GASEOUS STATES

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 then.

(c) Gases have unlimited dispensability and high compressibility.

(d) They have very low densities because of negligible intermolecular forces.

(e) Gases exerts 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,

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

so, $m = n \times M$

(b) Volume V-

(i) Gases occupy whole space available to then. The volume occupy 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 litre = 1 dm^3 = 1000 cm^3 = 1000 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 the direction.

(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 atm = 76 cm Hg = 760 mm Hg = 1.013×10^5 N/m^2 = 1.013×10^5 P_a = 1.013 bar = 760 torr.

(v) $$P\ (\text{Pressure}) = \frac{F\,(\text{Force})}{A\,(\text{Area})} = \frac{\text{Mass} \times \text{Accelaration}}{\text{Area}}$$

(vi) Pressure exerted by a gas is due to kinetic energy (K.E. = ½ mv^2) of the gases 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, ^oC, ^oF, ^oR.

(iii) K = ^oC + 273.15

(iv) $$\frac{x^0C}{5} = \frac{(y^0f - 32)}{9}$$

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.

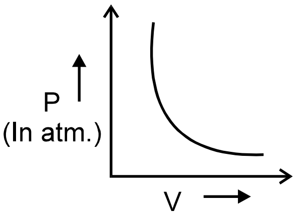
(b) Mathematically -

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

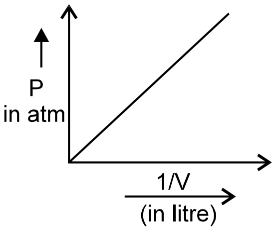
or $PV = K$ or $P_1V_1 = P_2V_2$

(c) Graphical representations

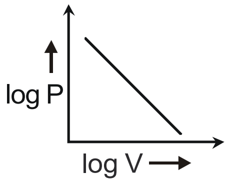
(i) P vs V



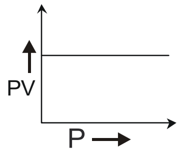
(ii) P vs $\frac{1}{V}$



(iii) log P vs, log V



(iv) PV vs P



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 = ^oC + 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,

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

or $\log V - \log T$ = constant

(c) For each degree change of temperature the volume of sample of a gas changes by the fraction of $1/273.15$ of its volume at 0 ºC

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

This equation is called Charles-gay-lussac equation.

where

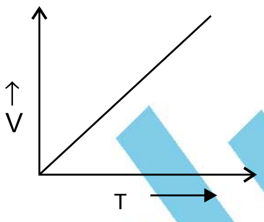
V_t = volume of gas at temperature t ºC

V_0 = volume of gas at 0 ºC temperature

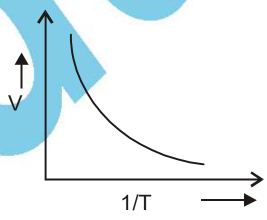
t = temperature in ºC.

(d) Graphical representations -

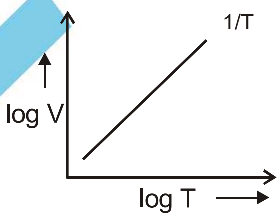
(i) V vs T



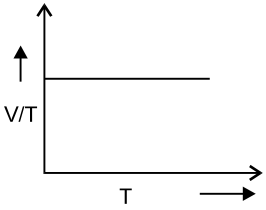
(ii) V vs $\frac{1}{T}$ or T vs $\frac{1}{V}$



(iii) log V vs log T,



(iv) $\frac{V}{T}$ vs T



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

$P = KT$

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.

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

$\log P - \log T$ = constant.

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

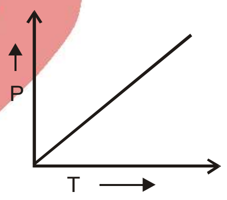
where P_t = Pressure of gas at t ºC

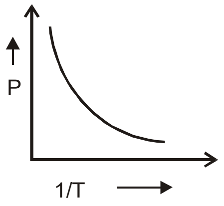
P_o = Pressure of gas at 0 ºC

t = Temperature in ºC.

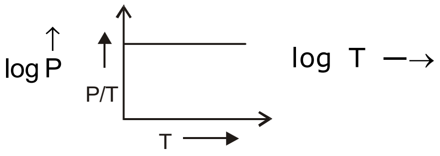
(d) Graphical representation -

(i) P vs T (ii) P vs $1/T$ or T vs $1/P$





(iii) log P vs log T (iv) P/T V_s T



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

$$PV = RT$$

$$\text{So } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = R \quad \text{or} \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

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 -

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

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

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

(f) Numerical values of R in different units -

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

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.- $P = P_1 + P_2 + P_3 + \ldots\ldots + P_n$

(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 -

$$PV = (n_1 + n_2)\, RT \qquad \ldots\ldots\text{(i)}$$

Partial pressure of individual gas calculates at

$$P_AV = n_1\, RT \qquad \ldots\text{(ii)}$$

$$P_BV = n_2\, RT \qquad \ldots\text{(iii)}$$

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

$$(P_A + P_B)\, V = (n_1 + n_2)\, RT \quad \ldots\ \text{(iv)}$$

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 × 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-

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

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

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

(**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.

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.

$V \propto n$ (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.

LAW

(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.

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.

$\because$ 2 × vapour density = Molecular mass

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

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.

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 ∝ 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 ∝ 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-

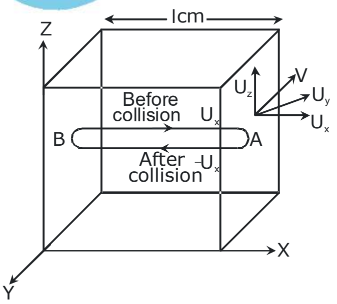
$$PV = \frac{1}{3}\,mnv^2$$

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.



for one mole of the gas -

PV = RT and n = N

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

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

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

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

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

$\because$ Boltzmann constant $K = R/N$

$= 1.38 \times 10^{-16}$ ergs K^{-1} molecules^{-1}

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

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

Average speed = .9213 × RMS speed
RMS speed = 1.085 × Average speed

(c) Most Probable Speed-

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

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

MPS = .816 × RMS ; RMS = 1.224 MPS
MPS : A.V. speed : RMS = 1 : 1 . 128 : 1.224

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 highT.
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)

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

at 1 °C ⇒ K.E = 3/2 R

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

Therefore ⇒ 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.

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

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 connect 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.

$$V_C = 3\ b$$

11. LIQUIFICATION OF GASES :

(a) On increasing the pressure gas can be easily liquified because intermolecular attraction between gaseous molecule 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.

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

13. INVERSION TEMPERATURE :

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

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

SOLVED PROBLEMS

Ex.1 ***A gas occupies 300 ml at 27ºC 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)

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

atm. P_2 = 1 atm., V_2 = ?

$$\because \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}, \frac{730 \times 300}{760 \times 1000 \times 300} = \frac{1 \times V_2}{273}$$

$\therefore$ V_2 = 0.2622 litre = 262.2 mℓ.

Ex.2 ***A truck carrying oxygen cylinders is filled with oxygen at -23ºC 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 30ºC -***

(A) 2.64 atm. ***(B) 1.64 atm.***

(C) 1 atm. ***(D) 3.64 atm.***

Sol. **(D)**

P_1 = 3 atm., P_2 = ?

T_1 = – 23 + 273 = 250 K

T_2 = 273 + 30 = 303 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 -23ºC and 780 torr is 1.40 gram per litre. Which one of the following gases is it -***

(A) CO_2 ***(B) SO_2*** ***(C) Cl_2*** ***(D) N_2***

Sol. **(D)**

PV = nRT

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

$$n = \frac{780}{760 \times 0.082 \times 250} \quad n = 0.0501 \text{ moles.}$$

$\because$ .0501 moles of the gas weigh = 1.40 gm.

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

So gas is N_2.

Ex. 4 ***Calculate the weight of CH_4 in a 9 litre cylinder at 16 atm and 27ºC (R = 0.08 lit. atm/k) -***

(A) 96 gm ***(B) 86 gm***

(C) 80 gm ***(D) 90 gm***

Sol. **(A)**

Given P = 16 atm, V = 9 litre.

T = 300 K, m_{CH4} = 16, R = 0.08 litre atm/k.

PV = w/m × R × T

$$16 \times 9 = \frac{w}{16} \times 0.08 \times 300 \quad w = 96 \text{ gm.}$$

Ex. 5 ***What is the density of sulphur dioxide (SO_2) 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 SO_2 = 64 gm/mole.

Since 1 mole of SO_2 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 XeF_4 gas was introduced into a vessel of 6 litre capacity at 80ºC. 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 = 5gm. m= 207.3

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

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

$$P = \frac{5 \times .082 \times 353}{6 \times 207.3} = .11 \text{ atm.}$$

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_{N2} = 28gm , P = 2.46 atm,

V = 10 litre. m_{N2} = 28,

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

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

T = 300 K.

Ex.8 ***A mixture of gases at 760 mm pressure contains 65% nitrogen, 15% oxygen and 20% Carbondioxide 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)**

$$P'_{N2} = 760 \times \frac{65}{100} = 494 \text{ mm}$$

$$P'_{O2} = 760 \times \frac{15}{100} = 114 \text{ mm}$$

$$P'_{CO2} = 760 \times \frac{20}{100} = 152 \text{ mm.}$$

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 75cm 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)**

No. of moles of gas 1 = n_1

$$= \frac{w_1}{m_1} = \frac{0.45}{60} = 0.0075$$

No. of moles of gas 2 = n_2

$$= \frac{w_2}{m_2} = \frac{0.22}{44} = 0.0050$$

Total no. of moles = $n_1 + n_2$

= 0.0075 + 0.0050 = 0.0125

P_2 partial pressure of gas 2

$$= \frac{0.0050}{0.0125} \times 75 = 30 \text{ cm of Hg.}$$

Ex.10 ***The total pressure of a sample of methane collected over water is 735 torr at 29ºC. The aqueous tension at 29ºC 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)**

$P_{total} = P_{dry\ methane} + P_{water}$

$735 = P_{dry\ methane} + 30$

$\therefore\ P_{dry\ methane} = 735 - 30 = 705$ 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)**

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}, \quad \frac{r_A}{r_B} = \sqrt{\frac{64}{46}} = 1.18$$

Time taken for the odour of B to reach the wall = 1.18 x 6 = 7.08 sec ≅ 7 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 $\frac{(21.2)^2}{(60)^2} = \frac{M_{He}}{32}$

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

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 27ºC -***

(A) 3527ºC ***(B) 4227ºC***

(C) 4527ºC ***(D) 4000ºC***

Sol. **(C)**

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

For oxygen $Co_2 = \sqrt{\frac{3RT}{32}}$

For CH_2 at 27ºC = $\sqrt{\frac{3 x R x 300}{2}}$

When $Co_2 = CH_2$

$$\sqrt{\frac{3RT}{32}} = \sqrt{\frac{3 x R x 300}{2}}, \quad \frac{3RT}{32} = \frac{3R x 300}{2}$$

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

Then in ºC T = 4800 - 273 = 4527ºC

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

(A) 1770.5 Joule ***(B) 1870.5 joule***

(C) 1970.5 joule ***(D) 1670.5 joule***

Sol. (B)

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

Total energy of 8 gm of methane

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

$= \frac{3741}{2} = 1870.5$ 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)

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

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

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

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

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

$$\text{now} \quad n = \frac{\text{Mass of gas}}{\text{Mol. mass of gas}} = \frac{1}{6} = \frac{7}{\text{Mol. mass of gas}}$$

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

Ex.18 Calculate density of ammonia at 30ºC 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 27ºC . Calculate the value of gas constant (R) in terms of kinetic energy of the molecules of gas -

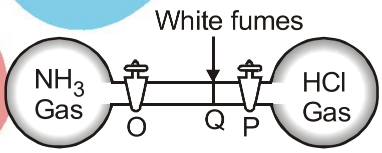
(A) 7.4×10^{-4} KE per degree kelvin.
(B) 4.5×10^{-4} KE per degree kelvin.
(C) 7.4×10^{-5} KE per degree kelvin.
(D) None of these

Sol. (A) K.E. for 1 mole = 3/2 RT

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

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

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 = 40cm, 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 × Distance travelled by NH_3 gas

$V_{NH_3} = A \times OQ = Ax$

{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 × Distance travelled by HCl gas

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

From Graham's Law of diffusion

$$\Rightarrow \frac{r_{NH_3}}{r_{HCl}} = \sqrt{\frac{M_{HCl}}{M_{NH_3}}}$$

$$\Rightarrow \frac{V_{NH_3/t}}{V_{HCl/t}} = \sqrt{\frac{36.5}{17}} = 1.46$$

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

$\Rightarrow$ x = 23.74 cm

$\therefore$ OQ = 23.74 cm