Gaseous State

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

so, m = n x M

(b) Volume V -

(i) Gases occupy whole space available to then. The volume occypy 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 of 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 (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 molecule 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) $\frac{x^oC}{5} = \frac{(y^of - 32)}{9}$

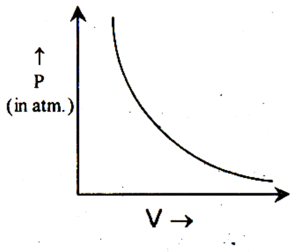
2. GAS LAWS

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

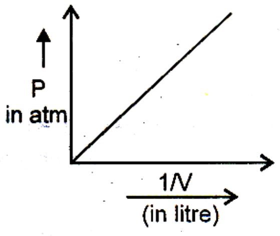
2.1 Boyle's law

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

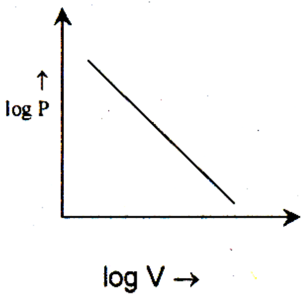
(b) Mathematically



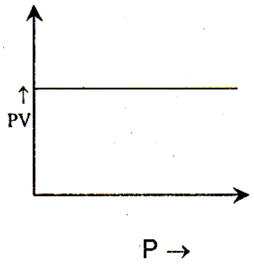
(i) P vs V



(ii) P vs $\frac{1}{v}$ or V Vs $\frac{1}{p}$



(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 = ^{o}C + 273.15)

(b) Mathematically – $V \propto T$ (at constant pressure)
V = volume of gas
T = Absolute temperature

$$V = KT \quad \text{or} \quad \frac{V}{T} = K$$

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

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 1/273.15 of its volume at 0^{o}C

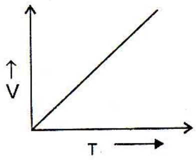
so $$V_t = V_0 \left[\frac{273.15 + t}{273.15}\right]$$

This equation is called charles-gay-lussac equation.

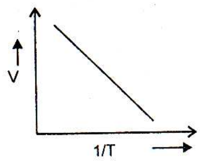
where V_t = volume of gas at temperature t^{o}C
V_0 = volume of gas at 0^{o}C temperature
t = temperature in ^{o}C

(d) Graphical representation

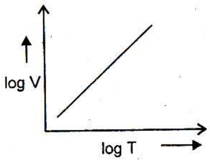
(i) V vs T



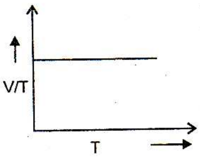
(ii) V vs $\frac{1}{T}$ or T vs $\frac{1}{V}$



(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.

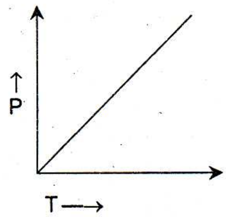
then $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ = constant

log P – log T = constant

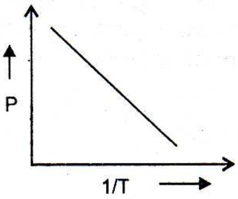
(c) $P_t = P_0 \left(1 + \frac{t}{273.15}\right)$

Where P_t = Presure of gas at t°C
P_o = Pressure of gas at 0°C
t = Temperature in °C

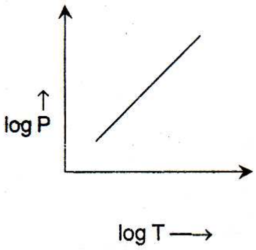
(d) Graphical representation -

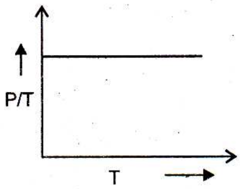


(i) P V_s T



(ii) P V_s ${}^1\!/_T$ or T V_s ${}^1\!/_P$





(iv) ${}^P\!/_T$ V_s T

2.4 Ideal gas equation

(a) It correlats 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 gas n = 1

$$PV = RT$$

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

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 -

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

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

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

(f) Numerical values of R in different units -

(i) R = 0.0821 litre atm deg^{-1} mole^{-1}

(ii) R = 62.4 litres mm deg^{-1} mole^{-1}

(iii) R = 8/314 x 10^7 ergs deg^{-1} mole^{-1}

(iv) R = 82.05 cc atm deg.^{-1} mole^{-1}

(v) R = 2 cals. deg^{-1} mole^{-1}

(vi) R = 8.314 J $K^{-1}\text{mole}^{-1}$

3. DALTON’S LAW OF PARTIAL PERSSURE

(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.

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

(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 -

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

Partial pressure of individual gas calcualtes at

$$P_AV = n_1RT \quad \ldots\ldots\ldots \text{(ii)}$$

$$P_BV = n_2RT \quad \ldots\ldots\ldots \text{(iii)}$$

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

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

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

$$P = P_A + P_B$$

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

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

$$P_A = x_A \text{ x } P$$

where x_A = mole fraction of 'A'

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

$$P_B = x_B = P$$

so

Partial pressure of a component = mole fraction x 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 $= \dfrac{\text{partial pressure of gas}}{\text{Total pressure}}$ x 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 $P_{dry\ gas}$ = Total measured pressure - Aqueous tension

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

3.2 Limitations of Dalton's law of partial pressure

(a) It is applicable only for 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. AVOGADRO'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 of 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 x 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 x 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.

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.

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

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

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 × 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 = $\dfrac{\text{Volume of gas diffused}}{\text{Time taken for diffusion}}$

5.1 Applications of graham's law of diffusion

(a) In 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 Clasius, 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 direction.

(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 law 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

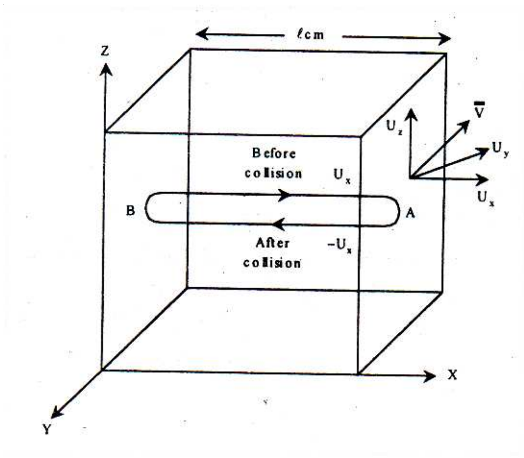
(h) The force of gravity has no effect on the speed of gas molecules.

6.2 Calculation of Kinetic energy

According to postulates of kinetic gas equation -

$$PV = \frac{1}{3}\, mn\bar{v}^2$$

where
P = pressure of the gas
V = volume of the gas
n = no of molecules present in the given amount of gas
$\bar{v}$ = Root mean square speed.



for one mole of the gas -

$$PV = RT \qquad \text{and} \qquad n = N \qquad \qquad \frac{1}{3}\, mNv^2 = RT$$

$$\text{or} \quad \frac{2}{3} \cdot \frac{1}{2}\, mNv^2 = RT \qquad \qquad (\therefore\ \frac{1}{2}\, m\, Nv^2 = \text{K.E. per mole})$$

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

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

$\therefore$ Boltzmann constant K = R/N = 1.38 x 10^{-16} ergs K^{-1} molecules^{-1}

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

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.

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

(b) Average Speed -

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

$$\text{Average speed} = \sqrt{\frac{8RT}{\pi m}}$$

Average speed = .9213 x RMS speed

RMS speed = 1.085 x Average speed

(c) Most probable speed -

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

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

MPS = .816 x RMS ; RMS = 1.224 MPS

MPS : A.V. speed : RMS = 1 : 1. 128 : 1224

7. MEAN FREE PATH (λ)

(a) The distance travelled by the gaseous molecule between the two successive collision is known as mean free path which is denoted by lambda (λ)

$$\lambda = \frac{1}{\sqrt{2}\,\pi\sigma^2} \cdot \frac{RT}{PV}$$

(b) As size of gas molecules increases mean free path decreases.

(c) At the room temperature and at the pressure of 1 atm, the order of mean free path is of 10^{-5} cm.

8. COLLISION FREQUENCY

(a) Collision frequency is defined as the no of collisions between gases molecules in per second per unit of volume and it is directly proportional to the pressure and density also to the

$Z = \frac{1}{2} \sigma \pi^2 N^2 v$ where Z = collision fequency

σ = collision diameter
v = average velocity

(b) $Z = \frac{1}{2}$ x no. of collision x total no. of molecules

9. DIFFERENCE BETWEEN REAL GAS & IDEAL GAS

S.No.	Ideal gas	Real gas
1.	Obeys gas law under all conditions of P and T.	Obeys only at low P and 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 neigligible 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

9.1 Explanation of deviation

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

(a) The force of attraction between gasesous molecules are negligible

(b) The volume occupied by the gaseseous molecules is negligible compared to total volume of gas at high pressure. Both the assumptions do not hold good hence deviating from ideal gas.

10. 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 1ºC (more accurately from 15.5 ºC to 16.5ºC).

The heat so supplied is used up in increasing the internal energy of themolecule, 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 1ºC when the volume is kept constant and pressure is allowed to increse.

(d) Specific Heat of gas at constant Pressure (Cp) –

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

$C_V = C_V \text{ x } M$ and $C_p = C_p \text{ x } M$

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

(e) Ratio for Molar Heats -

(i) For monoatomic gases $\frac{C_P}{C_V} = \gamma = \frac{5.0}{3.0} = 1.66$

∴ Molar heat capacity at constant volume $(C_V) = \text{K.E.} = \frac{3RT}{2}$

at 1 ºC ⇒ K.E. = 3/2 R ; K.E. = $\frac{3}{2}$ x 2 = 3

Therefore ⇒ C_V = 3 calories.

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

$$= \frac{3}{2}R + R = \frac{5}{2}R \text{ (Since R = 2) ;} \quad C_p = \frac{5}{2} \text{ x } 2 = 5 \text{ calories.}$$

(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 $\frac{C_P}{C_V} = \frac{5+3}{3+3} = \frac{8}{6} = 1.33$

11. VANDER WALL'S EQUATION - A MODIFICATION IN GAS EQUATION

According to Vander Waal Maxwell has made to 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.

11.1 Limitationsof van der Waal's gas equation

(a) It could not explain the quantitative aspect of deviation. So satisfactorily as it coluld explain the qualitative aspect of deviations.

(b) The values of 'a' and 'b' were also found to vary with P and T.

(c) The numerical value of V_C was not equal to 3b as derived by Vander Waals' equation.

(d) The numerical value of RT_C/P_CV_C was not found to be 8/3 as derived by van der Waal's equation.

12 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. $PC = \frac{a}{27b^2}$

(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. $V_c = 3b$

13. 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)

Solved Example

Ex.1 At fixed temperature and 600 mm pressure, the density of a gas is 42. At the same temperature and 700mm pressure, what is the denisty of the gas?

Sol. According to Boyle's law.

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

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

$$\text{or} \quad D \propto \frac{1}{V}$$

$$P \propto D$$

$$\frac{P_1}{D_1} = \frac{P_2}{D_2}$$

$$D_2 = \frac{P_2 \times D_1}{P_1}$$

$P_1 = 600$, $D_1 = 42$, $P_2 = 700$

$$D_2 = \frac{700 \times 42}{600} = 49$$

Ex.2 Volume of given amount of a gas at 57ºC and constant pressure is 425.8 cm^3. If the temperature is decreased to 37ºC at constant pressure, then the volume will be

Sol. According to charles law

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

$V_1 = 425.8 \text{ cm}^3 \quad V_2 = ?$
$T_1 = 273 + 57 = 330 \text{ K}$
$T_2 = 273 + 37 = 310 \text{ K}$

$$V_2 = \frac{425.8 \times 310}{300} = 400 \text{ cm}^3$$

Q.3 At what temeprature, the volume of given amount of a gas at 25º C becomes twice, when pressure is kept constant?

Sol. According to Charles law

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

$V_1 = v \qquad V_2 = 2v$
$T_1 = 273 + 25 = 298 \qquad T_2 = ?$

$$\frac{\nu}{298} = \frac{2\nu}{T_2}; \quad T_2 = \frac{298 \times 2}{1} = 596 \text{ K}$$

Q.4 4 gram oxygen and 2 gram hydrogen are kept in a one-litre container at 0ºC. Calculate the total pressure.

Sol. According to ideal gas equation

$$PV = nRT$$

$$P = ?, V = 1 \text{ litre}, n = \frac{4}{32} + \frac{2}{2} = 1.125$$

R = 0.821 L. atm. mol^{-1} K^{-1}
T = 0ºC + 273 = 273 K

$$P = \frac{1.125 \times 0.0821 \times 273}{1}$$

= 25.2 atm

Ex.5 A gaseous mixture of 0.336 gram N_2 and 11.2 gram litre O_2 is taken at NTP in a ten-litre flask at 27ºC. Calculate the partial pressure and total pressure of each gas.

Sol. Mole of N_2 = $\frac{\text{Mass of } N_2}{\text{Molecular mass of } N_2}$

$$= \frac{0.336}{28} = 0.012 \text{ mole}$$

Mole of O_2 = $\frac{11.2}{22.4}$ = 0.5 mole

Partial pressure of N_2 $P_{N_2} = \frac{nRT}{V}$

$$= \frac{0.012 \times 0.0821 \times 300}{10}$$

= 0.0296 atm

Partial pressure of O_2, $P_{O_2} = \frac{0.05 \times 0.0821 \times 300}{10}$

Total pressure = 0.0296 + 1.2315 = 1.2611 atm

Ex.6 If pressure of a gas is reduced by 25%, then what should be the temperature required to make its volume twice at NTP?

Sol. According to combined gas equation

$$\frac{PV}{T} = \frac{P_1V_1}{T_1}$$

P = 76.0, P_1 = 57.0 (on reducing by 25%)
V = ν, V_1 = 2ν, T = 273K, T_1 = ?

$$\frac{76 \times \nu}{273} = \frac{57 \times 2\nu}{T_1}$$

$$T_1 = \frac{273 \times 57 \times 2}{76} = 409.3 \text{ K}$$

Ex.7 Total pressure of a gaseous mixture of 2.8 gram N_2, 3.2 gram O_2 and 0.50 gram H_2 is 4.5 atmosphere, calculate the partial pressure of each gas.

Sol. Number of moles or = $\frac{\text{Mass}}{\text{Molar mass}}$

$nN_2 = \frac{2.8}{28}$ = 0.1 mole ; $nO_2 = \frac{3.2}{32}$ = 0.1 mole ; $nH_2 = \frac{0.5}{2}$ = 0.25 mole

Total number of mole = 0.1 + 0.1 + 0.25 = 0.45

Partial pressure of a gas = $\frac{\text{Number of moles}}{\text{Total number of moles}}$ x Total pressure

$P_{N_2} = \frac{0.1}{0.45}$ x 4.5 atm = 1.0 atm ; $P_{O_2} = \frac{0.1}{0.45}$ = 4.5 atm = 1.0 atm

$P_{H_2} = \frac{0.25}{0.45}$ x 4.5 atm = 2.5 atm

Ex.8 Under similar conditions, 160 ml hydrocarbon diffuses in 16 minutes and 150 ml of sulphur dioxide in 30 minutes. If molecular weight of sulphur dioxide is 64, then find out the molecular weight of hydrocarbon.

Sol. Hydrocarbon	Sulphur dioxide
V_1 = 160 ml	V_2 = 120 ml
t_1 = 16 minute	t_2 = 20 minute
M_1 = ?	M_2 = 64

From Graham's law $\frac{V_1/t_1}{V_2/t_2} = \sqrt{\frac{M_2}{M_1}}$; $\frac{160/16}{150/30} = \sqrt{\frac{64}{M_1}}$

$$2 = \sqrt{\frac{64}{M_1}} \quad M_1 = 4 = \frac{64}{M_1} = M_1 = \frac{64}{4} = 16$$

Ex.9 Average velocity of CO_2 at the temperature T_1K and maximum possible velocity of CO_2 at the temprature T_2 K is 9 x 10^{-4} cm $second^{-1}$ then calculate the value of T_1 and T_2.

Sol. Average velocity = $\sqrt{\frac{8RT}{\pi M}}$

and maximum possible velocity = $\sqrt{\frac{2RT}{M}}$

Average velocity = Maximum possible velocity = 9 x 10^4 cm sec^{-1} = 9 x 10^2 meter/ second

$$\therefore \quad 9 \times 10^2 = \sqrt{\frac{8 \times 8.314 \times T_1}{3.14 \times 44 \times 10^{-3}}}$$

$$9 \times 10^2 = \sqrt{\frac{2 \times 8.314 \times T_2}{44 \times 10^{-3}}}$$

On solving T_1 = 1682.5 K

T_2 = 2143.4 K

Ex.10 Calculate the average kinetic energy of 8 gram molecules of methane at 27ºC in Joule.

Sol. Total kinetic energy = $n(\frac{3}{2}RT)$

Mole number in 8 gram methane = $\frac{8}{16}$ = 0.5

(Molecular weight of methane = 16)

R = 8.314 Joule / K / mole

T = 27 + 273 = 300 K

Therefore, kinetic energy = 0.5 x 3/2 x 8.314 x 300 = 1870.65 Joule

$\therefore$ Average kinetic energy = $\frac{1870.65}{6.023 \times 10^{23} \times 0.5}$ = 6.21 x 10^{-21} joule