IDEAL GASES

INTRODUCTION:

Matter as we know broadly exists in three states.

There are always two opposite tendencies between particles of matter which determine the state of matter.

- Inter molecular attractive forces.
- The molecular motion / random motion.



In this chapter the properties and behaviour of the gases will be analysed and discussed in detail. These properties are measured with the help of the gas laws as proposed Boyle, Charles, Gay lussac etc.

SECTION (A) : IDEAL GAS EQUATION & GAS LAWS

Th1

1.1. Boyle's law and measurement of pressure :

D1 Statement : For a fixed amount of gas at constant temperature, the volume occupied by the gas is inversely proportional to the pressure applied on the gas or pressure of the gas.

$$V \propto \frac{1}{P}$$

1

Hence PV = constant.

This constant will be dependent on the amount of the gas and temperature of the gas.





Application of Boyles Law : For the two points 'A' and 'B' $P_1V_1 = K \& P_2V_2 = K$ Hence it follows that $P_1V_1 = P_2V_2$.



<u>Units</u>

Volume	Pressure	Temperature
Volume of the gas is the	Pressure = N/m ² = Pa \rightarrow S.I. unit	Kelvin scale \rightarrow Boiling point = 373 K
Volume of the container	C.G.S unit = dyne-cm ²	ice point = 273 K
S.I. unit $\rightarrow m^3$	Convert 1N/m ² into dyne/cm ²	Farenheit scale \rightarrow B.P. = 212°F
C.G.S. unit →cm ³	1N _ 10 ⁵ dyne	ice point = 32°F
$1 \ell = 10^{-3} m^3$	$\frac{1}{1} \text{m}^2 = \frac{10^4 \text{ cm}^2}{10^4 \text{ cm}^2}$	Celcius scale \rightarrow B.P. = 100°C
$1 \ \ell = 10^3 \ \mathrm{cm}^9$	$1 \text{ N/m}^2 = 10 \text{ dyne/cm}^2$	ice point = 0°C
1 dm ³ = 1 ℓ = 10 ⁻³ m ³	1 atm = 1.013 × 10 ⁵ N/m ²	$\frac{C-0}{C-0} = \frac{K-273}{K-273} = \frac{F-32}{K-R(0)} = \frac{K-R(0)}{K-R(0)}$
$1 \text{ ml} = 10^{-3} \ell = 1 \text{ cm}^3 = 1 \text{ cc}$	1 bar = 1 × 10^5 N/m ²	100-0 373-273 212-32 R(100)-R(0)
	1 atm = 760 mm of Hg = 760 torr	where R = Temperature on unknown scale.

Atmospheric pressure :

The pressure exerted by atmosphere on earth's surface at sea level is called 1 atm.

1 atm = 1.013 bar

 $1 \text{ atm} = 1.013 \times 10^5 \text{ N/m}^2 = 1.013 \text{ bar} = 760 \text{ torr}$

Solved Example

Example 1. A rubber balloon contains some solid marbles each of volume 10 ml. A gas is filled in the balloon at a pressure of 2 atm and the total volume of the balloon is 1 litre in this condition. If the external pressure is increased to 4atm the volume of Balloon becomes 625 ml. Find the number of marbles present in the balloon.

Solution: Let the no. of marbles be = n.

volume of marble = 10 n ml.

volume of balloon earlier = 1000 ml.

later = 625 ml.

Now for the gas inside the balloon temperature and amount of the gas is constant, hence Boyles law can be applied

$$P_1V_1 = P_2V_2$$

$$4 \times (625 - 10n) = 2 \times (1000 - 10n)$$

$$625 \times 4 = 2000 - 20n + 40n$$

$$625 \times 4 - 2000 = 20n$$

$$\frac{625 \times 4 - 2000}{20} = n. \qquad \frac{125}{5} = n \quad ; \qquad n = 25$$



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Th2

1.2. Measurement of Pressure

Barometer : A barometer is an instrument that is used for the measurement of pressure. The construction of the barometer is as follows





A mercury barometer is used to measure atmospheric pressure by determining the height of a mercury column supported in a sealed glass tube. The downward pressure of the mercury in the column is exactly balanced by the outside atmospheric pressure that presses down on the mercury in the dish and pushes it up the column.



A thin narrow calibrated capillary tube is filled to the brim, with a liquid such as mercury, and is inverted into a trough filled with the same fluid. Now depending on the external atmospheric pressure, the level of the mercury inside the tube will adjust itself, the reading of which can be monitored. When the mercury column inside the capillary comes to rest, then the net forces on the column should be balanced. Applying force balance, we get,

 $P_{atm} \times A = m \times g$ ('A' is the cross-sectional area of the capillary tube)

If ' ρ ' is the density of the fluid, then m = $\rho \times v$

Hence, $P_{atm} \times A = (\rho \times g \times h) \times A$ (v = A × h)

('h' is the height to which mercury has risen in the capillary)

F2

or,

Normal atmospheric pressure which we call 1 atmosphere (1 atm), is defined as the pressure exerted by the atmosphere at mean sea level. It comes out to be 760 mm of Hg = 76 cm of Hg. (at mean sea level the reading shown by the barometer is 76 cm of Hg)

or

$$1 \text{ torr} = 1 \text{ mm of Hg}.$$

 $P_{atm} = \rho g h$

$$1 \text{ bar} = 10^5 \text{ N/m}^2 (Pa)$$

Faulty Barometer : An ideal barometer will show a correct reading only if the space above the mercury column is vacuum, but in case if some gas column is trapped in the space above the mercury column, then the barometer is classified as a faulty barometer. The reading of such a barometer will be less than the true pressure.





For such a faulty barometer $P_0A = Mg + P_{gas} A$ $P_0 = \rho gh + P_{gas}$







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Gaseous State

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Gaseous State

D2 Charles law :

For a fixed amount of gas at constant pressure volume occupied by the gas is directly proportional to temperature of the gas on absolute scale of temperature.



Relation : T = t + 273

- Since volume is proportional to absolute temperature. The volume of a gas should be theoretically zero at absolute zero temperature.
- Infact no substance exists as gas at a temperature near absolute zero, though the straight line plots can be extra plotted to zero volume. Absolute zero can never be attained practically though it can be approached only.
- By considering –273.15°C as the lowest approachable limit, Kelvin developed temperature scale which is known as absolute scale.

Solved Example

Example 1. If the temperature of a particular amount of gas is increased from 27°C to 57°C, find final volume of the gas, if initial volume = 1 It and assume pressure is constant.

Solution:	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	\Rightarrow	$\frac{1}{(273+27)} = \frac{V_2}{(273+57)}$	So	V ₂ = 1.1 lt.
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Example 2. An open container of volume 3 litre contains air at 1 atmospheric pressure. The container is heated from initial temperature 27°C or 300 K to t°C or (t + 273) K the amount of the gas expelled from the container measured 1.45 litre at 17°C and 1 atm.Find temperature t.

Solution:

$$T_1 = 300$$

...

It can be assumed that the gas in the container was first heated to (t + 273), at which a volume ' ΔV ' escaped from the container.

Hence applying charles law :
$$\frac{3}{300} = \frac{3 + \Delta V}{t + 273}$$

Now, this volume ' ΔV ' which escapes when the container get cooled

$$\therefore \qquad \frac{\Delta V}{t+273} = \frac{1.45}{290}$$

Solve the two equations and get the value of ΔV and t.

Determine Δ V & calculate t that will be the answer.



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Th3

1.3. Calculation of pay load :

Pay load is defined as the maximum weight that can be lifted by a gas filled balloon.



For maximum weight that can be lifted, applying force balance

 $F_{buoyancy} = M_{balloon} \times g + M_{pay load} \times g$ \Rightarrow ρ_{air} v.g. = ρ_{gas} v.g + Mg + mg. mass of balloon = m net force on volume of balloon = vballoon = 0density of air = ρ_{air} (at equilibrium / when balloon is incoming density of gas inside the with constant speed) balloon = ρ_{gas}

Solved example

Example 1. A balloon of diameter 20 m weights 100 kg. Calculate its pay-load, if it is filled with He at 1.0 atm and 27°C. Density of air is 1.2 kg m⁻³. [R = 0.0082 dm³ atm K⁻¹ mol⁻¹] Solution: Weight of balloon = $100 \text{ kg} = 10 \times 10^4 \text{ g}$

Volume of balloon =
$$\frac{4}{3}\pi r^3 = \frac{4}{3} \times \frac{22}{7} \times \left(\frac{20}{2} \times 100\right)^3 = 4190 \times 10^6 \text{ cm}^3 = 4190 \times 10^3 \text{ litre}$$

Weight of gas (He) in balloon = $\frac{\text{PVM}}{\text{RT}} = \frac{1 \times 4190 \times 10^3 \times 4}{0.082 \times 300} = 68.13 \times 10^4 \text{ g} \quad \left(\because \text{PV} = \frac{\text{w}}{\text{M}}\text{RT} \right)$

Total weight of gas and balloon = $68.13 \times 10^4 + 10 \times 10^4 = 78.13 \times 10^4 \text{ g}$ ÷.

Weight of air displaced =
$$\frac{1.2 \times 4190 \times 10^6}{10^3}$$
 = 502.8 × 10⁴ g

Pay load = $502.8 \times 10^4 - 78.13 \times 10^4 = 424.67 \times 10^4 g$ *.*..

D3 Gay-lussac's law :

For a fixed amount of gas at constant volume, pressure of the gas is directly proportional to temperature of the gas on absolute scale of temperature.

ΡαΤ

= constant \rightarrow dependent on amount and volume of gas



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Gaseous State /

F3 $\frac{P_1}{T_1} = \frac{P_2}{T_2} \rightarrow \text{temperature on absolute scale}$

Originally, the law was developed on the centigrade scale, where it was found that pressure is a linear function of temperature $P = P_0 + bt$ where 'b' is a constant and P_0 is pressure at zero degree centigrade.





Ex.
$$PV = K \implies V = K_1/P$$

 $\frac{V}{T} = K_2 \implies V = K_2T$
 $\frac{K_1}{P} = K_2T$
 $PT = \frac{K_1}{K_2} = \text{const.} \Rightarrow P \alpha = \frac{1}{T} \Rightarrow ?$

This is wrong because we are varying temperature & $K_1 = f(1)$ thus K_1 will change according to temperature So $\frac{K_1}{K_2}$ will be a function of temp & not constant.

where are we wrong ?

Solved Example

- **Example 1.** The temperature of a certain mass of a gas is doubled. If the initially the gas is at 1 atm pressure. Find the % increase in pressure ?
- Solution:

 $\frac{P_1}{T_1} = \frac{P_2}{T_2}; \qquad \frac{1}{T} = \frac{P_2}{2T}$ % increase = $\frac{2-1}{1} \times 100 = 100\%$

Example 2. The temperature of a certain mass of a gas was increased from 27°C to 37°C at constant volume. What will be the pressure of the gas.

Solution: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$; $\frac{P}{300} = \frac{P_2}{310}$; $P_2 = \frac{31}{30}P$



Gaseous State

D4 Avogadro's Hypothesis :

For similar values of pressure & temperature equal number of molecules of different gases will occupy equal volume.

 $N_1 \longrightarrow V$ (volume of N_1 molecules at P & T of one gas)

 $N_1 \longrightarrow V$ (volume of N_1 molecules at P & T of second gas)

⇒ Molar volume & volume occupied by one mole of each and every gas under similar conditions will be equal.

One mole of any gas or a combination or gases occupies 22.413996 L of volume at STP.

The previous standard is still often used, and applies to all chemistry data more than decade old, in this definition **Standard Temperature and Pressure STP** denotes the same temperature of 0°C (273.15K), but a slightly higher pressure of 1 atm (101.325 kPa).

Standard Ambient Temperature and Pressure (SATP), conditions are also used in some scientific works. SATP conditions means 298.15 K and 1 bar (i.e. exactly 10^5 Pa) At SATP (1 bar and 298.15 K), the molar volume of an ideal gas is 24.789 L mol⁻¹ (Ref. NCERT)

Th4

1.4. Equation of State :

Combining all the gas relations in a single expression which discribes relationship between pressure, volume and temperature, of a given mass of gas we get an expression known as equation of state.

$$\frac{PV}{T} = \text{constant (dependent on moles of the gas n).}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$
Ideal gas Equation : $\frac{PV}{nT} = \text{constant}$ [universal constant]
$$= R \qquad (\text{ideal gas constant or universal gas constant})$$

$$R = 8.314 \text{ J/Kmole} \approx 25/3$$

$$= 1.987 \text{ cal/mole} \approx 2$$

$$= 0.08 \text{ Latm/mole} \approx 1/12$$

Solved Example

Example 3.

Some spherical balloons each of volume 2 litre are to be filled with hydrogen gas at one atm & 27°C from a cylinder of volume 4 litres. The pressure of the H_2 gas inside the cylinder is 20 atm at 127°C. Find number of balloons which can be filled using this cylinder. Assume that temperature of the cylinder is 27°C.

Solution: No. of moles of gas taken initially = $\frac{20 \times 4}{R \times 400}$ = 2.43 L No. of moles of gas left in cylinder = $\frac{1 \times 4}{R \times 300}$ = 0.162L No. of moles of gas to be filled in balloons = 2.43 - 0.162 = 2.268 Let we have 'n' balloons that we can fill No. of moles of gas that can be filled in 1 balloon = $\frac{1 \times 2}{0.082 \times 300}$ = 0.081 \therefore 0.081 x n = 2.268 n = 28 balloons.



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SECTION (B) : DALTONS LAW OF PARTIAL PRESSURES

Th5 1.5.



Partial pressure :

In a mixture of non reacting gases partial pressure of any component of gas is defined as pressure exerted by this component if whole of volume of mixture had been occupied by this component only.

Partial pressure of first component gas

$$P_1 = \frac{n_1 RT}{v};$$
 $P_2 = \frac{n_2 RT}{v};$ $P_3 = \frac{n_3 RT}{v}$

Total pressure = $P_1 + P_2 + P_3$.

D5 Daltons law :

For a *non reacting* gaseous mixture total pressure of the mixture is the summation of partial pressure of the different component gases.

F4
$$P_{Total} = P_1 + P_2 + P_3 = \frac{(n_1 + n_2 + n_3)RT}{V}$$

 $\frac{P_1}{P_T} = \frac{n_1}{n_T} = x_1$ (mole fraction of first component gas)

 $\frac{P_2}{P_T} = \frac{n_2}{n_T} = x_2$ (mole fraction of second component gas)

 $\frac{P_3}{P_T} = \frac{n_3}{n_T} = x_3$ (mole fraction of third component gas)

SECTION (C) : MIXING OF GASES

Solved Example

Example 1.

The stop cock connecting the two bulbs of volume 5 litre and 10 litre containing as ideal gas at 9 atm and 6 atm respectively, is opened. What is the final pressure if the temperature remains same.

Solution: After the opening of the stop cock the pressure of the each bulb will remain same.



At the beginning, the no. of moles of gas in A = $\frac{10 \times 6}{RT}$

At the beginning, the no. of moles of gas in B = $\frac{5 \times 9}{RT}$

... Total no. of mole at the beginning = $\frac{105}{RT}$ Total no. of mole of gas before opening the stop cock Total no. of moles of gas after opening stop cock = $\frac{105}{RT}$

... Pressure after the opening of the stop cock

$$P = \frac{105}{RT} \times \frac{RT}{V_{total}} = \frac{105}{10+5} = 7 \text{ atm}$$





Th6

1.6. Analysis of gaseous mixture :

D6 Vapour Density : Vapour density of any gas is defined as the density of any gas with respect to density of the H₂ gas under identical conditions of temperature T and pressure P.

$$Vapour density = \frac{density of gas at T \& P}{density of H_2 under same P \& T}$$

$$P = \frac{m}{V} \cdot \frac{RT}{M} \implies P = \rho \frac{RT}{M} \qquad \rho = \frac{PM}{RT}$$

$$Vapour density = \frac{PM_{gas}RT}{RT PM_{H_2}} = \frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2}$$

F5

Average molecular mass of gaseous mixture :

total mass of the mixture divided by total no. of moles in the mixture

 $M_{mix} = \frac{Total mass of mixture}{Total no. of moles in mixture}$

If we have ' n_1 ', ' n_2 ' and ' n_3 ' are moles of three different gases having of molar mass ' M_1 ', ' M_2 ' and ' M_3 ' respectively.

$$M_{min} = \frac{n_1 M_1 + n_2 M_2 + n_3 M_3}{n_1 + n_2 + n_3}$$

 $M_{gas} = 2 \times vapour density$

Solved Example

Example 1.

Calculate the mean molar mass of a mixture of gases having 7 g of Nitrogen, 22 g of CO_2 and 5.6 litres of CO at STP.

Solution:

Moles of
$$N_2 = 7/28 = 1/4$$

Moles of $CO_2 = 22/44 = 1/2$

Mean molar mass =
$$M_{min} = \frac{n_1M_1 + n_2M_2 + n_3M_3}{n_1 + n_2 + n_3} = (7 + 7 + 22) / 1 = 36$$

SECTION (D) : GRAHAM'S LAW OF DIFFUSION

Th7

1.7. Graham's Law of Diffusion/Effusion :

D7 Diffusion : Net spontaneous flow of gaseous molecules from region of high concentration (higher partial pressure) to the region of lower concentration or lower partial pressure

when removed



flow will be from both sides, N₂ will try to equalise its partial Pressure in both the vessels, and so will O₂.

$$P_{N_2} = 2.5 \text{ atm}$$
 $P_{N_2} = 2.5 \text{ atm}$
 $P_{O_2} = 1.5 \text{ atm}$
 $P_{O_2} = 1.5 \text{ atm}$



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Gaseous State

Graham's Law :

"Under similar conditions of temperature and pressure (partial pressure) the rate of diffusion of different gases is inversely proportional to square root of the density of different gases."

d = density of gas

dx

dt

F7

Rate of diffusion
$$r \propto \frac{1}{\sqrt{d}}$$
 $d = density of gas$
 $\frac{r_1}{r_2} = \frac{\sqrt{d_2}}{\sqrt{d_1}} = \frac{\sqrt{M_2}}{\sqrt{M_1}} = \sqrt{\frac{V.D_2}{V.D_1}}$ V.D is vapour density
 $r = volume flow rate = \frac{dV_{out}}{dt}$
 $r = mole flow rate = \frac{dn_{out}}{dt}$
 $r = distance travelled by gaseous molecule per unit time =$

The general form of the grahams law of diffusion can be stated as follows, when one or all of the parameters are varied.

rate
$$\propto \frac{P}{\sqrt{TM}} A$$
; P – Pressure, A – area of hole, T – Temp., M – mol. wt.

If partial pressure of gases are not equal. Then rate of diffusion is found to be proportional to partial pressure & inversely proportional to square root of molecular mass.

$$\label{eq:radius} \begin{array}{c} r \propto P \\ r \propto \frac{1}{\sqrt{M}} \qquad ; \qquad \qquad \frac{r_1}{r_2} \, = \frac{P_1}{P_2} \, \sqrt{\frac{M_2}{M_1}} \end{array}$$

Selective diffusion :

If one or more than one components of a mixture are allowed to diffuse and others are not allowed then it is selective diffusion of those components.



P Platinum allows only H₂ gas to pass through

Effusion : (forced diffusion) a gas is made to diffuse through a hole by the application of external pressure.





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Solved Example

In a tube of length 5 m having 2 identical holes at the opposite ends. H₂ & O₂ are made to Example 1. effuse into the tube from opposite ends under identical conditions. Find the point where gases will meet for the first time.

	H_2 O_2
Solution:	$\frac{r_1}{r_2} = \frac{ax}{dt} \times \frac{dt}{dx} = \sqrt{\frac{m_2}{m_1}} = \frac{dx_1}{dx_2} = \sqrt{\frac{32}{2}}$
	$\frac{dx_1}{dx_2} = 4 \qquad \Rightarrow \qquad \frac{distance \ travelled \ by \ H_2}{distance \ travelled \ by \ O_2} = 4$
	$\frac{x}{(5-x)} = 4 \qquad \Rightarrow \qquad x = (5-x) \ 4 \qquad \Rightarrow \qquad x = 20 - 4x$
	$5x = 20$ \Rightarrow $x = 4$ from H ₂ side
Example 2.	Assume that you have a sample of hydrogen gas containing H ₂ , HD and D ₂ that you want to separate into pure components (H = ¹ H and D = ² H). What are the relative rates of diffusion of the three molecules according to Graham's law ?
Solution:	Since D_2 is the heaviest of the three molecules, it will diffuse most slowly, and let we call its relative rate 1.00. We can then compare HD and H_2 with D_2 . Comparing HD with D_2 , we have
	$\frac{\text{Rate of HD diffusion}}{\text{Rate of D}_2 \text{ diffusion}} = \sqrt{\frac{\text{Molecular mass of D}_2}{\text{Molecular mass of HD}}} = \sqrt{\frac{4.0 \text{ amu}}{3.0 \text{ amu}}} = 1.15$

Comparing H_2 with D_2 we have

$$\frac{\text{Rate of H}_2 \text{ diffusion}}{\text{Mass of D}_2} = \frac{4.0}{4.0}$$

Rate of D_2 diffusion \sqrt{Mass} of H_2 1√2.0 amu Thus, the relative rates of diffusion are $H_2(1.41) > HD(1.15) > D_2(1.00)$.

SECTION (E) : KINETIC THEORY OF GASES

Th8

1.8. Kinetic Theory of Gases :

Postulates / assumptions of KTG :

A gas consists of tiny spherical particles called molecules of the gas which are identical in shape & size (mass)

amu = 1.41

The volume occupied by the molecules is negligible in comparision to the total volume of the gas.

For an ideal gas, volume of the ideal gas molecule ~ 0.

- Gaseous molecules are always in random motion and collide with other gaseous molecules & with the walls of the container.
- Pressure of the gas is due to these molecular collisions among themselves and with walls of the container
- These collisions are elastic in nature
- Molecular attraction forces are negligible. Infact, for an ideal gas attractive or repulsive forces are equal to zero.
- Newton's laws of motion are applicable on the motion of the gaseous molecules.
- Effect of gravity is negligible on molecular motion.
- The average K.E. of gaseous molecules is proportional to the absolute temperature of the gas.

 $\frac{1}{2}M(\overline{u^2}) \propto T$ (bar is for average)

Kinetic equation of gaseous state (expression for pressure of gas).



Der.1 m = mass of one molecule

 $\vec{u} = u_x \hat{i} + u_y \hat{j} + u_z \hat{k}$



Consider collision with face ABCD

inital $\vec{P}_i = mu_x \hat{i}$; final $\vec{P}_f = -mu_x \hat{i}$

change in momentum due to collision $= 2 U_x m$

time taken between two successive collision with face ABCD = t = $\frac{2\ell}{\mu_x}$

frequency of collisions (f) = $\frac{1}{t} = \frac{U_x}{2\ell}$

change in momentum in one sec. = force = 2 m $\frac{U_x \times U_x}{2\ell} = \frac{2mU_x^2}{\ell}$

force due to all the molecules = $\frac{m}{\ell} \{ U_{x_1}^2 + U_{x_2}^2 + \dots + U_{x_N}^2 \}$

average value of
$$U_N^2 = \overline{U_N^2} = \frac{Ux_1^2 + U_{x_2}^2 + \dots Ux_{x_N}^2}{N}$$

$$F_x = \frac{m}{\ell} \{ N \quad \overline{U_x^2} \}$$

all the three directions are equal as the motion is totally random in all directions, hence

$$U_x^2 = U_y^2 = U_z^2$$

$$\overline{U}^2 = \frac{U_1^2 + U_2^2 + U_3^2 + \dots + U_N^2}{N} = \overline{U}_x^2 + \overline{U}_y^2 + \overline{U}_z^2 = 3\overline{U}_x^2$$

$$F = \frac{m}{\ell} \cdot N \frac{1}{3} \overline{U}^2$$

$$Pressure = \frac{F_x}{\ell^2} = \frac{1}{3} \frac{mN}{\ell^3} \overline{U}^2 \qquad \text{The volume of the container 'V'} = \ell^3$$

$$\therefore \qquad PV = \frac{1}{3} mN\overline{U}^2 \quad \text{Kinetic equation of gases}$$

where $\overline{U^2}$ is mean square speed

root mean square speed =
$$U_{rms} = \sqrt{U^2} = \sqrt{\left(\frac{U_1^2 + U_2^2 + U_3^2 + \dots + U_N^2}{N}\right)}$$



八

Th9

1.9. Verification of Gaseous Laws Using Kinetic Equation :

From postulates ; $PV = \frac{1}{2} mN \overline{U^2}$ **F8** $\frac{1}{2} m \overline{U^2} \propto T = \lambda T$ Where ' λ ' is a proportionality constant $PV = \frac{2}{3} \left(\frac{1}{2} m \overline{U^2} \right) N$; $PV = \frac{2}{3} \lambda NT$ (N = Total number of molecules) N : constant Boyle's Law : T : constant PV = constantCharles law : N : constant P: constant $T \propto V$ Kinetic energy of gaseous molecule (translational K.E.) To calculate λ we have to use ideal gas equation (experimental equation) PV = nRTkinetic equation PV = nRT = $\frac{2}{3}\lambda$ NT = $\frac{2}{3}\lambda$ (nN_A) T (n = number of moles of gas) on comparing $\lambda = \frac{3}{2} \times \frac{R}{N_{A}}$ $\lambda = \frac{3}{2}$ K where K = $\frac{R}{N_{A}}$ = Boltzmann constant Average K.E. of molecules = $\frac{1}{2} m \overline{U^2} = \lambda T$ F9 Average K.E. = $\frac{3}{2}$ KT (only dependent on temperature not on nature of the gas.) Average K.E. for one mole = $N_A \left(\frac{1}{2}m\overline{U^2}\right) = \frac{3}{2}KN_AT = \frac{3}{2}RT$ F10 Root mean square speed $U_{rms} = \sqrt{U^2} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{mN_A}}$ Where m-mass of one molecule Dependent on nature of gas i.e. mass of the gas 0 $U_{rms} = \sqrt{\frac{3RT}{M}}$ F11 M = molar mass Average speed : $U_{av} = U_1 + U_2 + U_3 + \dots + U_N$ $U_{av} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{\pi m}}$ F12 K is Boltzmman constant Most probable speed: The speed possessed by maximum number of molecules at the given temperature $U_{mp} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2KT}{m}}$ F13 Velocity can be decribed by maximum number of molecule. Molecular speed Most probable speed Average speed Root mean square $\sqrt{\frac{2KT}{m}} = \sqrt{\frac{2RT}{M}} \qquad \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{m}} \qquad \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3KT}{m}}$ U_{mp} U_{av} $U_{ms} = \sqrt{2}$ $\sqrt{\frac{8}{\pi}}$ $\sqrt{3}$



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Solved Example

Example 1.

In a container of capacity 1 litre there are 10^{23} molecules each of mass 10^{-22} gm. If root mean square speed is 10^5 cm/sec then calculate pressure of the gas.

Solution:

$$PV = \frac{1}{3} \text{ mN } U_{\text{rms}}^{2}$$

$$P = ? ; V = 10^{-3} \text{ m}^{3}$$

$$m = 10^{-25} \text{ kg} ; N = 10^{23}$$

$$\sqrt{U^{2}} = 10^{5} \text{ cm/sec} = 10^{3} \text{ m/sec}$$

$$U^{2} = 10^{6} \text{ m}^{2}/\text{sec}^{2}$$

$$P \times 10^{-3} = \frac{1}{3} \times 10^{-25} \times 10^{23} \times 10^{6}$$

$$P = \frac{1}{3} \times 10^{-2} \times 10^{6} \times 10^{3} ; P = \frac{1}{3} \times 10^{7} \text{ pascals}$$

Th10

1.10. Maxwell's distributions of molecular speeds :

Postulates/Assumptions of speed distributions

- It is based upon theory of probability.
- It gives the statistical averages of the speed of the whole collection of gas molecules.
- Speed of gaseous molecules of may vary from 0 to ∞. The maxwell distribution of speed can be plotted against fraction of molecules as follows.



- The area under the curve will denote fraction of molecules having speeds between zero and infinity
- Total area under the curve will be constant and will be unity at all temperatures.
- Area under the curve between zero and u₁ will give fraction of molecules racing speed between 0 to u₁. This fraction is more at T₁ and is less at T₂.
- The peak corresponds to most probable speed.
- At higher temperature, fraction of molecules having speed less than a particular value decreases.
- For Gases with different molar masses will have following graph at a given temperature.





SECTION (F) : EUDIOMETRY

Th11

1.11. Eudiometry :

The analysis of gaseous mixtures is called eudiometry. The gases are identified by absorbing them in specified and specific reagents.

Some Common Facts :

- Liquids and solutions can absorb gases.
- If a hydrocarbon is burnt, gases liberated will be CO₂ & H₂O. [H₂O is seperated out by cooling the mixture & CO₂ by absorption by aqueous KOH]
- If organic compound contains S or P, then these are converted into SO₂ & P₄O₁₀ by burning the organic compound.
- If nitrogen is present, then it is converted into N₂.
- [The only exception : if organic compound contains NO2 group then NO2 is liberated]
- If mixture contains N₂ gas & this is exploded with O₂ gas, do not assume any oxide formation unless specified.
- Ozone is absorbed in turpentine oil and oxygen in alkaline pyragallol.

Solved Example

Example 1. Carbon dioxide gas (CO₂) measuring 1 litre is passed over heated coke the total volume of the gases coming out becomes 1.6 litre. Find % conversion of CO₂ into carbon monoxide.

Solution:

$CO_2 + C \longrightarrow$	2CO		CO ₂	CO
1	0	at time t	1 – x	2x
Initial volume =	1 litre			
final volume =	1.6 litre		Final volume =	(1 + x) litres
1 + x = 1.6				
x = 0.6				
x = 0.6				
$\frac{0.6}{1} \times 10 = 60\%$	% of CO	2 will be converte	ed into CO	

- **Example 2.** 100 ml of hydrocarbon is mixed with excess of oxygen and exploded. On cooling, the mixture was reported to have a contraction of 250 ml. The remaining gas when passed through a solution of aqueous KOH, the mixture shows a further contraction of 300 ml. Find molecular formula of the hydrocarbon.
- Solution: $C_x H_y + \left(x + \frac{Y}{4}\right) O_2 \longrightarrow x CO_2 + \frac{y}{2} H_2O$ 100 mlx 100 $\frac{y}{2} 100.$ mixture contains CO_2 , H_2O & excess O_2 on cooling, H_2O is separated, volume of $H_2O = 250$ ml $\frac{y}{2} \times 100 = 250$; y = 5,as KOH absorbs CO_2 , hence 300 ml contraction is because of CO_2 that has been absorbed.Volume of $CO_2 = 100 \times = 300$; x = 3Empirical formula = $C_3 H_5$; molecular formula = $C_6 H_{10}$.
- **Note:** If water is already condensed out then total contraction in reaction mixture = {volume of reactants} {volume of products + volume of unused species excluding volume of H₂O}.





Example 3. 100 ml of an hydrocarbon is burnt in excess of oxygen in conditions so that water formed gets condensed out the total contraction in volume of reaction mixture was found to be 250 ml when the reaction mixture is further exposed to aqueous KOH a further contraction of 300 ml is observed find molecular formula of hydrocarbon.

Solution:
$$C_xH_y + O_2 \longrightarrow CO_2 + H_2O$$
100 mlexcess300 mlBy POAC on 'C' atoms
 $x \times 100 = 300$
 $x = 3$ $x \times 100 = 300$
 $x = 3$ POAC on 'H' atoms
 $y \times 100 = 2 \times$ moles of H_2O
POAC on O atoms
 $2 \times v = 2 \times 300 + 1 \times H_2O$
 $2 \times v = 600 + 50 y$
 $v = \frac{600 + 50 y}{2}$ volume of O_2 consumed
The total volume contraction is 250 ml.
Hence,
 $100 + V - 300 = 250$
 $-200 + V = 250$
 $2 \times 450 - 600 = 50 y$
 $\frac{300}{50} = y = 6$
Hydro carbon will be C₃ H₆

Alternative :

Using balanced chemical equation

	C _x H _y	+ $\left(\frac{x+y}{4}\right)O_2 \longrightarrow$	xCO ₂	+	<u>y</u> 2 H₂O
t = 0	100 ml	V	0		0
t	0	$V - 100 \left(x + \frac{y}{4} \right)$	100 x ml		<u>100 y</u> 2

volume remained

$$V-100\left(x+\frac{y}{4}\right) + 100x + 50y$$

-100 - V = 250
-25y + 50y = 150
25y = 150
y = 6



Gaseous State

SUMMARY

A gas is a collection of atoms or molecules moving independently through a volume that is largely empty space. Collisions of the randomly moving particles with the walls of their container exert a force per unit area that we perceive as pressure. The SI unit for pressure is the pascal, but the atmosphere and the millimeter of mercury are more commonly used. The physical condition of any gas is defined by four variables; pressure (P), temperature (T), volume (V) and molar amount (n). The specific relationship among these variables are called the gas laws :

Boyle's law :	The volume of a gas varies inversely with its pressure. That is, V \propto 1/P or
	PV = k at constant n. T.
Charles Low:	The volume of a gas varias directly with its Kalvin, temperature. That is

Charles Law :	The volume of a gas varies directly with its Kelvin temperature. That i	is,
	$V \propto T$ or V/T = k at constant n, P	

Avogadro's Law : The volume of a gas varies directly with its molar amount. That is, $V \propto n$ or V/n = k at constant T,P.

The three individual gas laws can be combined into a single ideal gas law, PV = nRT. If any three of the four variables P, V, T and n are known, the fourth can be calculated. The constant R in the equation is called the gas constant and has the same value for all gases. At standard temperature and pressure (STP; 1 atm and 0°C), the standard molar volume of an ideal gas is 22.414 L.

The gas laws apply to the mixture of gases as well as to pure gases According to Dalton's law of partial pressures, the total pressure exerted by a mixture of gases in a container is equal to the sum of the partial pressure of each individual gas would exert alone.

The behaviour of gases can be accounted for using a model called the kinetic-molecular theory, a group of five postulates:

- 1. A gas consists of tiny particles moving at random.
- 2. The volume of the gas particles is negligible as compared with the total volume.
- 3. There are no forces between particles, either attractive or repulsive.
- 4. Collisions of gas particles are elastic.

5. The average kinetic energy of gas particles is proportional to their absolute temperature.

The connection between temperature and kinetic energy obtained from the kinetic- molecular theory makes it possible to calculate the average speed of a gas particle at any temperature. An important practical consequence of this relationship is Graham' law, which states that the rate of a gas effusion or spontaneous passage through a pinhole in a membrane depends inversely on the square root of the molar mass of gas.

Real gases differ in their behaviour from that predicted by the ideal gas law, particularly at higher pressure, where gas particles are forced close together and intermolecular forces become significant.







MISCELLANEOUS SOLVED PROBLEMS (MSPS)

Problem 1. The diameter of a bubble at the surface of a lake is 4 mm and at the bottom of the lake is 1 mm. If atmospheric pressure is 1 atm and the temperature of the lake water and the atmosphere are equal. What is the depth of the lake ? (The density of the lake water and mercury are 1 g/ml and 13.6 g/ml respectively. Also neglect the contribution of the pressure due to surface tension) Solution: $P_1V_1 = P_2V_2$:. (760 mm × 13.6 × g) $\frac{4}{3}\pi$ (4 mm/2)³ = (760 mm × 13.6 × g + h × 1 × g) $\frac{4}{3}\pi$ (1 mm/2)³ $760 \times 13.6 \times 64 = (760 \times 13.6 + h)$ $h = 64 \times 760 \times 13.6 - 760 \times 13.6$ $h = 63 \times 760 \times 13.6 \text{ mm}$ $h = \frac{63 \times 760 \times 13.6}{1000 \times 1000} \text{ km} = 0.6511 \text{ km} = 651.1 \text{ m Ans.}$ Problem 2. A gas is initially at 1 atm pressure. To compress it to 1/4 th of initial volume, what will be the pressure required? Solution: $P_1 = 1$ atm $V_1 = V$ $P_2 = ?$ $V_2 = V/4$ $P_1V_1 = P_2V_2$ at const. T & n $P_2 = \frac{P_1V_1}{V_2} = \frac{1 \text{ atm} \times V}{V/4} = 4 \text{ atm Ans.}$ case II case I Problem 3. A gas column is trapped between closed end of a tube and a mercury column of length (h) when this tube is placed with its open end upwards the length of gas column is (ℓ_1) , the length of h gas column becomes (ℓ_2) when open end of tube is held downwards. Find atmospheric pressure in terms of height of Hg column. $P_1 = (P_0 + h)$ Solution: $P_2 = (P_0 - h)$ For gas $V_1 = \pi r^2 \ell_1$ $V_2 = \pi \mathbf{r}^2 \ell_2$ at const T. and moles. $P_1V_1 = P_2V_2$; (P₀ + h) $\pi r^2 \ell_1 = (P_0 - h) \pi r^2 \ell_2$ $P_0\ell_2 + h\ell_1 = P_0\ell_2 - h\ell_2$ $P_0\ell_2 - P_0\ell_1 = h\ell_1 + h\ell_2$ $P_0 = \frac{h(\ell_1 + \ell_2)}{(\ell_2 - \ell_1)} \text{ cm of Hg column Ans.}$ Problem 4. If water is used in place of mercury then what should be minimum length of Barometer tube to measure normal atmospheric pressure. Solution: $P_{H_{a}} = P_{H_{2}O} = P_{atm}$. $0.76 \text{ m} \times 13.6 \times \text{g} = \text{h}_{H_{20}} \times 1 \times \text{g};$ $h_{H_{AO}} = 0.76 \times 13.6 = 10.336 \text{ m Ans.}$



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Gaseous State Problem 5. A tube of length 50 cm is initially in open atmosphere at a pressure 75 cm of Hg. This tube is dipped in a Hg container ろ upto half of its length. Find the level of mercury column in side the tube. (x-火)↓ 1/2 Solution: If after dipping the tube, the length of air column be x cm (situation shown in the adjoining figure). Then by using, $P_iV_i = P_fV_f$ We have, 75 cm Hg × ℓ A = P_f × x × A (1) (ℓ = 50 cm) & also, $P_f = 75 \text{ cm Hg} + (x - \frac{\ell}{2})$ (2) $\begin{array}{rcl} (2) \& (1) & \Rightarrow & [75 + (x - 25)] \times x = 75 \times 50 \\ & \Rightarrow & x^2 + 50 \times - 3750 = 0 \\ & \therefore & x = 41.14 & \text{or} & -91 \\ \text{But,} & x \text{ can't } be - ve & \therefore & x = 41.14 \end{array}$ -91.14Mercury column inside the tube = (50 - 41.14) cm = 8.86 cm Ans. *.*.. Problem 6. An open container of volume V contains air at temperature 27°C or 300 K.The container is heated to such a temperature so that amount of gas coming out is 2/3 of (a) amount of gas initially present in the container. (b) amount of gas finally remaining in the container. Find the temperature to which the container should be heated. Solution: (a) Here, P & V are constant, n & T are changing. Let, initially the amount of gas present be n & temp is 27°C or 300K. Finally amount of gas present in container = $n - \frac{2}{3}n = \left(\frac{1}{3} \times n\right)$ & final temperature be T. Then using $n_1T_1 = n_2T_2$, we have, $n \times 300 = \frac{n}{2} \times T_2 \Rightarrow T_2 = 900$ K i.e., final temp = 900K Ans. Let there be x moles of gas remaining in the container, $\frac{2}{3}$ of x come out (b) $\left(\frac{2}{3}x+x\right)=n \Rightarrow \frac{5x}{3}=n \therefore x=\frac{3n}{5}$ *.*.. $n \times 300 \text{ K} = \frac{3n}{5} \times T_2$ Using $n_1 T_1 = n_2 T_2$ ÷. $T_2 = 500K$ ÷. Final temperature = 500 K Ans. Find the lifting power of a 100 litre balloon filled with He at 730 mm and 25°C. (Density of air = Problem 7. 1.25 g /L). Solution: Since, PV = nRT $PV = \frac{W}{M}RT \qquad \therefore \qquad W = \frac{PVM}{RT} = \frac{730}{760} \times \frac{100 \times 4}{0.082 \times 298}g$ i.e., Wt. of He = 15.72 g Wt. of air displaced = 100×1.25 g/L = 125 g Lifting power of the balloon = 125 g - 15.72 g = 109.28 g Ans. *.*.. Problem 8. A weather balloon filled with hydrogen at 1 atm and 300 K has volume equal to 12000 litres. On ascending it reaches a place where temperature is 250 K and pressure is 0.5 atm. The volume of the balloon is : (A) 24000 litres (B) 20000 litres (C) 10000 litres Using $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$; $\frac{1 \text{ atm } \times 12000 \text{ L}}{300 \text{ K}} = \frac{0.5 \text{ atm } \times V_2}{250 \text{ K}}$ (D) 12000 litres Solution: V₂ = 20,000 L Hence, Ans. (B)



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Gaseous Stat	<i>te</i>
Problem 9.	Four one litre flasks are separately filled with the gases, O ₂ , F ₂ , CH ₄ and CO ₂ under the same conditions. The ratio of number of molecules in these gases :
Solution:	(A) $2: 2: 4: 3$ (B) $1: 1: 1: 1$ (C) $1: 2: 3: 4$ (D) $2: 2: 3: 4$ According to Avogadro's hypothesis.All the flasks contains same no. of molecules \therefore Ratio of no. of molecules of O_2 , F_2 , $CH_4 & CO_2$ $= 1: 1: 1: 1$ Hence, (B)
Problem 10.	A sample of water gas has a composition by volume of 50% H ₂ , 45% CO and 5% CO ₂ . Calculate the volume in litres at STP at which water gas which on treatment with excess of steam will produce 5 litre of H ₂ . The equation for the reaction is: $CO + H_2O \longrightarrow CO_2 + H_2$
Solution:	If x L CO in needed then volume of H ₂ in water gas = $\left(\frac{x}{0.45} \times 50\%\right) L = \left(\frac{x}{0.45} \times \frac{1}{2}\right) L = \frac{x}{0.9} L$
	But, from equation : $CO + H_2O \longrightarrow CO_2 + H_2$ & Gay-Lussac's law, we get, that the volume of H_2 Produced = volume of CO taken. \therefore Volume of H_2 due to reaction = x L \therefore Total volume of $H_2 = \left(\frac{x}{0.9} + x\right)L = 5 L$ 1.9x
	$\Rightarrow \frac{1}{0.9} = 5 L$
	:. $x = \frac{0.9 \times 5}{1.9}$:: Volume of water gas $= \frac{x}{0.45} L = \frac{0.9 \times 5}{1.9 \times 0.45} L = 5.263 L$ Ans.
Problem 11.	The partial pressure of hydrogen in a flask containing two grams of hydrogen and 32 gm of sulphur dioxide is :(A) 1/16th of the total pressure(B) 1/9th of the total pressure(C) 2/3 of the total pressure(D) 1/8th of the total pressure
Solution:	$n_{H_2} = \frac{2g}{2g/mol} = 1 \text{ mol.}$; $n_{SO_2} = \frac{32g}{64g/mol} = 0.5 \text{ mol}$
	$\therefore \qquad P_{H_2} = \frac{n_{H_2}}{\left(n_{H_2} + n_{SO_2}\right)} \times P_{T} = \frac{1}{(1+0.5)} \times P_{T} = \frac{2}{3} P_{T}. \text{Hence, Ans. (C)}$
Problem 12.	Equal volume of two gases which do not react together are enclosed in separate vessels. Their pressures are 10 mm and 400 mm respectively. If the two vessels are joined together, then what will be the pressure of the resulting mixture (temperature remaining constant) : (A) 120 mm (B) 500 mm (C) 1000 mm (D) 205 mm
Solution:	Let, vol of containers be V & temperature be T
	P ₁ = 10mm P ₂ = 400mm ∴ $n_1 = \frac{P_1 V}{RT}$ & $n_2 = \frac{P_2 V}{RT}$
	$\therefore \qquad n_1 + n_2 = \frac{(P_1 + P_2) \times V}{PT}$
	After joining two containers final vol = $(V + V) = 2V$ (for gases)
	$P_{\text{final}} = \frac{(n_1 + n_2)RT}{V_{\text{final}}} = \frac{(P_1 + P_2) \times V}{RT} \times \frac{RT}{2V} = \frac{(P_1 + P_2)}{2} = \frac{(10 + 400)mm}{2} = 205 \text{ mm}.$
	Hence, Ans. (D)
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Gaseous State		——————————————————————————————————————
Problem 13.	5 ml of H₂ gas diffu the same hole unde	ses out in 1 sec from a hole. Find the volume of O_2 that will diffuse out from r identical conditions in 2 sec.
Solution:	Rate of diffusion of	$H_2 = \frac{5mI}{1 \text{ sec}} = 5mI/s = r_{H_2}$ (say)
	$\therefore \mathbf{r}_{O_2} = \mathbf{r}_{H_2} \times \frac{1}{4} = \$$	$5 \text{ml/s} \times \frac{1}{4}$
	∴ Volume of O₂ di	ffused in 2.0 seconds = $\frac{5}{4} \times 2$ ml = 2.5 ml Ans.
Problem 14.	A vessel contains allowed to diffuse the	H_2 & O_2 in the molar ratio of 8 : 1 respectively. This mixture of gases is arough a hole, find composition of the mixture coming out of the hole.
Solution:	Here, n_{H_2} : n_{O_2} = 8	$: 1 \& \qquad \frac{I_{H_2}}{I_{O_2}} = \frac{I_{H_2}}{I_{O_2}} \sqrt{\frac{I_{O_2}}{M_{H_2}}}$
	$\Rightarrow \qquad \frac{r_{_{H_2}}}{r_{_{O_2}}} = \frac{8}{1} \times$	$\sqrt{\frac{32}{2}} = \frac{32}{1} \qquad \Rightarrow \qquad \frac{(\text{no. of moles of H}_2 \text{ coming out})/\Delta t}{(\text{no. of moles of O}_2 \text{ coming out})/\Delta t} = \frac{32}{1}$
	Required composit	ion of H_2 : O_2 coming out = 32 : 1 Ans.
Problem 15.	If for two gases of r which property has (A) Density	nolecular weights M_A and M_B at temperature T_A and T_B ; $T_AM_B = T_BM_A$, then the same magnitude for both the gases. (B) Pressure (C) KE per mol (D) RMS speed
Solution:	Given that $T_A M_B =$	$T_B M_A \implies \frac{T_A}{M_A} = \frac{T_B}{M_B}$
	But, r.m.s. = $\sqrt{\frac{3RT}{M}}$	$\frac{1}{3RT_{B}}$
	$r.m.s_A = \sqrt{M_A}$ $r.m.s_A = r.m.s_B$ An	s. (D)
Problem 16.	It has been conside	red that during the formation of earth H_2 gas was available at the earth. But
	due to the excessive earth during its form	re heat on the earth this had been escaped. What was the temperature of nation? (The escape velocity is 1.1×10^6 cm/s)
Solution:	Escape velocity of I ∴ Avg velocity	H_2 should be equal to average velocity of H_2 . / of $H_2 = 1.1 \times 10^6$ cm/s = 1.1 × 10 ⁴ m/s
	But, avg. velocity =	$\sqrt{\frac{8RT}{\pi M}} \qquad \Rightarrow 1.1 \times 10^4 = \sqrt{\frac{8 \times 8.314 \times T}{\pi \times 2 \times 10^{-3}}} \qquad \left(M_{H_2} = 2g = 2 \times 10^{-3} \text{kg}\right)$
	$\therefore \qquad T = \frac{(1.1 \times 10^{\circ})}{10^{\circ}}$	$\frac{4}{3} \times \pi \times 2 \times 10^{-3}$ K = 11430.5 K = 11157.5°C Ans.
Problem 17.	A gaseous mixture mixture is exploded observed & when to observed. Find the	containing CO, methane $CH_4 \& N_2$ gas has total volume of 40 ml. This d with excess of oxygen on cooling this mixture a contraction of 30 ml is his mixture is exposed to aqueous KOH a further contraction of 30 ml is composition of the mixture
Solution:	Let vol of CO be x r On explosion with e	nL, vol of CH ₄ be y mL, vol of N ₂ be z mL. xcess of oxygen the following reactions takes place
	$CO(g) + \frac{1}{2}$	$D_2(g) \longrightarrow CO_2(g)$ (By Gaylussac's law of combing volume)
	x mL	x mL
	CH₄(g) + 20 y mL	$\begin{array}{c} D_2(g) \longrightarrow CO_2(g) + 2H_2O(g) \\ & y \text{ mL} 2 \text{ ymL} \end{array}$
	N_2 remains unreact	ed liguifies bence volume reduction of 30 mL is observed
	2y = 30	y = 15
	This is absorbed in	III = (x + y) mL KOH & vol reduction of 30 mL is observed.
	x + y = 30	$\Rightarrow \qquad x = 30 - y = (30 - 15) = 15$
	Composition of mix	$\frac{1}{2} = \frac{1}{2} = \frac{1}{2} = \frac{1}{2} = \frac{1}{2} = \frac{1}{10} = 1$
	vol. of CO = 15 mL	, VOI OF $CH_4 = 15 \text{ mL}$, VOI OF $N_2 = 10 \text{ mL}$ Ans.
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CHECK LIST

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F-4.	$P_{Total} = P_1 + P_2 + P_3$	
F-5.	$M_{gas} = 2 \times vapour density$	

F-6. $\frac{r_1}{r_2} = \frac{\sqrt{d_2}}{\sqrt{d_1}} = \frac{\sqrt{M_2}}{\sqrt{M_1}} = \sqrt{\frac{V.D_2}{V.D_1}}$ F-7. Rate $\propto \frac{P}{\sqrt{TM}} A$ F-8. $PV = \frac{1}{3} mN \overline{U^2}$

- **F-9.** Average K.E. of molecules = $\frac{3}{2}$ K T
- **F-10.** $N_A \left(\frac{1}{2}m \quad \overline{U^2}\right) = \frac{3}{2}KN_AT = \frac{3}{2}RT$
- **F-11.** $U_{rms} = \sqrt{\frac{3RT}{M}}$
- **F-12.** $U_{av} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{\pi m}}$

F-13.
$$U_{MPS} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2KT}{m}}$$

Table (Tab.)

Tab-1. Units of V, P & T

Derivation (Deri.)

Derivation of Kinetic Gas Equation Deri-1.



Reg. & Corp. Office : CG Tower, A-46 & 52, IPIA, Near City Mall, Jhalawar Road, Kota (Raj.) – 324005 Website : www.resonance.ac.in | E-mail : contact@resonance.ac.in ADVGST - 24 Toll Free : 1800 258 5555 | CIN : U80302RJ2007PLC024029