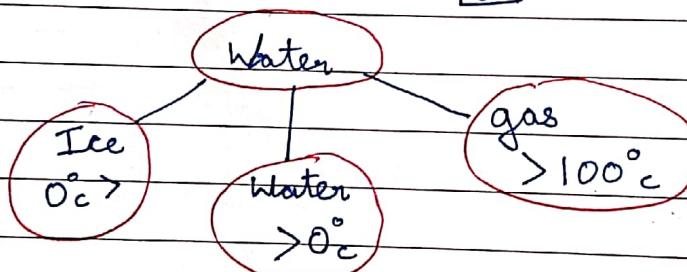
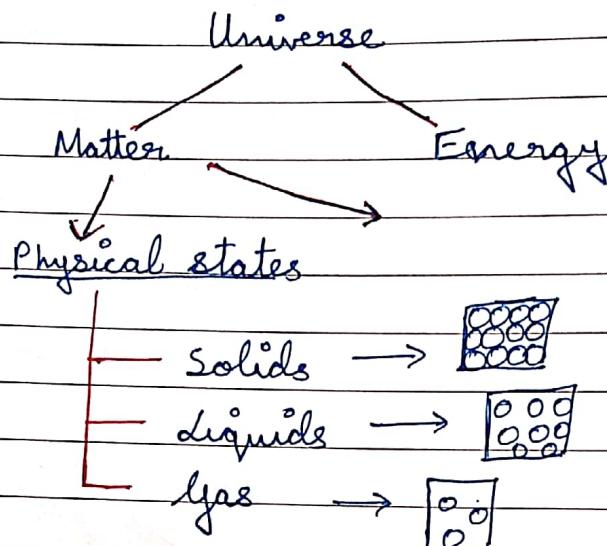


STATES OF MATTER

Gaseous States



Solids	Liquids	Gases
① The interparticle forces are maximum	The interparticle forces are weaker	The interparticle forces are minimum
② They have definite shape and volume	They have definite volume but no shape	They don't have shape and no volume
③ Solids are rigid	Fluid	Fluid
④ K.E minimum	K.E is intermediate	K.E is maximum
⑤ Density is max	Density is ..	Density is minimum

(i) Compressibility is min.

Compressibility is intermediate

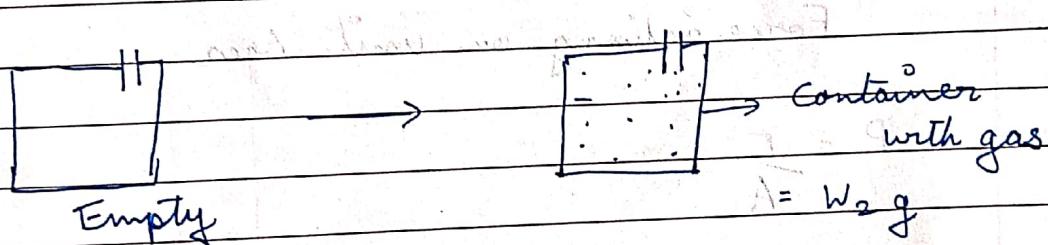
Compressibility is max.

* In periodic table there are 11 gases

(ii) $H_2, He, O_2, N_2, F_2, Cl_2, Ne, Ar, Kr, Xe, Rn$

* Measurable properties of gas:-

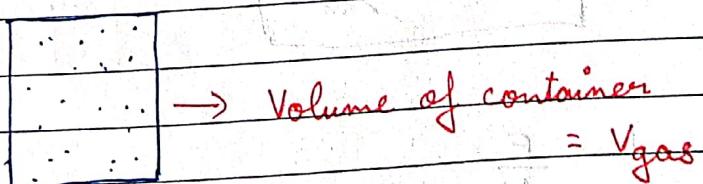
① Weight of gas:



$$\text{Weight of gas} = (W_2 - W_1) \text{ g} = W_{\text{gas}}$$

$$\text{No. of moles} = \frac{(W_2 - W_1)}{M \cdot W} = \frac{W_{\text{gas}}}{M \cdot W_{\text{gas}}}$$

② Volume of gas



③ Temperature (degree of hotness)

Temperature mainly measured in $^{\circ}\text{C}$ and K

$$T \text{ } ^{\circ}\text{K} = \text{ } ^{\circ}\text{C} + 273$$

$$0 \text{ } ^{\circ}\text{C}$$

$$= 273 + 0$$

$$= 273 \text{ K}$$

④ Pressure of the gas

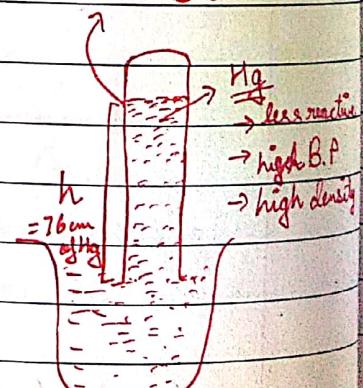
Force acting on unit Area

$$P = \frac{F}{A}$$

Area of cross section = A

$$F = m \times a = \frac{m \times v}{t} = \frac{m \times d}{t \times t}$$

$$= \text{kg. m. sec}^{-2} = \text{1 N}$$



$$P = \frac{1 \text{ N}}{\text{m}^2} = 1 \text{ Pa}$$

(Pascal)

$$1 \text{ Pa} = 1 \text{ N/m}^2$$

$$P = \frac{F}{A}$$

$$F = P \times A$$

$$mg = p \times A$$

$$\begin{aligned} m &= d \times v \\ v &= A \times h \end{aligned}$$

$$d \times v \times g = p \times A$$

$$p \times A = d \times A \times g \times h$$

$$p = hdg$$

$h = 76 \text{ cm of Hg}$

$d_{Hg} = 13.6 \text{ g/ml}$

$g = 9.8 \text{ m/s}^2$

$$p = 76 \times 10^{-2} \text{ m} \times \frac{13.6 \times 10^{-3} \text{ kg}}{10^{-6} \text{ m}^2} \times 9.8 \text{ m/s}^2$$

$$p = 101325 \text{ N/m}^2$$

$1 \text{ N/m}^2 = 1 \text{ pascal}$

$$= 101325 \text{ Pa}$$

$$1 \text{ atm} = 76 \text{ cm of Hg} = 760 \text{ mm of Hg} = 101325 \text{ Pa}$$

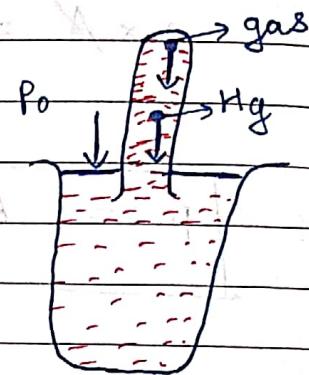
$$= 101325 \text{ N/m}^2 = 101.325 \text{ kPa} = 1.01325 \text{ bar}$$

$$= 760 \text{ torr}$$

Barometer and manometer are used

→ Faulty Barometer

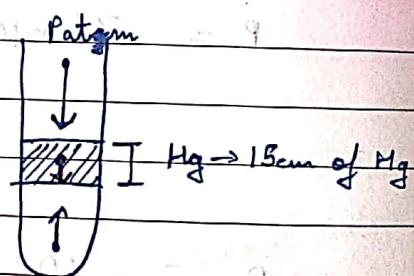
→ This barometer mainly used to measure the pressure of the gas



$$P_0 = P_{\text{gas}} + P_{\text{Hg}}$$

~~$P_{\text{gas}} = P_0 - P_{\text{Hg}}$~~

Vertically downward Upward



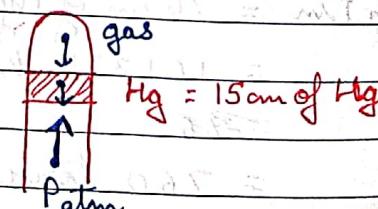
$$P_{\text{gas}} = P_{\text{atm}} + P_{\text{Hg}}$$

$$= 76 \text{ cm} + 15 \text{ cm}$$

$$P_{\text{gas}} = 91 \text{ cm of Hg}$$

Vertically downwards

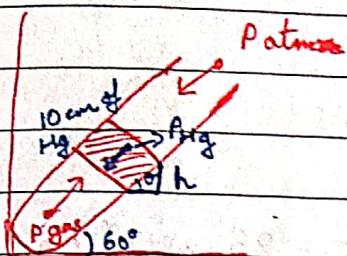
$$P_{\text{gas}} + P_{\text{Hg}} = P_{\text{atm}}$$



$$P_{\text{gas}} = 76 - 15$$

$$= 61 \text{ cm of Hg}$$

Q)



$$\sin \theta = \frac{h}{10}$$

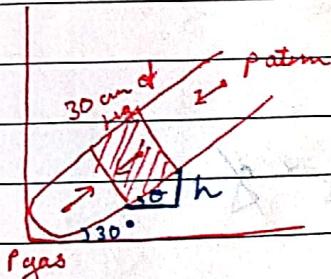
$$P_{Hg} = h = 10 \sin \theta$$

$$P_{atm} + P_{Hg} = P_{gas}$$

$$76 + 10 \times \frac{\sqrt{3}}{2} = P_{gas}$$

$$P_{gas} = 76 + 5\sqrt{3} \text{ cm of Hg}$$

Q)



$$\sin \theta = \frac{h}{30}$$

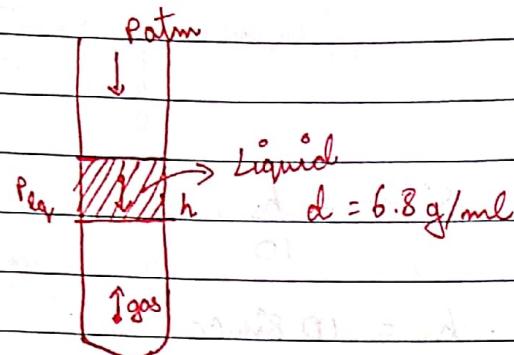
$$P_{Hg} = 30 \sin 30^\circ$$

$$P_{atm} + P_{Hg} = P_{gas}$$

$$76 + 30 \times \frac{1}{2} = P_{gas}$$

$$P_{gas} = 91 \text{ cm of Hg}$$

Q)



$$h = 25 \text{ cm of Hg}$$

$$P_{atm} = h d g$$

~~$$= 25 \times 6.8 \times 10^{-3}$$~~

cm

cm^2

~~$$= 25 \times$$~~

$$P_e = P_{Hg}$$

$$h_{Hg} \times d_{Hg} \times g = h_{Hg} \times d_{Hg} \times g$$

$$h_{Hg} = \frac{h \times d e}{d_{Hg}}$$

$$= \frac{25 \times 6.8}{13.6}$$

$$h_{Hg} = 12.5 \text{ cm of Hg}$$

$$P_{gas} = P_{atm} + P_{Hg}$$

$$= 76 + 12.5$$

$$= 88.5 \text{ cm of Hg}$$

Patm

J

Liquid = 1

 $d = 3.14 \text{ g/cm}^3$

Vessel = 2

 $d = 6.8 \text{ g/cm}^3$ Liquid = 3, $d = 27.2 \text{ g/cm}^3$

$$\text{Patm} + P_{\text{g-1}} + P_{\text{g-2}} + P_{\text{g-3}} = P_{\text{gas}}$$

$$P_{\text{g-1}} = P_{\text{Hg}}$$

$$h_{\text{mercury}} = h_{\text{Hg}} \times d_{\text{Hg}}$$

$$h_{\text{Hg}} = \frac{20 \text{ cm}}{13.6} \times \frac{3.14}{4}$$

$$= 20 \cancel{13.6} \text{ cm of Hg}$$

 $P_{\text{g-2}}$

$$h_{\text{Hg}} = \frac{20 \times 6.8}{13.6} = 10 \text{ cm of Hg}$$

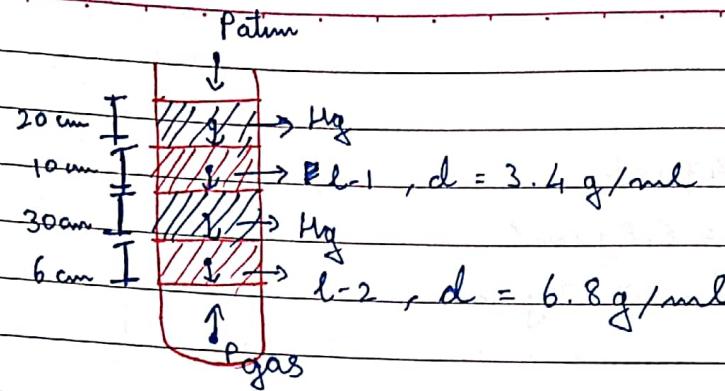
 $P_{\text{g-3}}$

$$h_{\text{Hg}} = \frac{10 \times 27.2}{13.6} = 20 \text{ cm of Hg}$$

$$P_{\text{gas}} = 76 + \cancel{4.5} + 10 + \overset{2.0}{\cancel{2.0}}$$

$$= \cancel{10.6} + 4.5 - \cancel{10.6} + 2.0 = 12.6 \text{ cm of Hg}$$

a)

P_{l-1}

$$h_{\text{Hg}} = \frac{h \times d_e}{d_{\text{Hg}}}$$

$$= \frac{10 \times 3.1}{13.6} = 2.5 \text{ cm of Hg}$$

P_{l-2}

$$h_{\text{Hg}} = \frac{6 \times 6.8}{13.6} = 3 \text{ cm of Hg}$$

Patum
↑

$$\therefore P_{\text{gas}} = 76 + 2.5 + 20 + 30 + 3$$

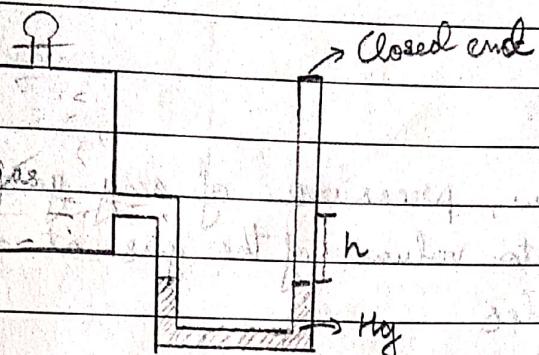
$$h_{\text{Hg}} = 126 + 5.5$$

$$= 131.5 \text{ cm of Hg}$$

Manometer

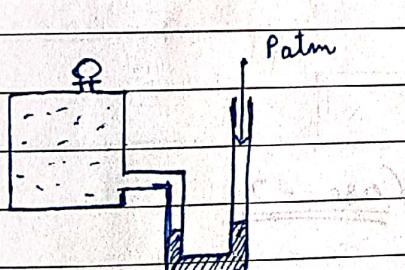
- It is used to measure the pressure of gas

① Closed end Manometer :-

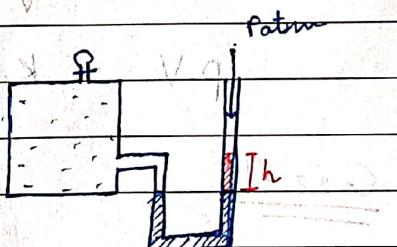


$$P_h = P_{gas}$$

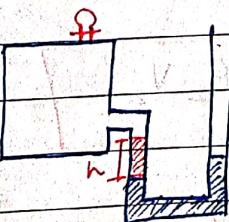
② Open end Manometer :-



$$P_{gas} = P_{atm}$$



$$P_{gas} = P_{atm} + P_h$$



$$P_{gas} = P_{atm} - P_h$$

⇒ Gas Laws

- These laws are mainly used to relate the measurable properties (p, v, n, T) of the gas.

⇒ Boyle's law

- Acc. to this law, pressure of any gas is inversely proportional to volume of the gas at constant temp. & moles.

$$\boxed{P \propto \frac{1}{V}} \text{ at constant } n, T$$

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

$$PV = K$$

Case - 1

$$P_1 V_1 = K$$

Case - 2

$$P_2 V_2 = K$$

$$\therefore \boxed{P_1 V_1 = P_2 V_2}$$

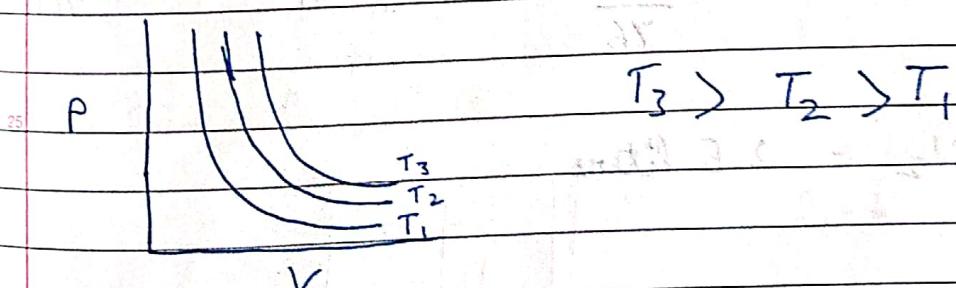
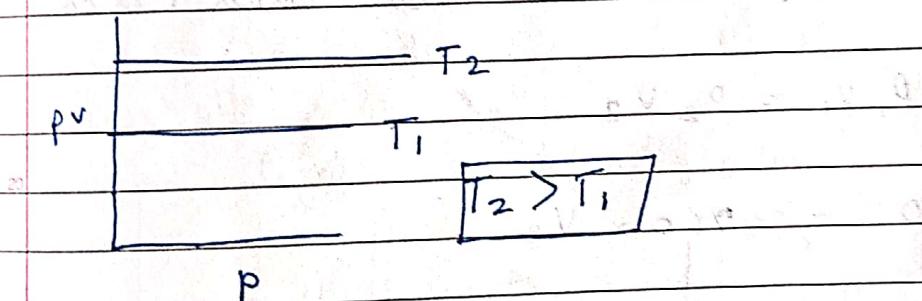
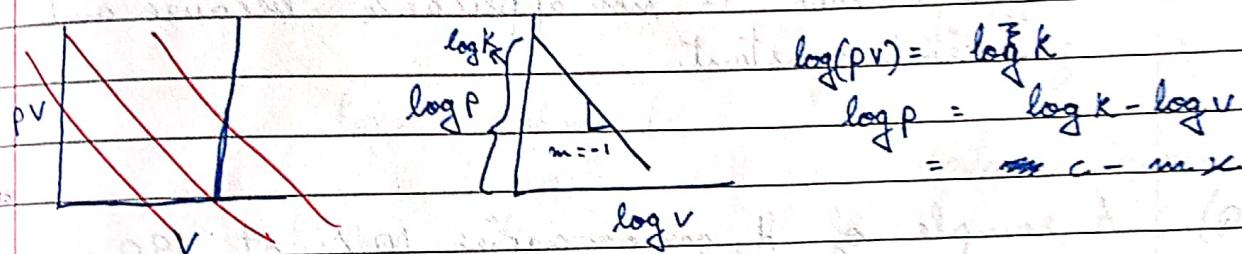
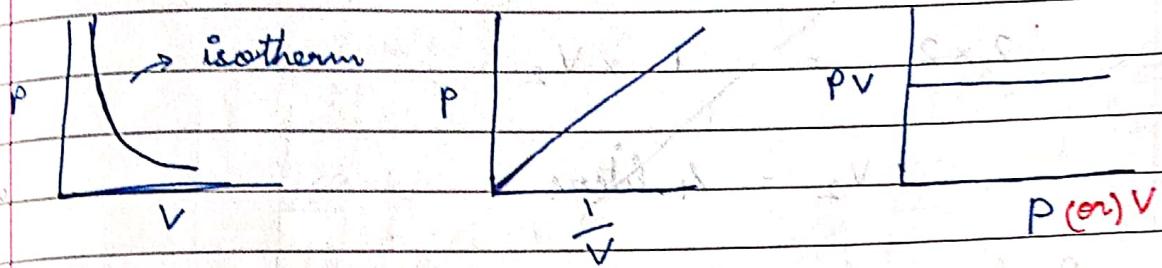
$$\boxed{\frac{P_1}{P_2} = \frac{V_2}{V_1}}$$

T - constant \rightarrow Isotherm

P - constant \rightarrow Isobar

V - constant \rightarrow Isochore

\Rightarrow Graph



- Q) A balloon contains gas with pressure 2 atm & and $V = 2 \text{ liters}$
after gas leak $P = 1 \text{ atm}$, $V = ?$

$$P_1 V_1 = P_2 V_2$$

$$2 \times 2 = 1 \times V_2$$

$$V_2 = 4 \text{ liters}$$

Boyle's law is not applicable because amt of gas isn't constant.

- Q) A sample of H_2 gas occupies 10 L at 190 mm pressure what would be the volume at 1 atm when T is constant

$$P_1 V_1 = P_2 V_2$$

$$190 \times 10 = 760 \times V_2$$

$$V_2 = \frac{190}{76}$$

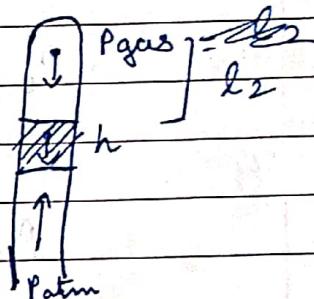
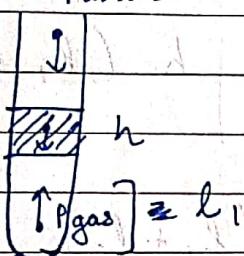
$$\begin{array}{r} 2.5 \\ 76 \sqrt{190} \\ \quad 152 \\ \quad 380 \\ \quad 380 \end{array}$$

$$V_2 = 2.5 \text{ liters}$$

Q) A gas column trapped b/w closed end of a tube and Hg column of length $\rightarrow h$

When it is placed open end upwards, length is l_1 and becomes l_2 when open end is downwards
 Find atm pressure in terms of 'Hg' columns

Patm is



$$P_{\text{gas}} = \text{Patm} + P_h$$

$$\text{Patm} = P_{\text{gas}} - P_h$$

~~$$\text{Patm} = P_{\text{gas}} - P_h$$~~

~~$$P_{\text{gas}} = \text{Patm} + P_h$$~~

~~$$P_{\text{gas}} l_1 - h = l_2 + h$$~~

~~$$l_1 - l_2 = 2h$$~~

$$\text{Patm} = P_0$$

$$P_1 = P_0 + h \quad , \quad P_0 - h = P_2$$

$$V_1 = l_1 \times A \quad , \quad V_2 = l_2 \times A$$

$$P_1 V_1 = P_2 V_2$$

$$(P_0 + h) (l_1 \times A) = (P_0 - h) l_2 \times A$$

$$P_0 l_1 + h l_1 = P_0 l_2 - h l_2$$

~~cancel~~

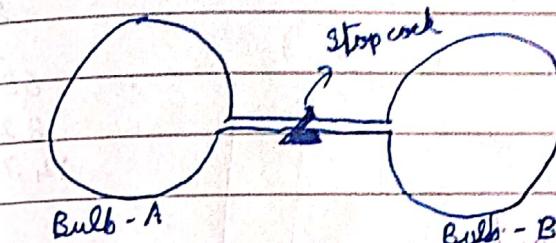
$$P_0 (l_2 - l_1) = h (l_1 + l_2)$$

$$P_0 = \frac{h (l_1 + l_2)}{(l_2 - l_1)}$$

→ Concept of connecting bulb problem

Case (i)

Temperature is same for both the bulbs



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

$$V_1 = 10 \text{ L}$$

$$W_1 = 20 \text{ g of H}_2$$

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

$$V_2 = 15 \text{ L}$$

$$W_2 = 160 \text{ g of O}_2$$

$$\begin{aligned} pV &= n(R)T \\ &= \frac{Wg}{M.W.} \times R \times T \end{aligned}$$

Calculate pressure in both the bulb, when stopcock is closed?

$$p_1 V_1 = \frac{W_1}{M.W.} \times 0.0821 \times 300$$

$$10 p_1 = \frac{20}{2} \times 0.0821 \times 300$$

$$\begin{aligned} p_1 &= 3 \times \frac{821}{1000} \times 1000 \\ &= 25 \text{ atm} \end{aligned}$$

$$p_2 V_2 = \frac{W_2}{M.W.} \times R \times T$$

$$15 p_2 = \frac{160}{32} \times 0.0821 \times 300$$

$$p_2 = 100 \times 0.0821$$

$$= 38.42 \text{ atm}$$

When stopcock is open, what is the pressure?

$$p_f V_f = n RT \quad (\text{T is constant})$$

$$p_f \cdot 25 = (n_1 + n_2) RT$$

$$25 p_f = 15 \times 0.0821 \times 300$$

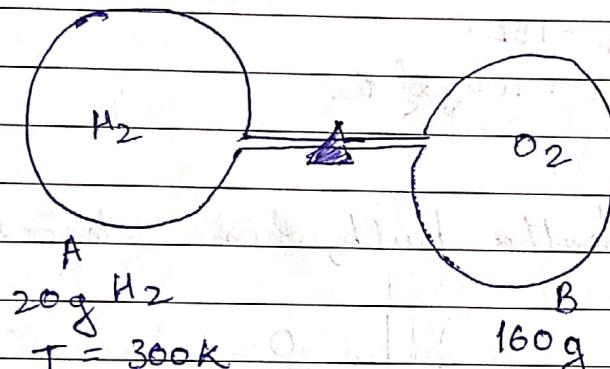
$$p_f = 3 \times 60 \times 0.0821$$

$$p_f = 18 \times 0.821$$

$$= 14.778$$

821
 18
 6568
 8210
 14.778

Case(ii) If T isn't constant



$$T_A V_1 = 10 L \quad T_B V_2 = 15 L$$

$$T_A = 300 K \quad T_B = 400 K$$

$$V_1 = 10 L \quad V_2 = 15 L$$

When stopcock is closed

H₂

$$pV = nRT$$

$$p_{H_2} = 25 \text{ atm}$$

(atmosphere & T)

O₂

$$PV = nRT$$

$$P = \frac{160}{32} \times 0.0821 \times \frac{400}{15}$$

$$(x+0.1)P = (x+2)R$$

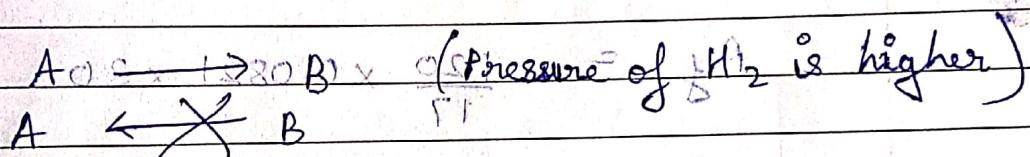
$$= 11.1 \text{ atm}$$

J.F.E

Add.

When stopcock is open

Calculate no. of moles of gas and final pressure after stopcock is open



$$n_{H_2} = 10 \quad (\text{Initial})$$

$$n_{O_2} = 5 \quad (\text{Initial})$$

$$n_f = 10 - x$$

$$n_f = 5 + x$$

In Δ Volume V_2 , we can't add because T isn't constant

$$\therefore P_f \times V_1 = (10-x) R \times 300 - ① \quad T \text{ is fixed}$$

$$P_f \times V_2 = (5+x) R \times 400 - ②$$

$$\left[\frac{dP}{dT} + 1 \right]_T V = P_f V$$

①
 ②

$$\frac{P_1}{2} \times 10 = (10-x) R \times 300$$

$$\frac{P_2}{2} \times 15 = (5+x) R \times 400$$

$$8(5+x) = 9(10-x)$$

$$x = \frac{50}{17}$$

$$= 2.9$$

$$P_1 \times 10 = (10 - \frac{50}{17}) \times 0.0821 \times 300$$

$$P_1 = \frac{120}{17} \times 0.0821 \times 300$$

$$P_1 = 17.1 \text{ atm}$$

\Rightarrow Charles law

- According to Charles law, The volume of a given mass of gas increases or decreases by $\frac{1}{273}$ of its volume at 0° Celsius , for each 1° rise or fall, provided T is constant ($T = 0^\circ \text{ Celsius}$)

$$V_1 = V_0 \times (1 + \frac{\Delta T}{273})$$

$$V_{\text{at} T} = V_0 \left[1 + \frac{\Delta t}{273} \right]$$

$$V_{\text{at } T} = V_0 \left[\frac{273 + t}{273} \right] \rightarrow T \quad \begin{matrix} V_{\text{at } T} \\ T \end{matrix}$$

$$T = 273 + t$$

$$V_T = \frac{V_0 (273 + T)}{273} \Rightarrow V_{\text{at } T}$$

$V(\text{cc})$

\propto

$$V \propto T$$

at constant

m, P

$$V = kT$$

$$273.15^\circ\text{C} \rightarrow T^\circ\text{C} \uparrow$$

$$\leftarrow T \quad \frac{V}{T} \rightarrow 273.15\text{K}$$

Absolute
Temperature
(Kelvin)

Case i)

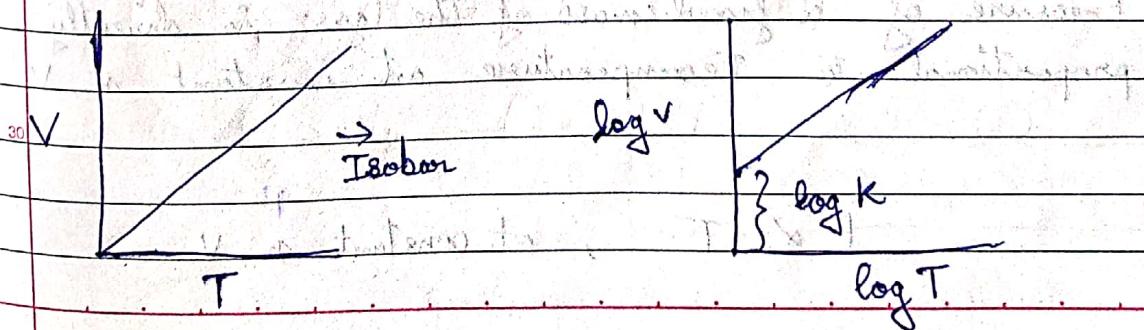
Case ii)

$$\frac{V_1}{T_1} = K$$

$$\frac{V_2}{T_2} = K$$

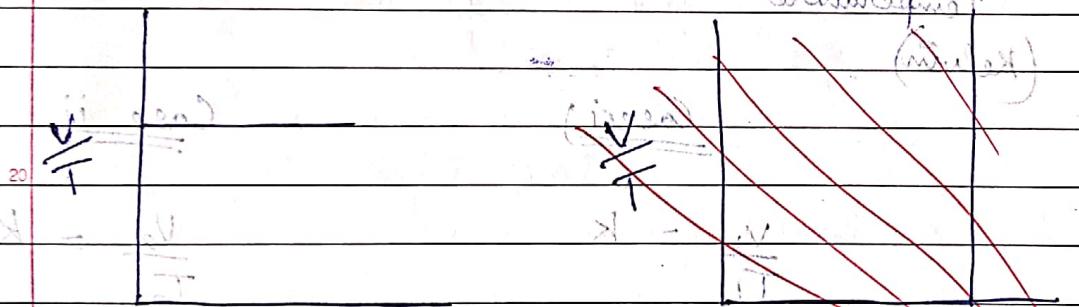
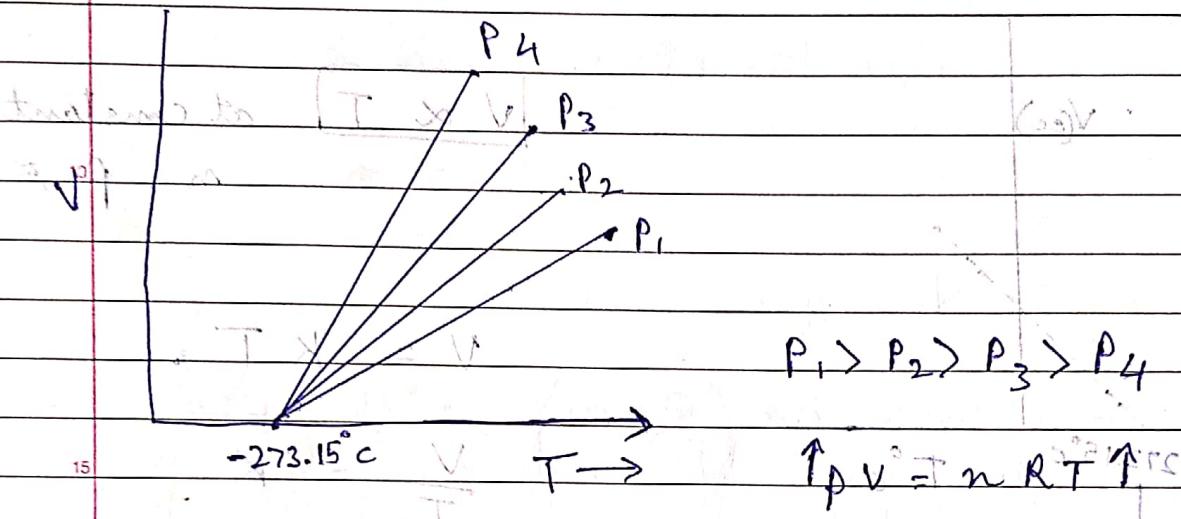
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

\Rightarrow graphs



$$\frac{\log V}{\log T} = \frac{\log k}{\log T}$$

$$\log V = \log K + \log T$$



$$V \propto T$$

$$\frac{V}{T} = \text{constant}$$

\Rightarrow Gay Lussac's law

Pressure of a fixed mass of the gas is directly proportional to Temperature at constant n, V

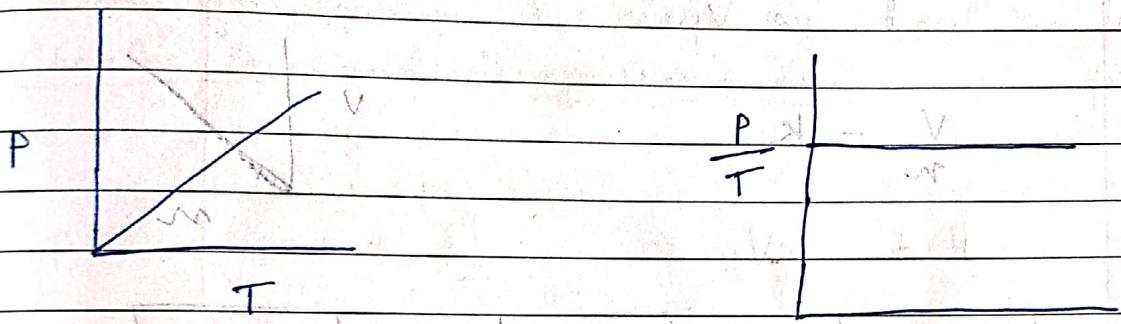
$$P \propto T$$

$$P \propto T, \text{ at constant } n, V$$

$$P = kT$$

$$\frac{P}{T} = k = \text{constant}$$

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



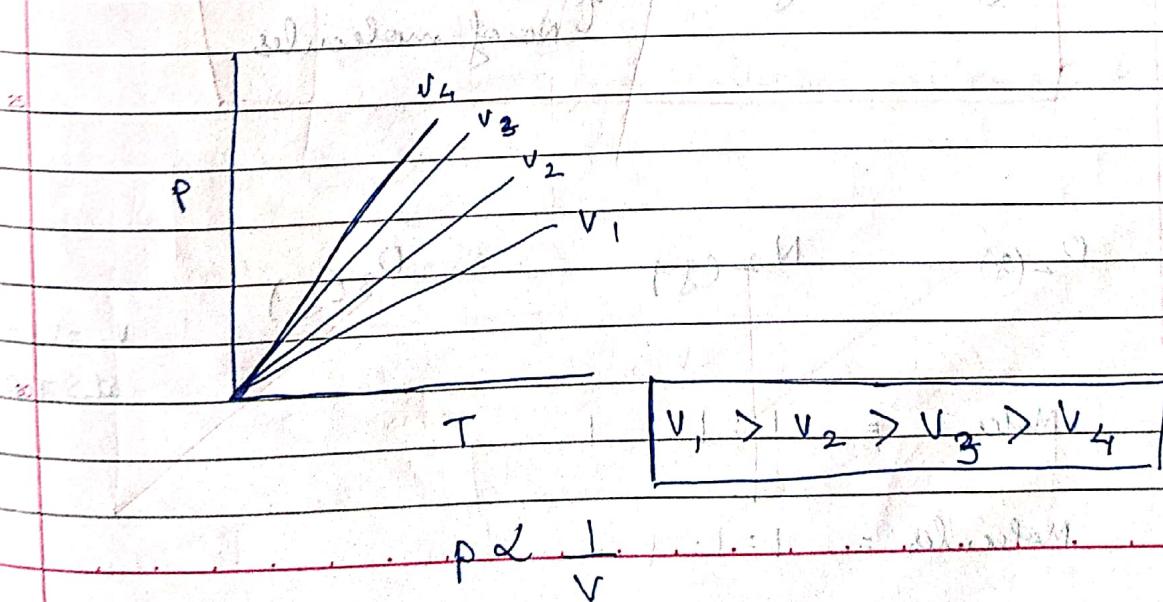
$$\log P - \log T = \log k$$

$\log P$

{ $\log k$

$\log T$

$$\log P = \log k + \log T$$

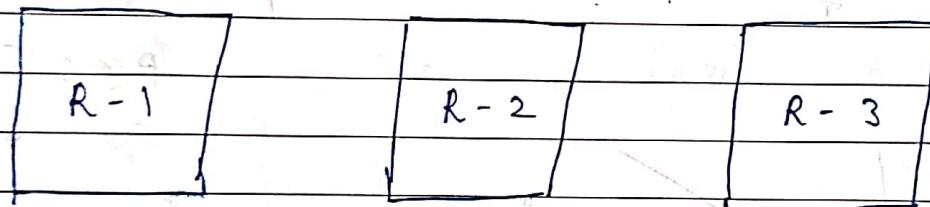
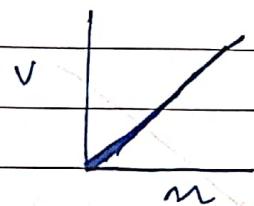


→ Avogadro's Law

- Acc. to this law, Volume of the gas is directly proportional to no: of moles at constant T, P

$V \propto n$ at constant T, P

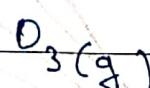
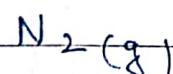
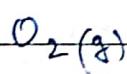
$$\frac{V}{n} = k$$



If all rooms have same volume at STP

These rooms contains same moles of gas

$V_{\text{gas}} \propto n_{\text{gas}} \propto \frac{N}{\uparrow \text{No of molecules}}$



$$V_1 = V_2 = V_3 \\ \text{at STP}$$

Moles $\propto V \propto 1 : 1 : 1$

Molecules $\propto 1 : 1 : 1$

No. of atoms = 2 : 2 : 3

Ideal Gas Equation :-

- Ideal Gas is a hypothetical gas, In nature no gas is ideal gas.

Ideal gas equation is derived by Boyle's law, Charles' law and Avogadro's law

$$p \propto \frac{1}{V} [n, T] \Rightarrow pV = K$$

$$V \propto T [n, p] \Rightarrow \frac{V}{T} = K$$

$$V \propto n [P, T] \Rightarrow \frac{V}{n} = K$$

$$\therefore V \propto \frac{1}{P} \propto T \propto n$$

$$V \propto \underline{Tn}$$

$\rightarrow R$ is universal gas constant

$$pV = \textcircled{R} nT$$

$$\boxed{pV = nRT}$$

R → Universal Gas Constant

$$R = \frac{pV}{nT} = \frac{\text{Force} \times \text{Volume}}{\text{Area} \times \text{Moles} \times \text{Kelvin}}$$

$$= \frac{\text{Force} \times L^3}{\text{Moles} \times \text{Kelvin}} = \frac{\text{Work}}{\text{Moles} \times \text{Kelvin}}$$

$$= J \times K^{-1} \times \text{mole}^{-1}$$

$$R = \frac{101325 \text{ N/m}^2 \times [22.4 \times 10^{-3} \text{ m}^3]}{1 \text{ mol} \times 273 \text{ K}}$$

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

$$(2) R = \frac{1 \text{ atm} \times 22.4 \text{ L}}{1 \text{ mole} \times 273 \text{ K}}$$

$$R = 0.0821 \text{ L atm mole}^{-1} \text{ K}^{-1}$$

$$(3) 1 \text{ cal} = 4.184 \text{ J}$$

$$R = 2 \text{ cal K}^{-1} \text{ mole}^{-1}$$

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

$$R = 0.0821 \text{ L atm K}^{-1} \text{ mole}^{-1}$$

$$R = 2 \text{ cal} \text{ K}^{-1} \text{ mole}^{-1}$$

$$R = 8.314 \times 10^7 \text{ erg K}^{-1} \text{ mole}^{-1}$$

$$1 \text{ cal} = 4.184 \text{ J}$$

$$1 \text{ J} = 10^7 \text{ ergs}$$

$$1 \text{ KJ} = 10^3 \text{ J}$$

At STP / NTP ^{Normal}

STP

$$p = 1 \text{ atm} \quad T = 0^\circ\text{C} = 273 \text{ K} \quad V = 22.4 \text{ L - 1 mole}$$

NTP

$$p = 1 \text{ bar} \quad T = 0^\circ\text{C} = 273 \text{ K}, \quad V = 22.7 \text{ L} \quad (1 \text{ mole})$$

$$1 \text{ bar} = 0.987 \text{ atm}$$

SATP

Standard Ambient Temperature and Pressure

$$p = 1 \text{ atm}, \quad T = 25^\circ\text{C} = 298 \text{ K}, \quad V_g = 24.4 \text{ L}$$

$$p = 1 \text{ bar}, \quad T = 25^\circ\text{C} = 298 \text{ K}, \quad V_g = 24.7 \text{ L}$$

$$\underline{pV = nRT}$$

1 mole of gas, Initially P_1, V_1 and T_1

1 mole of gas, Finally P_2, V_2 , and T_2

$$\therefore \frac{P_1 V_1}{T_1} = nR \quad \text{--- (1)}$$

$$\frac{P_2 V_2}{T_2} = nR \quad \text{--- (2)}$$

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

$$\Rightarrow pV = nRT$$

$$pV = \frac{W \times RT}{MW}$$

$$m = \frac{W}{MW}$$

$$MW \times p = \frac{W}{V} \times RT$$

$$d = \frac{W}{V}$$

$$\boxed{MW = \frac{dRT}{p}}$$

Note :-
* In case

Note:- * In case of ideal gas there is no intermolecular force of attraction b/w two gas molecules

* Volume of ideal gas is negligible to the volume of the container

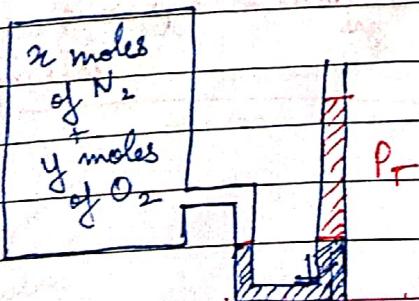
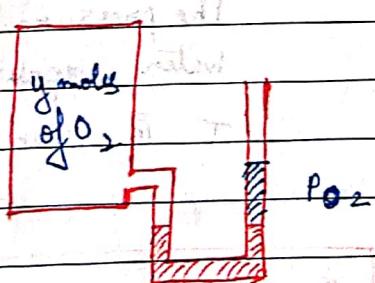
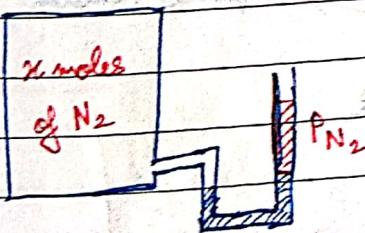
Dalton's law of Partial pressure

→ Partial pressure - The pressure exerted by individual called partial pressure

→ This law is applicable for mixture of nonreacting gases

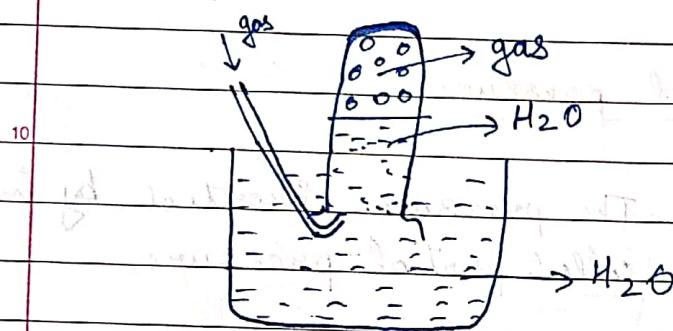
→ According to Dalton's Law Total pressure of gas is equal to sum of partial pressures of individual gases

$$P_T = P_1 + P_2 + P_3 \dots \quad [T, V]$$



$$P_T = P_{N_2} + P_{O_2}$$

* Generally, gases are collected over water.



$$P_{\text{moisture}} = P_{\text{dry gas}} + P_{\text{water vapour}}$$

$$P_{\text{dry gas}} = P_{\text{moisture}} - P_{\text{water vapour}}$$

↑
constant at particular T

The pressure exerted by water vapour at particular T is called aqueous tension.

$$P_{\text{dry gas}} = P_{\text{observed}} - \text{aqueous tension}$$

$$P_A = m_A \times \frac{RT}{V}$$

$$P_B = m_B \times \frac{RT}{V}$$

$$P_C = m_C \times \frac{RT}{V}$$

$$P_T = P_A + P_B + P_C$$

$$= (m_A + m_B + m_C) \frac{RT}{V}$$

$$\frac{P_A}{P_T} = \frac{n_A}{n_A + n_B + n_C}$$

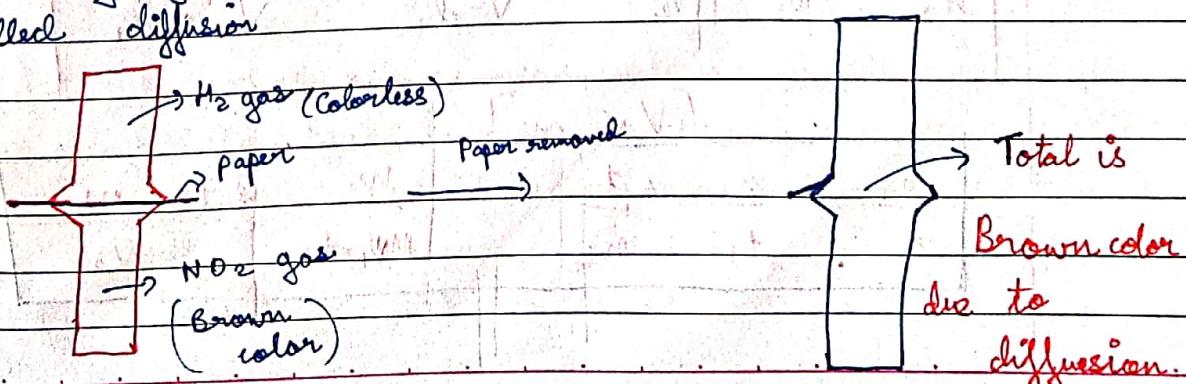
$$P_A = X_A \cdot P_T$$

$$P_B = X_B \cdot P_T$$

$$P_C = X_C \cdot P_T$$

Diffusion

- Intermixing of 2 gases without applying any external force is called diffusion



Graham's law of Diffusion

Graham studied the rate of diffusion of different gases and given a mathematical relationship

i.e.

$$\alpha_1 \propto \frac{1}{\sqrt{d_1}}$$

$$d = \frac{M.W}{2} \quad (\text{Vapour Density})$$

$$\alpha_1 \propto \frac{1}{\sqrt{d_1}}$$

$$\alpha_2 \propto \frac{1}{\sqrt{d_2}}$$

$$\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{\frac{M.W_2}{2}}{\frac{M.W_1}{2}}}$$

$$\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{M.W_2}{M.W_1}}$$

$$\text{Rate} = \frac{\text{Volume of gas diffused}}{\text{time taken}} = \frac{V}{t}$$

$$\frac{\alpha_1}{\alpha_2} = \frac{\frac{V_1}{t_1} \times t_2}{\frac{V_2}{t_2}} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M.W_2}{M.W_1}}$$

$$[V \propto n]$$

$$\frac{\alpha_1}{\alpha_2} = \frac{n_1}{n_2} \times \frac{t_2}{t_1} = \frac{W_1}{M.W_1 \times t_1} \times \frac{t_2 \times M.W_2}{W_2}$$

σ vs $p \& d$

$$\sigma \propto p \propto \frac{1}{\sqrt{d}}$$

$$\boxed{\frac{\sigma_1}{\sigma_2} = \frac{P_1}{P_2} \sqrt{\frac{d_2}{d_1}}}$$

σ vs $T \& MW$

$$\sigma \propto \sqrt{T} \propto \frac{1}{\sqrt{M.W}}$$

$$\frac{n_1}{n_2} = \sqrt{\frac{T_1 \cdot M.W_2}{T_2 \cdot M.W_1}}$$

$$\frac{v_1}{v_2} = \frac{n_1 \cdot t_2}{n_2 \cdot t_1} = \frac{n_1 \cdot t_2}{n_2} \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M.W_2}{M.W_1}} = \frac{P_1}{P_2} \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{T_1 \cdot M.W_2}{T_2 \cdot M.W_1}}$$

$$\boxed{\sigma \propto \frac{1}{\sqrt{d}} \propto \frac{1}{\sqrt{M.W}} \propto p \propto \sqrt{T}}$$

- Q) At an under water depth of 250 ft the pressure is 8.38 atm, what should be the molar % of O_2 in the diving gas for the partial pressure of O_2 in the mixture to be 0.21 atm the same as it is in air at 1 atm?

$$P_f = 8.38 \text{ atm} \quad P_{O_2} = 0.21 \text{ atm}$$

$$P_{O_2} = 1 \text{ atm}$$

$$P_{O_2} = X_{O_2} \cdot P_T$$

$$X_{O_2} = \frac{P_{O_2}}{P_T} = \frac{0.21}{8.38}$$

$$= 0.025$$

$$= 2.5\%$$

Q) Which of the 2 gases NH_3 , HCl will diffuse faster by what factor

$$\frac{n_1}{n_2} = \sqrt{\frac{MW_2}{MW_1}}$$

$$= \sqrt{\frac{36.5}{17}} = 1.46$$

$$= \sqrt{2.2}$$

$$n_{NH_3} = 1.46 \times n_{HCl}$$

Kinetic Molecular Theory of Gases

This Theory that provides an explanation for all the gas laws. It is assumed that all gases are made up of identical shape and size which are in random motion.

Postulates

- 1) A gas consists of a large no: of molecules of mass 'm' - the dimensions of these are very small compared to the space b/w them.
 - 2) So these molecules are treated as point masses which are in random motion.
 - 3) There are practically no attractive forces b/w the molecules so they can move individually.
 - 4) The molecules are in random motion so they collide with each other and the walls of the container because of the collisions direction of their motion will change and these collision are elastic so the energy and momentum is conserved.
 - 5) The press of the gas is due to collisions with the walls of the container.
 - 6) Avg K.E of gas molecules & Temperature of the gas
- The kinetic gas eqn. derived on the basis of postulates of kinetic Theory
- $$PV = \frac{1}{3} m N (\bar{u}_{rms}^2)$$
- p = pressure
V = volume
m = mass of one gas molecule
N = No: of molecules with volume (V)
 \bar{u}_{rms} = Root Mean Square Velocity.

$$m \times N = M$$

$M \rightarrow$ Molecular weight

$$pV = \frac{1}{3} M(u_{rms})^2 \times \frac{2}{3}$$

$$pV = \frac{2}{3} \times \frac{1}{2} M(u_{rms})^2 \rightarrow K.E_{avg}$$

$$pV = \frac{2}{3} K.E_{avg}$$

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

$$\begin{bmatrix} n = 1 \\ pV = RT \end{bmatrix}$$

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

$$K.E \propto T$$

$$K.E_{avg} = \frac{3}{2} \times \frac{R}{N} \times T$$

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

$$K_B = \frac{R}{N_A}$$

Boltzmann constant

$$K_B = 1.38 \times 10^{-23} J/K$$

$$K.E. = \frac{3}{2} P V$$

$$P V = R T$$

$$= \frac{3}{2} \frac{P M}{d}$$

$$M w = \frac{d R T}{P}$$

$$K.E. = \frac{3}{2} \frac{P M}{d}$$

Molecular Velocities

① Root Mean square velocity (v_{rms}): (R)

② Average velocity (v_{avg})

③ Most Probable Velocity (v_{mp})

$$R > A > M$$

① Root Mean square velocity (v_{rms}): (R)

The molecules in a container going at high speeds/velocity

$$U_{rms}^2 = n_1 v_{rms}^2 + n_2 v_{2rms}^2 + n_3 v_{3rms}^2$$

$$n = n_1 + n_2 + n_3$$

$$U_{RMS}^2 = \frac{3RT}{M}$$

$$U_{RMS} = \sqrt{\frac{3RT}{M}}$$

$$U_{RMS} = \sqrt{\frac{3PV}{M}}$$

$$U_{RMS} = \sqrt{\frac{3P}{d}}$$

(2) Average Velocity (A)

$$U_{avg} = \frac{n_1 U_1 + n_2 U_2 + n_3 U_3}{n_1 + n_2 + n_3}$$

$$U_{avg} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8PV}{\pi M}} \leq \sqrt{\frac{8P}{\pi d}}$$

As per Kinetic Theory each molecule moves with different velocity, then the average velocity is given above:

(3) Most Probable Velocity (U_{mp}) (M)

The velocity possessed by most of the molecules

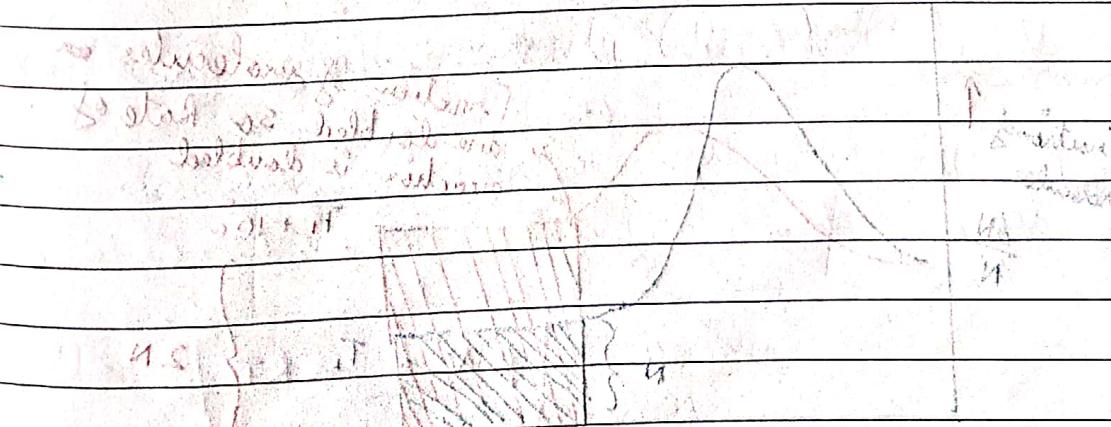
$$U_{MP} = \sqrt{\frac{2RT}{M}}$$

$$U_{MP} = \sqrt{\frac{2PV}{d}} = \sqrt{\frac{2P}{d}}$$

$$U_{RMS} > U_{AVG} > U_{MP}$$

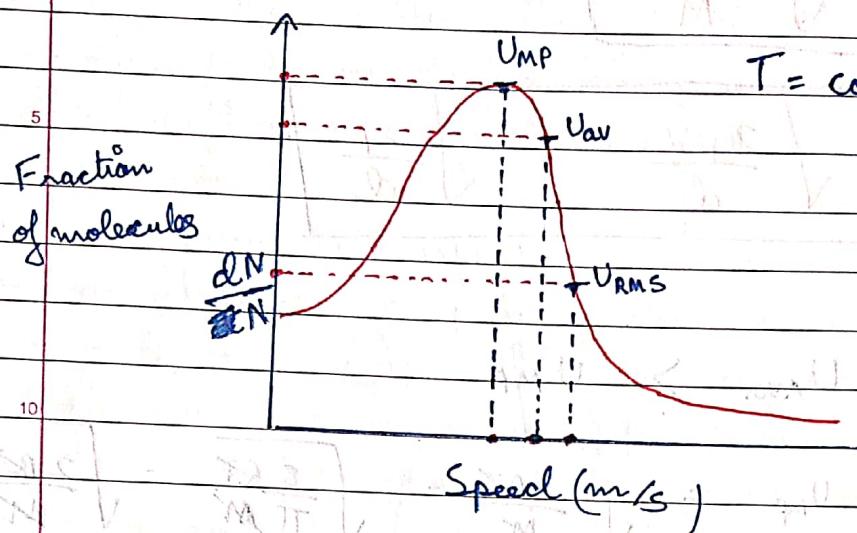
$$U_{RMS} : U_{AVG} : U_{MP} = \sqrt{\frac{3RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{2RT}{M}}$$

$$= \sqrt{3} = \sqrt{\frac{8}{3 \cdot 14}} = \sqrt{2}$$



intensity of light is measured along the x-axis.

Maxwell distribution of Speeds

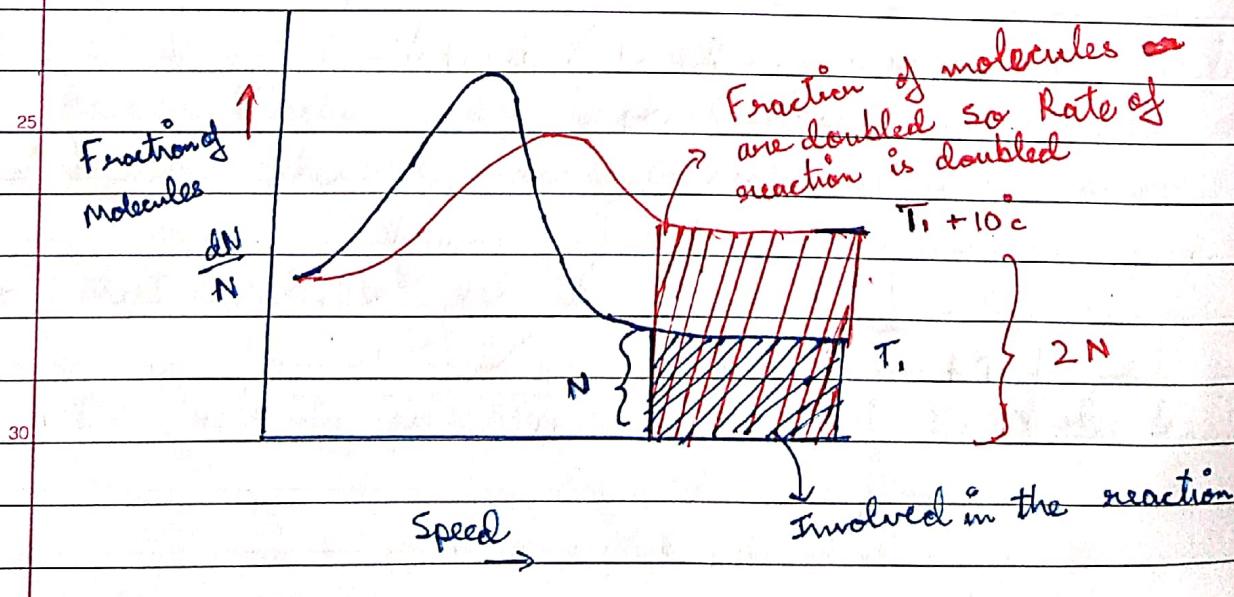


Speed

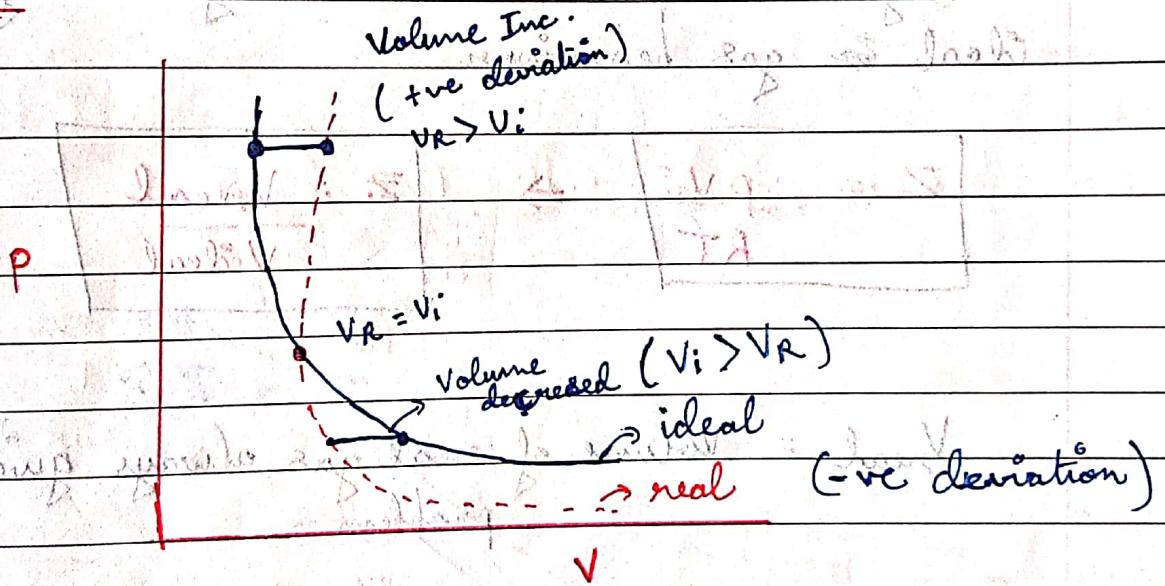
$$U_{RMS} > U_{avg} > U_{MP}$$

Kinetics

- For every 10°C rise in temperature, the rate of reaction is doubled or tripled.



⇒ Real Gas. $P \text{ not } \propto V^{-1}$ because $V_i < V_R$ and $T \text{ is constant}$



Ideal Gas

(i) A gas that follows the Ideal gas equation at all P, T

(ii) The volume occupied is negligible compared to the volume of container and also b/w 2 molecules there is no force of attraction

Real Gas

Gas that follows the Ideal gas equation at low P and high T

The volume occupied is not negligible compared to the volume of container and also b/w 2 molecules there is significant force of attraction

Compressibility Factor (\bar{z})

This factor explains deviation of Real gas from Ideal gas behaviour

$$Z = \frac{PV}{RT} \Rightarrow Z = \frac{V_{\text{real}}}{V_{\text{ideal}}}$$

V_{real} = Volume of real gas always given in the problem

V_{ideal} = Volume of ideal gas we have to calculate

$$PV = nRT$$

$$z = 1 \Rightarrow V_R = V_i \Rightarrow \text{Ideal gas}$$

$$z = f(p, T)$$

$$z \neq 1$$

$$z < 1$$

$$z > 1$$

$$V_i > V_R$$

$$V_i < V_R$$

Attractive forces b/w
gas molecules

Repulsive forces b/w
gas molecules

Q) Volume of 1 mole of gas is 25 l at STP, calc. (i)
the nature of gas

$$PV = 1 \times 25$$

$$V_R = 25 \text{ liters}$$

$$V_i = \frac{nRT}{P} = 22.4 \text{ liters}$$

$$PV > nRT$$

$$\therefore z > 1$$

$$\therefore V_R > V_i$$

$$z = \frac{25}{22.4} = 1.07 \therefore \text{Repulsive forces are there}$$

Q) Gases A, B, C, D have Z values 1.6, 0.8, 0.6, 1.8 respectively.

Arrange in inc. order of ideal behavior.

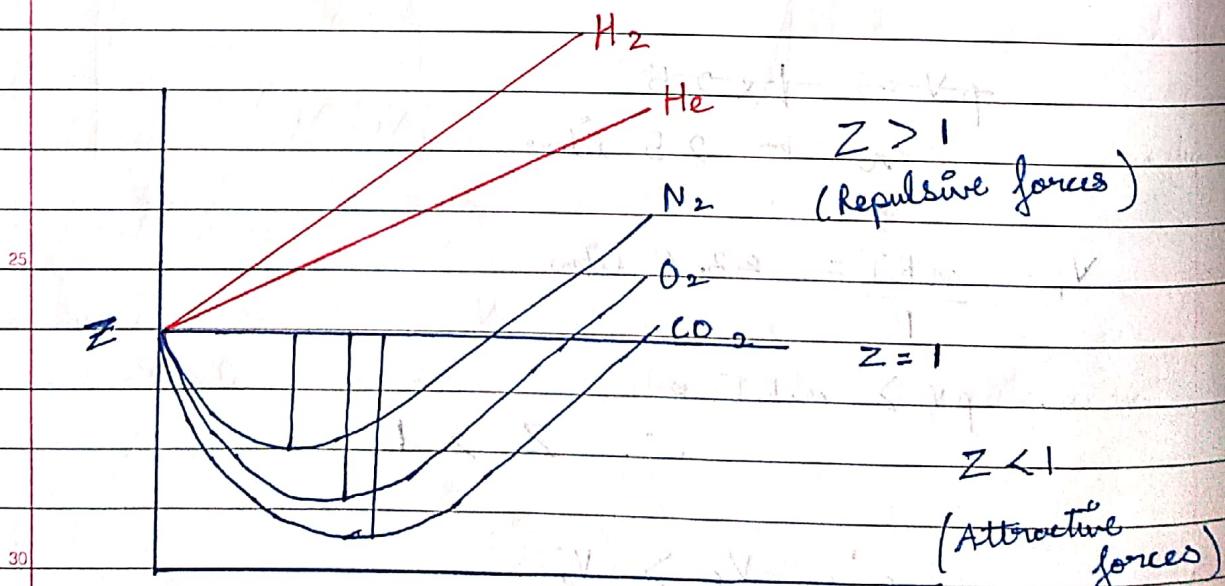
A	B	C	D
$Z = 1.6$	0.8	0.6	1.8

$$\text{deviation} = (Z - 1)$$

A	B	C	D
0.6	0.2	0.4	0.8

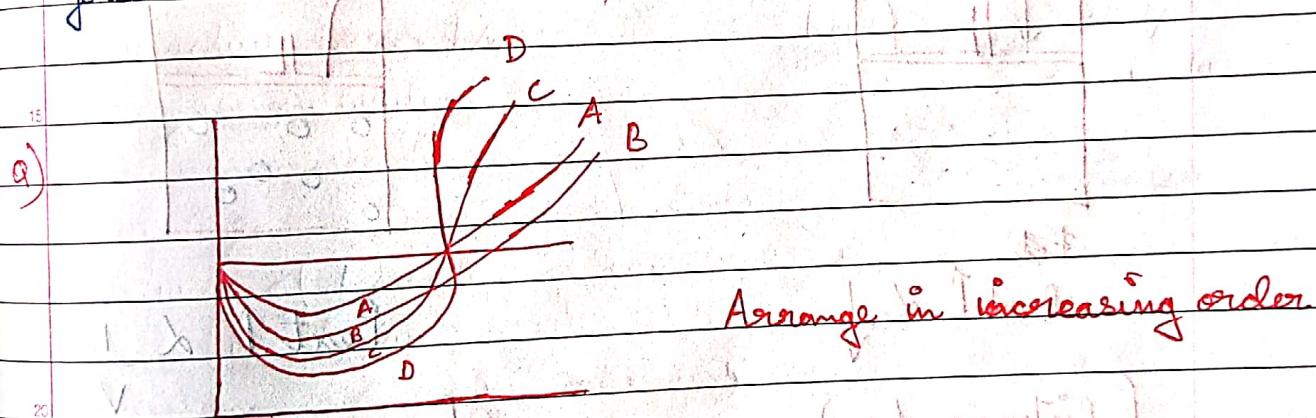
$$B < C < A < D$$

⇒ Variation of Z with P and T



- H_2, He shows +ve deviation at all P and T and graph is straight line.
- All other gases shows both +ve, -ve deviation and graph is curved.
- +ve deviation \rightarrow Low pressure.
- ve deviation \rightarrow High pressure.

Higher the dip of the curve, more will be the deviation from the ideal gas and more the attractive forces.



$$A < B < C < D$$

$\Rightarrow Z \text{ vs } T$ (For same gas)

$$T_1 > T_2 > T_3$$

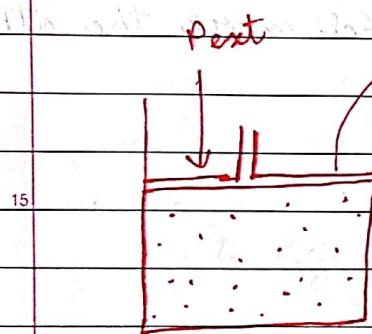
$$T \propto \frac{1}{\text{attraction forces}}$$

Reasons of Deviation from ideal gases

- Ideal gas molecules are point masses but real gas are rigid & spherical which have significant volume

I.G

point Mass
(Volume is negligible)



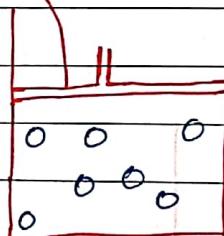
Pext

$$P_{ext} \propto \frac{1}{V}$$

Piston reaches
the bottom of
the container

R.G

rigid & spherical
(Volume is significant)



$$P_{ext} \propto \frac{1}{V}$$

Piston won't reach the
bottom because these
molecules have significant
Volume

- There is no force of attraction b/w ideal gas molecules but there is significant force of attraction b/w real gas

Derivation of Real Gas

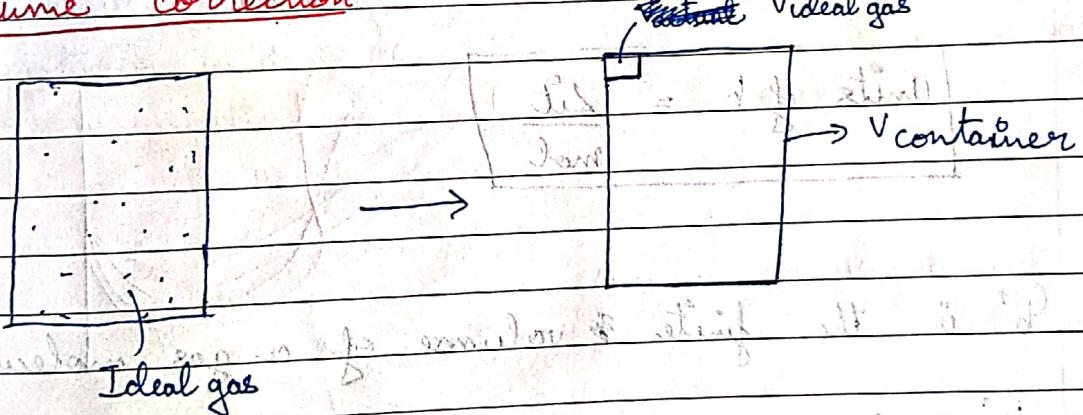
Equation derived by the following :

(i) Volume of gas molecules aren't negligible compared to the total volume of container

(ii) b/w gas molecules there is force of attraction.

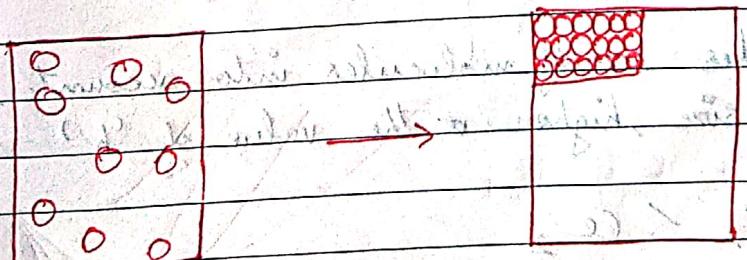
* Mandeswall is the first scientist who gave the real gas equation

Volume correction



$$V_{\text{Ideal gas}} < V_{\text{container}}$$

volume of ideal gas
molecules is negligible



$$V_{\text{actual}} = V_{\text{container}} - V_{\text{gas molecules}}$$

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

$V_{\text{actual}} = V - nb \rightarrow n \text{ moles of gas}$

b = Van der Waal's constant

$$b = \frac{V}{n} = \frac{\text{Lit}}{\text{mole}}$$

Units of b = $\frac{\text{Lit}}{\text{mol}}$

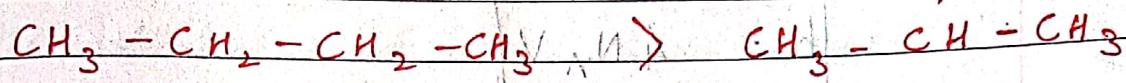
'b' is the finite volume of a gas molecule.

'b' is also known as 'Excluded' or 'co-volume'

→ Comparison of 'b' for different gases

- Since 'b' takes size of molecules into account, hence larger the size, higher is the value of 'b'





n-butane

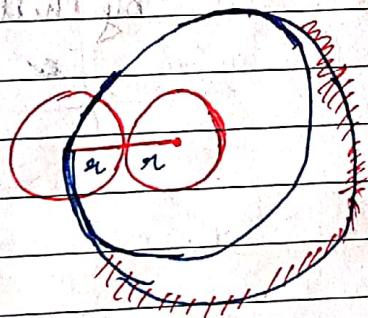
>

Isobutane

⇒ Branching decreases the surface area

⇒ Value of 'b'

* *b* is a constant which numerical value is given in the question but the help of radius of molecules '*b*' can be calculated



Excluded volume

$$\text{for 2 molecules} = \frac{4}{3} \pi (2a)^3$$

Excluded Volume of one molecule

$$= \frac{1}{2} \times \frac{4}{3} \pi \times 8a^3$$

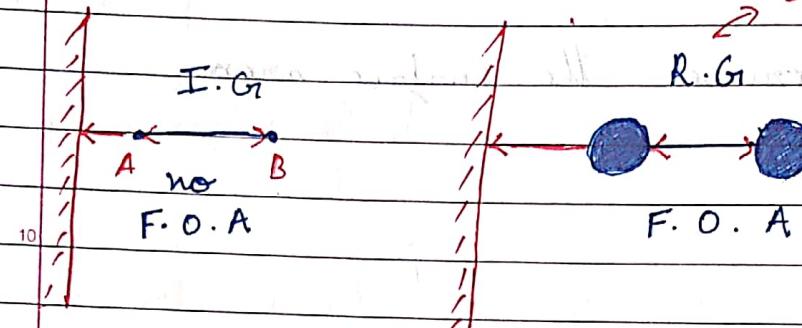
$$= 4 \times \frac{4}{3} \pi a^3$$

Excluded Volume by one

$$\text{mole of molecules} = 4 \times \frac{4}{3} \pi a^3 \times N_A$$

$$b = 4 N_A V$$

⇒ Pressure Correction



Speed is reduced b/w A and B because of F.O.A

Wall of container

Pressure is developed

due to the hitting the wall.

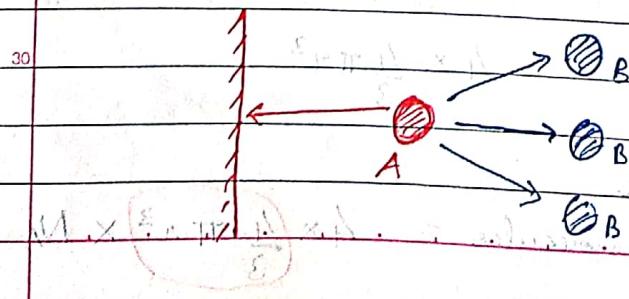
Pressure developed by hitting wall is $\downarrow p$

Pressure developed by hitting the wall is p_i

$$p_i > p$$

$$p_i = p + \beta$$

⇒ Calculation of β



β depend on attraction forces b/w (A) & (B) gas molecules

All the molecules present in the container will of A type and B type and their no. depends on density

$$\text{no: A type of molecules} = \frac{n}{V}$$

$$\text{no: B type of molecules} = \frac{n}{V}$$

$n \rightarrow \text{no: moles}$

$$\beta \propto \frac{n}{V} \propto \frac{n}{V}$$

$$\beta \propto \frac{n^2}{V^2}$$

$$B = \frac{an^2}{V^2} \quad (a \rightarrow \text{vanderwall's constant})$$

$$P_i = P + \frac{an^2}{V^2}$$

$$(\text{Units of } a) = \frac{\text{atm} \cdot \text{litre}^2}{\text{Mole}^2}$$

$$PV = nRT$$

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

Vanderwall's gas equation
a, b are vanderwall's constants.

$\Rightarrow \#$ Significance of 'a'

- * Higher the value of a , easier the liquification
- * For ideal gas, $a = 0$ so liquification is not possible

\Rightarrow Liquification Pressure

10 The pressure required to liquefy a gas is called as liquification Pressure.

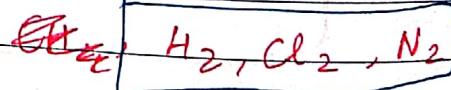
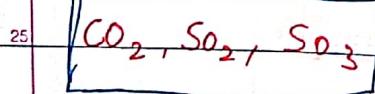
$$a \propto \frac{1}{L.P}$$

\Rightarrow Comparison of 'a'

15 * a is constant whose numerical value is given but we can compare the value of a for different gases

* 20 a value is higher for polar gases compared to non-polar gases

a polar \rightarrow a non-polar



* Between gas of same nature (polar/Non-polar) higher is the Molar mass, higher the value of a .

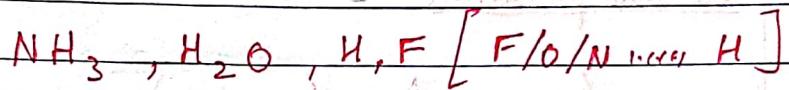
Non Polar

$$O_2 > N_2$$

(32) (28)

$$a_{O_2} > \cancel{a_{N_2}}$$

- Larger is the size, higher the force of attraction
- If hydrogen bond is present, then the value of a is higher than Vanderwall's force



- Due to very small size of H_2, He the value of a is negligible

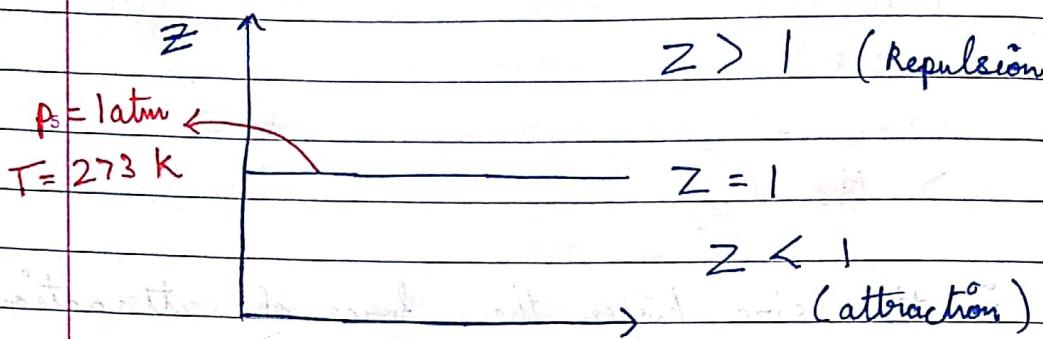
$$[a_{H_2, He}] \approx 0$$

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

for 1 mole of gas

$$\left(p + \frac{a}{V^2} \right) (V - b) = nRT$$

\Rightarrow Calculation of Z from Van der Waals gas equation



\Rightarrow at low pressure :-

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

$$V - b \approx V$$

$$PV + \frac{a}{V} = RT$$

$$\left(Z = \frac{PV}{RT} \right)$$

$$\frac{PV}{RT} + \frac{a}{VRT} = 1$$

$$Z = 1 - \frac{a}{VRT}$$

$$Z < 1$$

attractive force

(Negative deviation)

→ At high pressure :-

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

$$P + \frac{a}{V^2} \approx P$$

$$PV - Pb = RT$$

$$\frac{PV}{RT} - \frac{Pb}{RT} = 1$$

$$Z = 1 + \frac{Pb}{RT}$$

$$Z > 1$$

Repulsive force
(+ve deviation)

→ At high Temperature

$$\left(P + \frac{a}{V^2} \right) (V - b) \underset{T \uparrow, P \uparrow}{\approx} RT$$

$$T \uparrow \quad P \uparrow$$

$$P + \frac{a}{V^2} \approx P$$

$T \uparrow \quad K.E. \uparrow \quad$ Attractions ↓ $V \uparrow$

$$PV = RT$$

⇒ Exceptional behaviour of H₂ and He

H₂, He has lower size so attraction forces are negligible

$$a_{H_2, He} \approx 0$$

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

$$a = 0$$

$$PV - Pb = RT$$

$$\frac{PV}{RT} = -\frac{Pb}{RT} = 1$$

$$\frac{PV}{RT} = \frac{1 + Pb}{RT}$$

$$\bar{z} > 1$$

(+ve deviation)

⇒ Virial Equation of state

- It is a generalised eqn of real gas and all other gas equation can be expressed in Virial equation.

- In this, \bar{z} is expressed in terms of "powers of P " and " $\frac{1}{V_m}$ "

$$Z = 1 + \frac{B}{P} + \frac{C}{P^2} + \frac{D}{P^3} \dots \dots \quad [\text{In terms of } P]$$

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

$$Z = 1 + \frac{B}{V_m} + \frac{C}{(V_m)^2} + \frac{D}{(V_m)^3} \dots \dots \quad (\text{In terms of } V)$$

→ Conversion of Van der waal's gas equation to Virial Form

$$\left(P + \frac{a}{V_m^2}\right) (V_m - b) = RT$$

$$T \left(P + \frac{a}{V_m^2}\right) = \frac{RT}{V_m - b}$$

(Multiply by V_m)

$$V_m \left(P + \frac{a}{V_m^2}\right) = \frac{RT \cdot V_m}{V_m - b}$$

$$\frac{PV_m + \frac{a}{V_m}}{V_m} = \frac{RT \cdot V_m}{V_m - b}$$

$$PV_m = \frac{RT \cdot V_m}{V_m - b} - \frac{a}{V_m} \quad (\div RT)$$

$$Z = \frac{V_m}{V_m - b} - \frac{a}{V_m RT}$$

$$Z = \frac{1}{1 - \frac{b}{V_m}} - \frac{a}{V_m RT}$$

$$Z = \left(1 - \frac{b}{V_m}\right) - \frac{a}{V_m RT}$$

$$\Rightarrow (1-x)^{-1} = 1 + x + x^2 + x^3 + \dots$$

$$\therefore Z = \left(1 + \frac{b}{V_m} + \frac{b^2}{V_m^2} + \frac{b^3}{V_m^3} \dots\right) - \frac{a}{V_m RT}$$

$$Z = \left(1 + \cancel{b} \cdot \frac{1}{V_m} + b^2 \cdot \frac{1}{V_m^2} \dots\right) - \frac{a}{V_m RT}$$

$$Z = \left(1 + \cancel{\frac{b-a}{RT}} + b^2 \cdot \frac{1}{V_m^2} + b^3 \cdot \frac{1}{V_m^3} \dots\right)$$

↓ ↓ ↓ ↓
 B C D

25 B → Second Virial co-efficient

$$\therefore B = b - \frac{a}{RT}$$

Neglecting the higher value of volume terms

$$Z = 1 + \left(\frac{b-a}{RT}\right) \frac{1}{V_m}$$

For Ideal gas, $Z = 1$

$$1 = 1 + \left(b - \frac{a}{RT} \right) \frac{1}{V_m}$$

$$0 = \left(b - \frac{a}{RT} \right) \frac{1}{V_m}$$

$$\frac{b-a}{RT} = 0$$

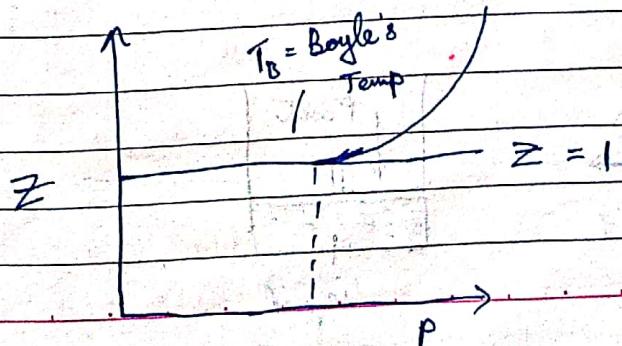
$$\frac{a}{RT} = b$$

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

Boyle's Temperature

- The temperature at which gas follows ideal behavior for large range of pressures is known as Boyle's Temperature.

- At Boyle's Temp. Effect produced by molecular attraction will be cancelled by effect produced by molecular size.



$$\textcircled{2} \quad T_B = \frac{a}{Rb} \quad T_i = \frac{2a}{Rb}$$

T_i = inversion Temperature

- The temperature at which the Joule Thompson effect for a given gas changes sign so that the gas is neither heated nor cooled when allowed to expand without expending energy

$$\textcircled{3} \quad T_B = \frac{1}{2} T_i$$

\Rightarrow Relative Humidity

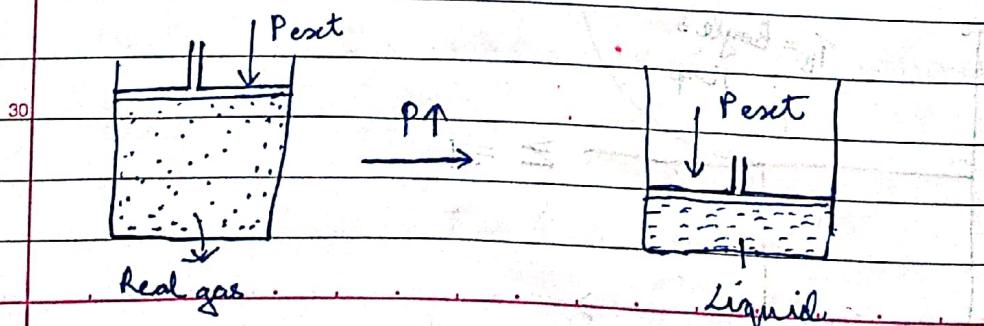
At given temp,

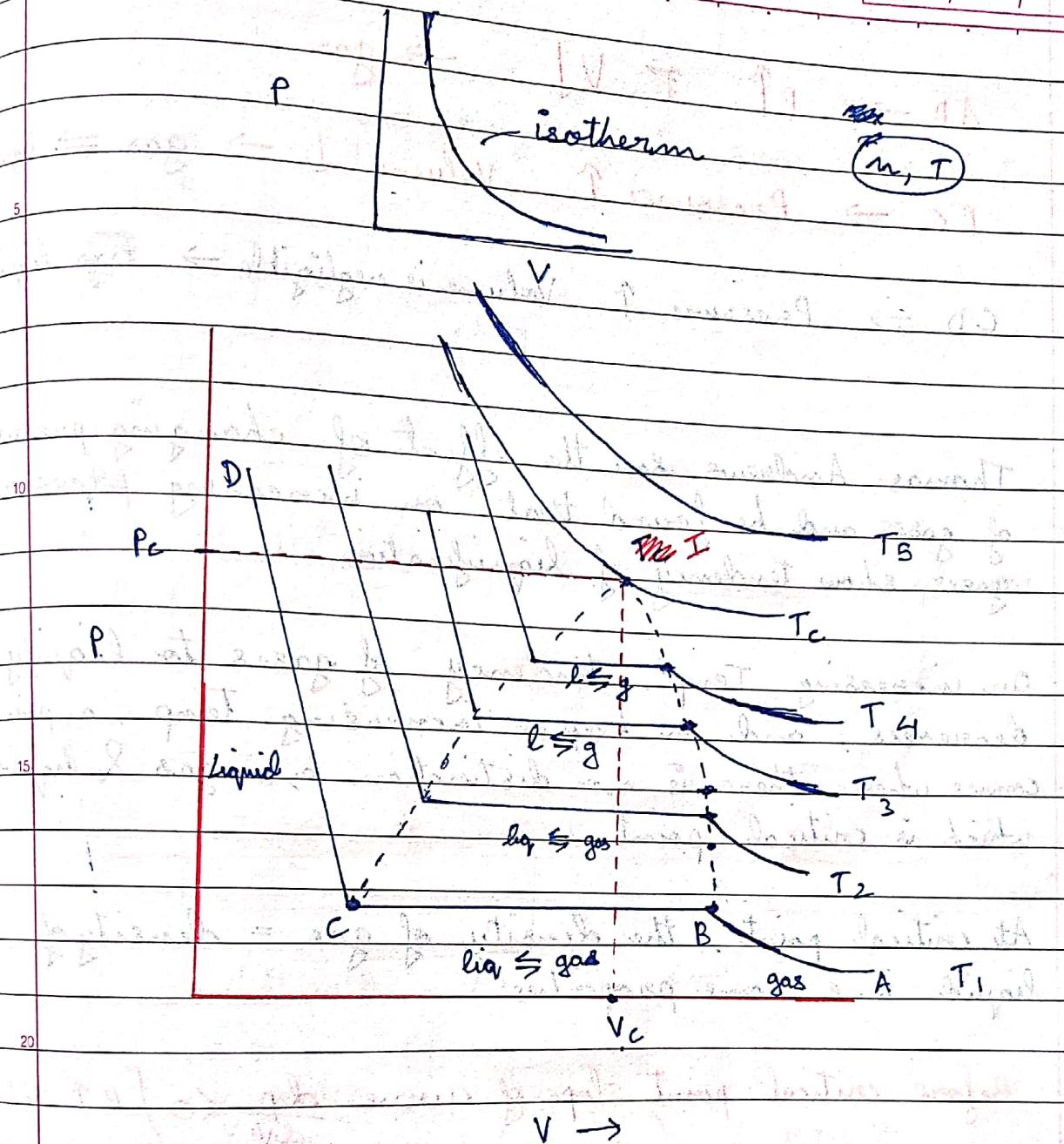
$R_H = \frac{\text{Partial pressure of } H_2O \text{ in air}}{\text{Equilibrium vapour pressure of } H_2O \text{ at this temperature}}$

\Rightarrow Liquification of gases

Critical phenomenon:

CO_2 gas - At different $T \rightarrow$ Andrews Isotherms





B → Dew Point (The last drop to convert to gas)

C → Bubble Point ("first" drop)

$$T_5 > T_c > T_4 > T_3 > T_2 > T_1$$

Max. Temp upto this gas can be liquified

Gas can't
be liquified

$$P_c = \text{Critical Presure}, \quad T_c = \text{Critical Temp}$$

$$V_c = \text{Critical volume}, \quad I = \text{Critical Point}$$

$AB \rightarrow P \uparrow \nparallel V \downarrow \rightarrow \text{gas}$

5 $BC \rightarrow \text{Pressure} \uparrow \text{Volume} \downarrow \rightarrow \text{gas} \rightleftharpoons \text{liquid}$

CD $\rightarrow \text{Pressure} \uparrow \text{Volume is negligible} \rightarrow \cancel{\text{gas}} \text{ liquid}$

- Thomas Andrews saw the effect of changing pressure of gases and he found that on increasing pressure gases show tendency of liquification

- On increasing Temp, tendency of gases to liquify decreases, and on keep on increasing Temp, a point comes where there is no distinction b/w gas & liquid which is critical point

- At critical point, the density of gas = density of liquid i.e. some properties

- Before critical point slope of curve $\frac{dp}{dV} < 0 [P \uparrow V]$

- At critical point slope of curve, $\frac{dp}{dV} = 0 [P = V]$

⇒ 25 Critical Temperature (T_c)

- The temperature above which gas can't be liquified

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

$$B T_b : T_i : T_c = 1 : 2 : \frac{8}{27}$$

Critical Pressure (P_c)

It is the pressure required to cause liquefaction at T_c

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

Critical Volume (V_c)

It is the volume occupied by one mole of gas at T_c and P_c

$$V_c = 3b$$

Determination of critical constants from VGF:

Vanderwall

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

$$(PV_m^2 + a)(V_m - b) = RTV_m^2$$

$$PV_m^3 + aV_m - PV_m b - ab - RTV_m^2 = 0$$

$$PV_m^3 + V_m^2 (-pb - RT) + aV_m - ab = 0$$

$$PV_m^3 - V_m^2 (pb + RT) + aV_m - ab = 0$$

$$V_m^3 - \left(b + \frac{RT}{P} \right) V_m^2 + \frac{a}{P} V_m - \frac{ab}{P} = 0$$

at critical Point

$$T = T_c, V = V_c, P = P_c$$

$$V_m^3 - \left(b + \frac{RT_c}{P_c} \right) V_m^2 + \frac{a}{P_c} \cdot V_m - \frac{ab}{P_c} = 0 \quad (1)$$

V.G.E at Critical point

This eqn. contains 3 roots :-

$$V_m - V_c = 0 \quad (V_m - V_c)^3 = 0$$

$$V_m^3 - 3V_m^2 V_c + 3V_m V_c^2 - V_c^3 = 0 \quad (2)$$

Compare (1) & (2)

$$3V_c = \left(b + \frac{RT_c}{P_c} \right) \quad (3)$$

$$3V_c^2 = \frac{a}{P_c} \quad (4)$$

$$V_c^3 = \frac{ab}{P_c} \quad (5)$$

(3)
(4)

$$\frac{V_c}{3V} = \frac{ab}{P_c} \times \frac{P_c}{ab}$$

$$V_c = 3b \quad (6)$$

(5)
(6)

(6) in (5)

$$(3b)^3 = \frac{ab}{P_c}$$

$$TP_c V = \frac{ab}{27b^3}$$

$$P_c = \frac{a}{27b^2} \quad (7)$$

(7), (6) in (3)

$$3(3b) = \frac{\frac{a}{27b^2} \times b}{a} + RT_c$$

$$\frac{a}{27b^2} \times 9b = \frac{a}{27b} + RT_c$$

$$RT_c = \frac{a}{3b} - \frac{a}{27b} = \frac{8a}{27b}$$

$$T_c = \frac{8}{27 R b}$$

$$Z = \frac{P_c V_c}{R T_c}$$

$$= \frac{\frac{d}{27 b^2} \times 3b}{R \times \frac{8ac}{27 K b}}$$

$$Z = \frac{3}{8}$$

⇒ Vanderwaal's constant in terms of V_c, T_c, P_c

$$V_c = 3b$$

$$b = \frac{V_c}{3}$$

$$P_c = \frac{a}{27 b^2} = \frac{a}{27 \left(\frac{V_c}{3}\right)^2} = \frac{a}{27 \frac{V_c^2}{9}} = \frac{a}{3 V_c^2}$$

$$P_c = \frac{a}{3 V_c^2}$$

$$a = 3 V_c^2 P_c$$

5

10

15

⇒ Reduced Equation of State

⇒ Reduced Temperature (T_r):

Temp. in any state of gas wrt critical temp
is known as reduced temp.

$$T_r = \frac{T}{T_c}$$

25

$$P_r = \frac{P}{P_c}$$

$$V_r = \frac{V}{V_c}$$

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

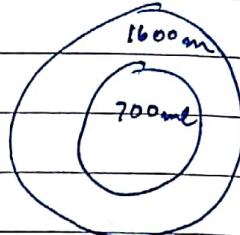
$$(P_r \cdot P_c + \frac{an^2}{V^2}) (V_r \cdot V_c - nb) = nRT$$

30

5 Q) Balloon A is inserted in balloon B at 300 K. Maximum capacity of A and B are 800 ml and 1800 ml. When this system is uniformly heated which balloon will burst first?

5

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



For Inner

10

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

15

$$\frac{700}{300} = \frac{800}{T_2}$$

$$T_2 = \frac{2400}{7}$$

20

$$T_2 = 342.8 \text{ K}$$

For Outer

25

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

$$\frac{800}{300} = \frac{1800}{T_2}$$

30

$$T_2 = \frac{5400}{16} = 337.5 \text{ K}$$

∴ Outer will ~~burst~~ burst first

(a)

Level - I (Pg - 38)

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

$$\frac{P_1 V}{T} = R$$

$$\frac{0.418}{300} = \frac{V_2}{273}$$

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

$$\cancel{V_2} =$$

$$\frac{0.418 \times \frac{74}{76}}{300} = \frac{1 \times V_2}{273}$$

$$V_2 = \frac{418 \times 74 \times 273}{100 \times 300 \times 76}$$

$$\frac{76}{3} \\ 228$$

$$= 0.37 \text{ Liters}$$

$$(b) M.W = ? , W = 3.08$$

$$P V = n R T$$

$$\frac{740}{760} \times \frac{0.418}{300} = \frac{3}{M.W} \times 0.0821$$

$$M.W = \frac{760 \times 300 \times 3 \times 0.0821}{740 \times 0.418}$$

$$= \frac{760 \times 300 \times 3 \times 82.1}{74 \times 418}$$

$$M.W = \frac{76 \times 90 \cancel{8} \times 82}{\cancel{74} \times 418}$$

$$M.W = 1.98 \times 90$$

$$= \cancel{17.82} \approx 178.2$$

$$= 181.6$$

$$c) W_2 = 7.5 \text{ g}$$

$$T_2 = 280 \text{ K}$$

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

$$\frac{\cancel{74} \times \cancel{418}}{\cancel{76} \times 100} \times \frac{\cancel{76} \times 90 \times 82}{\cancel{74} \times \cancel{418}} =$$

$$\frac{\cancel{74} \times \cancel{418}}{\cancel{76} \times 100} = \frac{T P_2 \times 0.418}{\cancel{7.5} \times 280}$$

$$\frac{\cancel{74} \times \cancel{W_1} \times \cancel{418} \times 1}{\cancel{76} \times \cancel{W_1} \times 100 \times 300} = \frac{P_2 \times 418 \times 10 \times M.W}{100 \times 280 \times 75}$$

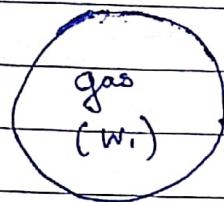
$$\frac{\cancel{74} \times \cancel{418}}{\cancel{76} \times \cancel{W_1} \times 300} = \frac{P_2 \times 418}{280 \times 75}$$

$$P_2 = \frac{74 \times 75 \times 28}{76 \times 900}$$

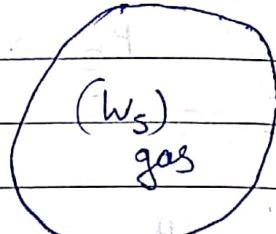
$$P_2 = \frac{74 \times 21}{76 \times 9}$$

$$\therefore P_2 = 0.97 \times 2.33 \\ = 2.2601 \text{ atm}$$

⇒ Payload :-



some amount
of gas displaced

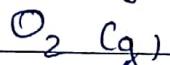
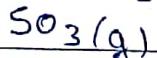
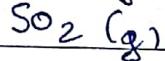
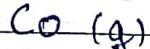
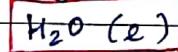
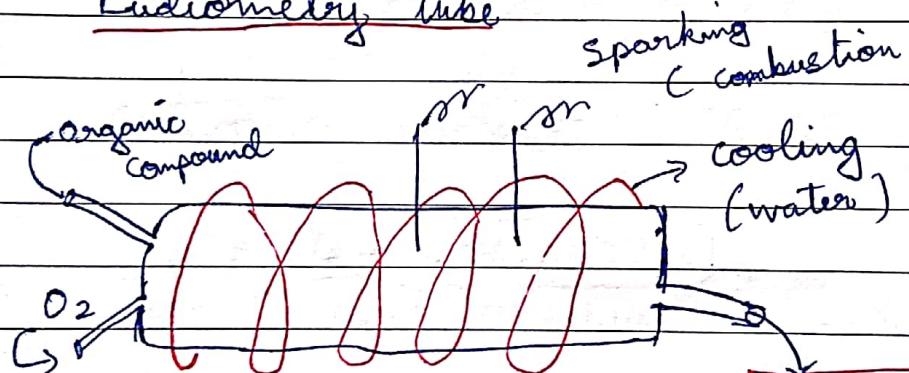


Balloon (w_2)

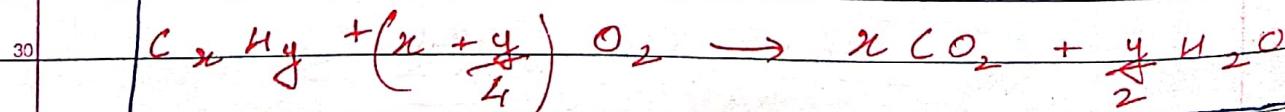
$$\text{Payload} = \frac{\text{Weight of the gas displaced}}{\text{Weight of } w_1 + \text{weight of balloon}}$$

⇒ Endiometry (or) Gas Analysis

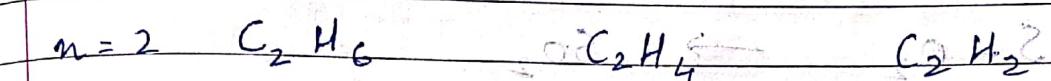
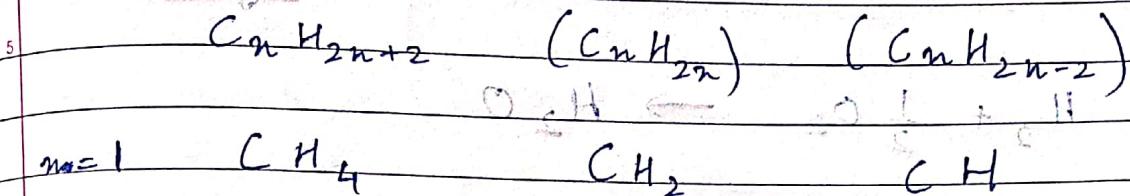
Endiometry tube



Combustion :

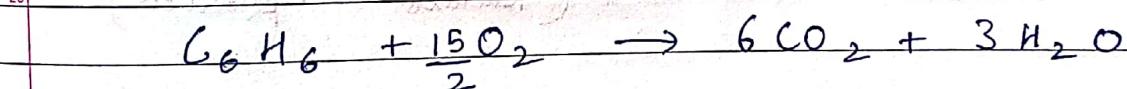


Alkanes Alkenes Alkynes
 $(-\text{H})$ $(=\text{H})$ $(\equiv \text{H})$

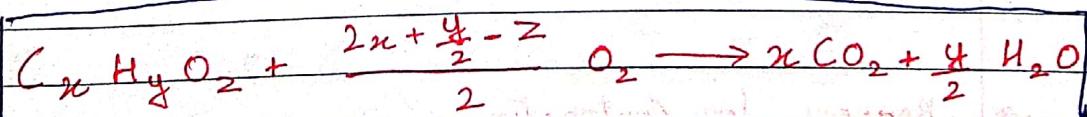


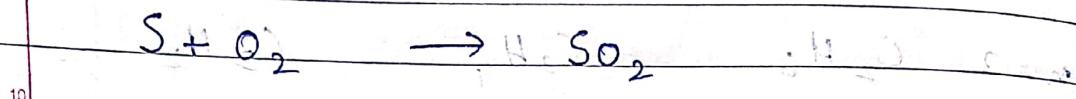
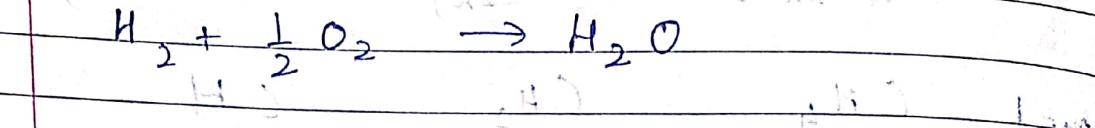
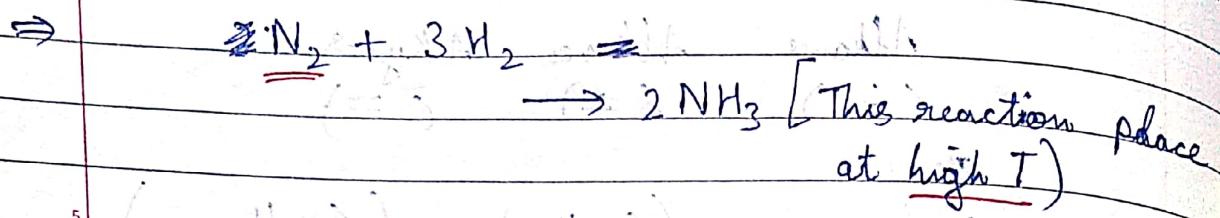
x.

Benzene



$\Rightarrow_{25} \text{C}_x\text{H}_y\text{O}_z$





⇒ Eudiometry (or) gas analysis

• Contraction: [decrease in volume]

$$\text{Contraction} = V_{\text{initial}} - V_{\text{final}}$$

$$\boxed{\Delta V = V_I - V_F}$$
 ($V_{\text{reactants}} - V_{\text{products}}$)

• Expansion: [Increase in volume]

$$\boxed{\Delta V = V_F - V_I}$$

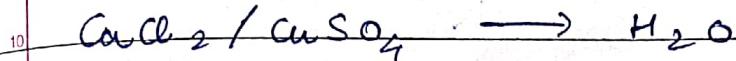
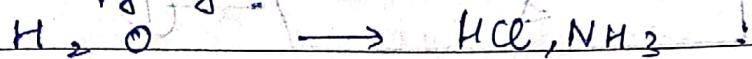
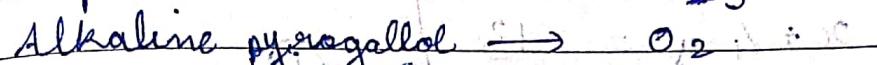
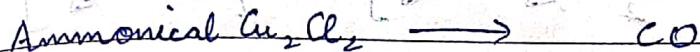
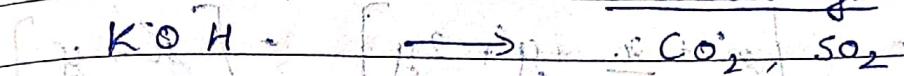
⇒ Reasons for Contraction

gas $\xrightarrow{\text{Expansion}}$ liquid / gas

Contraction

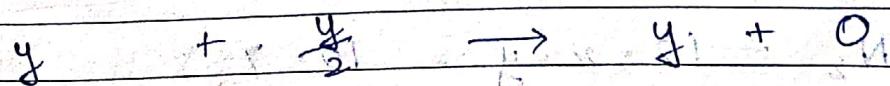
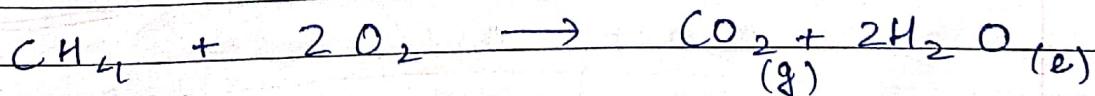
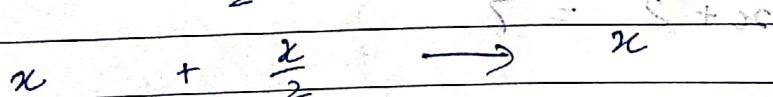
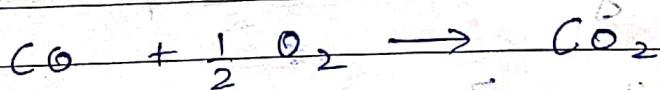
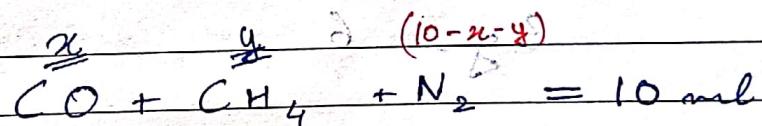
Some compounds absorb specific gases

Absorbed gases



Impact of acids NOx

Q) 10 ml of $(\text{CO}, \text{CH}_4, \text{N}_2)$ explodes with excess of (O_2) giving a contraction of 6.5 ml, there was a further contraction of 7 ml when treated with KOH. Find volume of $\text{CO}, \text{CH}_4, \text{N}_2$



$(10 - x - y)$

$$\Delta V = V_{\text{Initial}} - V_{\text{Final}}$$

$$6.5 = \left[x + \frac{x}{2} + y + 2y \right] - [x + y]$$

$$2x + 4y = 13 \quad \boxed{1}$$

KOH absorbs $\text{CO}_2 = 7 \text{ ml}$

$$x + y = 7 \text{ ml} \quad \boxed{2}$$

$$\boxed{1} - \boxed{2}$$

$$3y = 6$$

$$y = 2 \text{ ml} = \text{CH}_4$$

$$x + 2 = 7$$

$$x = 5 \text{ ml} = \text{CO}$$

$$\text{N}_2 = 10 - x - y = 10 - 7 + 2$$

$$= 3 \text{ ml}$$

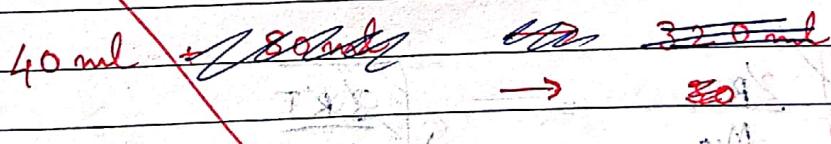
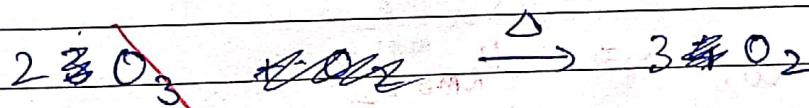
(a) When 100 ml of (O_2, O_3) mixture was passed through turpsentine oil there was reduction of volume by 20 ml. If 100 ml of this mixture is heated what will be the increase in volume.

$$O_2 + O_3 = 100 \text{ ml}$$

Passed through turpsentine oil.

$$\therefore O_3 \text{ absorbed} = 20 \text{ ml}$$

$$\therefore O_2 = 80 \text{ ml}$$



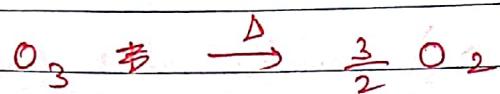
$$\Delta V = V_f - V_i$$

$$= 320 - 120 = 200 \text{ ml}$$

$$\text{Increase in volume} = 200 \text{ ml}$$



$$80 \text{ ml}$$



$$20 \text{ ml} \rightarrow \left[\frac{3}{2} \times 20 \right] \text{ } 30 \text{ ml}$$

$$80 + 30 = 110 \text{ ml}$$

$$\therefore \Delta V = 10 \text{ ml}$$

a) Calc. the T at which CO_2 has same Vrms as O_2 at STP

$$V_{rms, CO_2} = V_{rms, O_2}$$

$$\sqrt{\frac{3RT}{M_{CO_2}}} = \sqrt{\frac{3RT}{M_{O_2}}}$$

$$M_{CO_2} \times 0.0273 = M_{O_2} \times T$$

$$T = \frac{273 \times 44}{32}$$

168

$$T = 375.1 \text{ K}$$

a) calculate Z for CO_2 If one mole of it occupies 0.4 liter at 300 K and 40 atm, comment on the result.

$$Z = \frac{PV}{RT}$$

$$Z = \frac{40 \times 0.4}{0.0821 \times 300}$$

$$= \frac{16}{821 \times 3} = \frac{16}{2463} = 0.0065$$

$$\therefore Z = 0.07$$

From above $Z < 1$, it shows negative deviation, so it is easily liquifiable

Q) 50 gms of dry N_2 is passed through 36 gms of 300 K of H_2O

There is a loss of 11.20 gm in H_2O and calculate vapour pressure of H_2O ?

$$pV = nRT$$

$$pV = \frac{W}{M} \times RT$$

$$p \times 50 = \frac{12}{180} \times \frac{0.19}{12} \times 300$$

$$p = \frac{3}{90}$$

$$p = \frac{1}{30} \times \frac{100}{10} = \frac{0.033}{10}$$

$$p = 0.033 \text{ atm}$$

$$\frac{0.033}{760} \text{ mm Hg}$$

$$1 \text{ atm} = 760 \text{ mm of Hg}$$

~~$$\frac{33}{76000} = 0.033$$~~

$$0.033 = 760 \times 0.033$$

~~$$= 23.2 \text{ mm of Hg}$$~~

CO_2 is found to have a $d = 1.20 \text{ g/l}$ at 30°C
~~730 mm~~ what is the composition?

$$d = 1.112 \text{ g/l}$$

~~$T = 303 \text{ K}$~~

$$P = \frac{730}{760} - \frac{73}{76} \text{ atm}$$

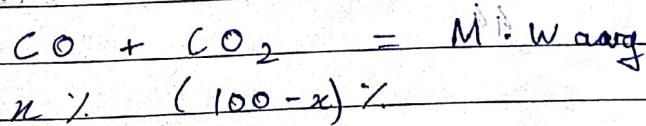
$$MW = \frac{d \cdot R \cdot T}{P}$$

$$= \frac{1.112 \times 0.082 \frac{1}{12} \times 303}{10}$$

$$MW = \frac{303 \times 76 \times 1.5 \times 821}{730 \times 1000}$$

$$= \frac{23028 \times 821 \times 1.5}{730 \times 100}$$

$$= 38.8 \text{ g}$$



$$38.8 = \frac{28x + 44(100-x)}{100}$$

$$3885 = -16x + 4400$$

$$16x = 515$$

$$\frac{16x = 515}{16}$$

$$CO(x) = 32.1\%$$

$$CO_2 = 67.9\%$$

- Q) A 20g chunk of dry ice is placed in an empty 0.75 l vial bottle tightly closed what would be the final pressure after one CO₂ has evaporated and T = 25°C

Really high.

$$W = 20 \text{ g}, V = 0.75 \text{ l}$$

$$PV = \frac{20 \times 0.0821 \times 298}{44.1} \text{ K}$$

$$P = \frac{20 \times 29.8 \times 821 \times 100}{44 \times 75 \times 1000}$$

$$= 14.827 \text{ atm}$$