuses through the same hole in 21.2 min. What is the atomic weight of Helium -

- (A) 2.99 (B) 3.99
(C) 2.08 (D) 1.99

Sol. (B)

$$\frac{r_{O_2}}{r_{He}} = \frac{1000/60}{1000/21.2} = \frac{21.2}{60} = \sqrt{\frac{M_{He}}{M_{O_2}}}$$
$$= \sqrt{\frac{M_{He}}{32}}$$

Squaring both of sides $\frac{(21.2)^2}{(60)^2} = \frac{M_{He}}{32}$

$$M_{He} = \frac{(21.2)^2 \times 32}{(60)^2} = 3.99$$

Since Helium is monoatomic so
Atomic weight = Molecular weight = 3.99

Ex.13 What is the temperature at which oxygen molecules have the same r.m.s. velocity as the hydrogen molecules at 27°C -

- (A) 3527°C (B) 4227°C
(C) 4527°C (D) 4000°C

Sol. (C)

$$\text{r.m.s. velocity } C = \sqrt{\frac{3RT}{M}}$$

$$\text{For oxygen } C_{O_2} = \sqrt{\frac{3RT}{32}}$$

$$\text{For } C_{H_2} \text{ at } 27^\circ\text{C} = \sqrt{\frac{3 \times R \times 300}{2}}$$

$$\text{When } C_{O_2} = C_{H_2}$$

$$\sqrt{\frac{3RT}{32}} = \sqrt{\frac{3 \times R \times 300}{2}}$$

$$\frac{3RT}{32} = \frac{3R \times 300}{2}$$

$$T = \frac{300 \times 32}{2} = 4800 \text{ K.}$$

$$\text{Then in } ^\circ\text{C} \quad T = 4800 - 273 = 4527^\circ\text{C}$$

Ex.14 Calculate the total kinetic energy in joules, of the molecules in 8 gm of methane at 27°C-

- (A) 1770.5 Joule (B) 1870.5 joule
(C) 1970.5 joule (D) 1670.5 joule

Sol. (B)

$$E_K \text{ (For 1 mole)} = \frac{3}{2} RT = 3741 \text{ Joule.}$$

Total energy of 8 gm of methane

$$= \frac{1}{2} \text{ mole of methane}$$

$$= \frac{3741}{2} = 1870.5 \text{ Joule.}$$

Ex.15 Calculate the root mean square velocity of SO_2 at S.T.P. -

- (A) 3.26×10^4 cm/sec
(B) 1.26×10^2 cm/sec
(C) 1.26×10^4 cm/sec
(D) 3.26×10^2 cm/sec

Sol. (A)

$$\therefore v_{\text{rms}} \text{ of } SO_2 = \sqrt{\frac{3RT}{M}}$$

$$= \sqrt{\frac{3 \times 8.314 \times 10^7 \times 273}{64}} = 3.26 \times 10^4 \text{ cm/sec.}$$

Ex.16 Calculate the number of atoms in 1 g of helium -

- (A) 1.506×10^{23} atoms
(B) 1.605×10^{32} atoms
(C) 1.056×10^{25} atoms
(D) None of these

Sol. (A)

Atomic mass of He = 4

4 g of He contain = 6.023×10^{23} atoms

\therefore 1 g of He contains =

$$\frac{6.023 \times 10^{23}}{4} = 1.506 \times 10^{23} \text{ atoms.}$$

Ex.17 7.00 g. of a gas occupies a volume of 4.1 litres at 300 K and 1 atmosphere pressure.

Calculate the molecular mass of the gas -

- (A) 40 g mol⁻¹ (B) 42 g mol⁻¹
(C) 48 g mol⁻¹ (D) 45 g mol⁻¹

Sol. (B)

$$PV = nRT \quad \therefore \quad n = PV/RT$$

$$n = \frac{(1 \text{ atm.})(4.1 \text{ L})}{(0.082 \text{ L atm K}^{-1} \text{ mol}^{-1})(300 \text{ K})} = \frac{1}{6} \text{ mol}$$

$$\text{now } n = \frac{\text{Mass of gas}}{\text{Mol. mass of gas}} = \frac{1}{6}$$

$$= \frac{7}{\text{Mol. mass of gas}}$$

$$\text{Thus, molecular mass of gas} = 7 \times 6 = 42 \text{ g mol}^{-1}$$

- Ex.18** Calculate density of ammonia at 30°C and 5 atm. pressure -
 (A) 3.03 g/litres. (B) 3.82 g/litres
 (C) 3.42 g/litres. (D) 4.42 g/litres.

Sol. (C)

$$PV = nRT, \text{ or } PV = \frac{m}{M} RT \text{ or } P = \frac{m}{V} \times \frac{RT}{M}$$

$$\text{or } P = d \times \frac{RT}{M}$$

$$d = MP/RT ; d = \frac{17 \times 5}{0.082 \times 303} = 3.42 \text{ g/litres.}$$

- Ex.19** 3 moles of a gas are present in a vessel at a temperature of 27°C . Calculate the value of gas constant (R) in terms of kinetic energy of the molecules of gas -
 (A) 7.4×10^{-4} KE per degree kelvin.
 (B) 4.5×10^{-4} KE per degree kelvin.
 (C) 7.4×10^{-5} KE per degree kelvin.
 (D) None of these

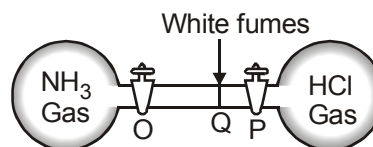
Sol. [A] K.E. for 1 mole = $3/2 RT$

$$\text{K.E. for 3 moles} = 9/2 RT. \text{ or } R = \frac{2}{9 T} \text{ KE}$$

$$= \frac{2}{9 (300)} \text{ KE}$$

$$= 7.4 \times 10^{-4} \text{ KE per degree kelvin.}$$

- Ex.20** In the following diagram, container of NH_3 gas and container of HCl gas, connected through a long tube, are opened simultaneously at both ends; the white NH_4Cl ring first formed will be at Q point. If $OP = 40\text{cm}$, then find OQ -



- (A) 35 cm (B) 23.74cm
 (C) 30 cm (D) 31.25 cm

Sol. [B]

Let $OQ = x$ cm so $QP = (40 - x)$ cm
 Diffused volume of NH_3 gas = Area of T.S. of tube \times Distance travelled by NH_3 gas

$$V_{\text{NH}_3} = A \times OQ = Ax$$

{Where A is area of T.S. of tube}

Similarly in the same time,

Diffused volume of HCl gas = Area of T.S. of tube \times Distance travelled by HCl gas

$$V_{\text{HCl}} = A \times QP = A (40 - x)$$

From Graham's Law of diffusion

$$\Rightarrow \frac{r_{\text{NH}_3}}{r_{\text{HCl}}} = \sqrt{\frac{M_{\text{HCl}}}{M_{\text{NH}_3}}}$$

$$\Rightarrow \frac{V_{\text{NH}_3}/t}{V_{\text{HCl}}/t} = \sqrt{\frac{36.5}{17}} = 1.46$$

$$\Rightarrow \frac{x}{(40 - x)} = 1.46$$

$$\Rightarrow x = 23.74 \text{ cm}$$

$$\therefore OQ = 23.74 \text{ cm}$$