EXAMS ROAD

"Bringing Excellence to Students"

Handwritten Notes

On

States of Matter

States of Matter

① Solid — Liquid — gas

Intermolecular space ↑↑ → Intermolecular F·O·A ↓↓ → Thermal Energy ↑↑

② Parameters of Gas

Pressure :

1 atm = 760 mm of Hg
= 76 cm of Hg.
= 760 torr
= 1·013 bar

Volume :

$1ml = 1cm^3 = 1cc$ $\quad 1L = 1dm^3$

$1m^3 = 1000l$

Temp. –

$T_K = 273 + T°_C$

$\frac{F-32}{9} = \frac{C}{5}$

③ Kinetic Theory of Gases :

(i) Actual Vol. occupied by gas molecule is negligible in comparision to empty space b/w them.

(ii) there is no force of attraction working b/w gas molecules.

(iii) Particles of Gas are always in Constant random motion.

(iv) Perfectly elastic collision ∴ there is h

v) Avg. K.E of the gaseous molecule is directly proportional to absolute temp.

④ Gas Laws

Boyle's Law – Isotherm

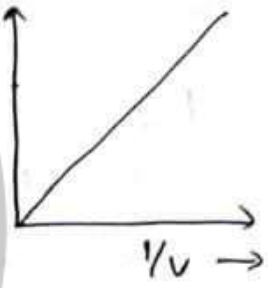
⇓

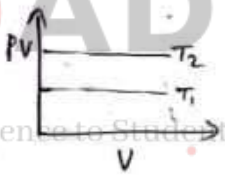
At Constant Temperature

$P \propto \frac{1}{V}$ [at constant n & P]

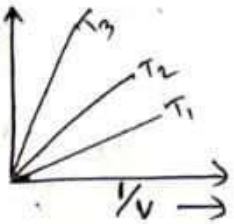
$PV = Constant$ $\quad P_1V_1 = P_2V_2$

Graph.





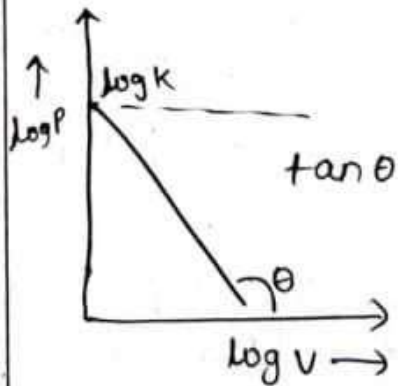
$T_2 > T_1$



$T_3 > T_2 > T_1$

$PV = K$

$\log_{\downarrow y} P = \log_{\downarrow c} K - \log_{\downarrow mc} V$



⑤ Charle's Law :

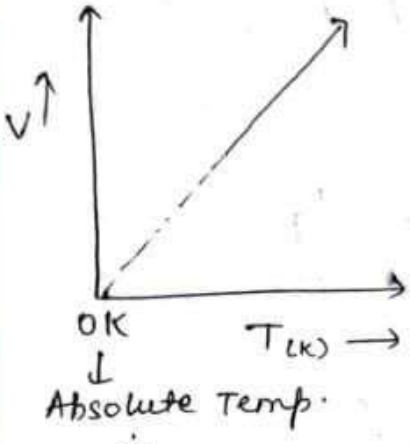
⇓

At Constant Pressure

$V \propto T_{in\ K}$

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

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

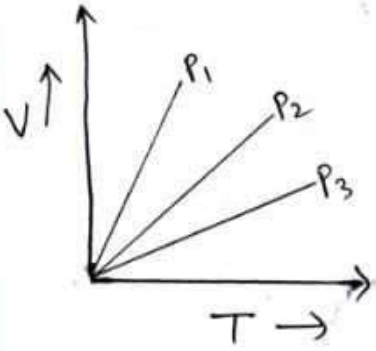


∴

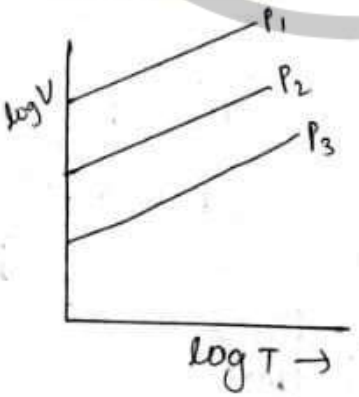
Note : Each degree Rise in temp. Vol. of gas Increase by 1/273 of the original Vol. of gas.

Let $0°c \rightarrow V_0$ $\quad t° \rightarrow V_t$

$$V_t = V_0\left[1 + \frac{t}{273}\right]$$



$P_3 > P_2 > P_1$



$P_3 > P_2 > P_1$

Each line of Vol. V/s Temp. graph is called → **Isobar**

⑥ Gay Lussac's Law

⇓

At Constant Volume

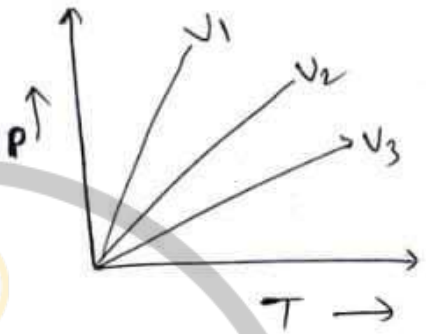
$P \propto T$

$\frac{P}{T} = \text{Constant}$

Each line of this graph

⇓

Isochore



$V_3 > V_2 > V_1$

⑦ Avogadro Law

It states that equal Vol. of all gases under the same Cond^n of temp. & pressure contain equal no. of molecules.

$V \propto n$

$PV = nRT$

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

$PM = dRT$

At STP/NTP : 1 mole of any gas occupies 22.4 L

$T = 273K$

$P = 1\ atm$

$R \rightarrow 0.0821\ L\ atm\ mol^{-1}\ K^{-1}$

$R \rightarrow 0.083\ L\ bar\ mol^{-1}\ K^{-1}$

(8) Dalton's law of Partial Pressure :-

Total pressure exerted by mixture of Non-reactive gases is equal to sum of partial pressure of individual gases

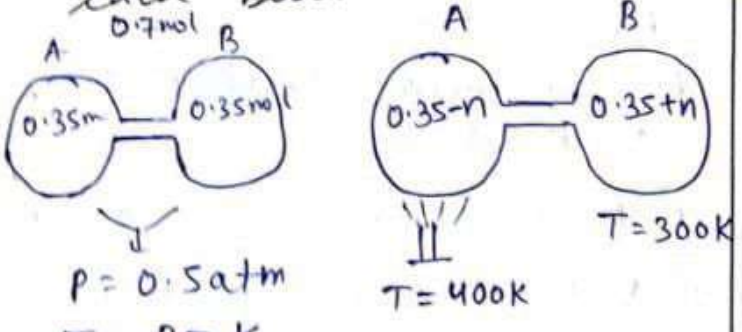
$$P_{Total} = P_1 + P_2 + P_3 + \dots$$

$$P_1 = \frac{n_1 RT}{V} \qquad P_2 = \frac{n_2 RT}{V} \quad \dots$$

$$P_1 = P_{total} \times X_{gas}$$

Partial Pressure of any gas $= X_{gas} \times P_{total}$

Qn) Two bulbs of equal vol. connected through stop clock contained 0.7 mol of H_2 gas at 0.5 atm pressure at 27°C (at open posn of stop clock). If first bulb is heated to 127°C keeping the other at same temp. i.e 27°C. What will be final pressure & moles in each bulb?



$$PV = nRT \quad (\text{Same})$$

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

Bulb A:

$$\frac{0.5}{P} = \frac{0.35 \times 300}{(0.35-n)\,400} \quad \text{— (I)}$$

Bulb B

$$\frac{0.5}{P} = \frac{0.35 \times 300}{(0.35+n)\,300} \quad \text{— (II)}$$

From (I) & (II) we get.

Note :- Daltons law of Partial pressure is applicable for → Non Reacting gases only

(9) Graham's law of Diffusion / effusion :-

Rate of diffusion/effusion $\propto \frac{1}{\sqrt{\text{density}}}$.

$$R \propto \frac{1}{\sqrt{d}}$$

$$R \propto \frac{1}{\sqrt{M}} \qquad M \propto d$$

Lighter gas → Rate of diffusion ↑↑

Heavy gas → Rate of diffusion ↓↓.

$$\text{Rate of diffusion/effusion} = \frac{\text{Dist. cov.}}{\text{Time}}$$

$$= \frac{\text{Change in Pressure}}{\text{Time}}.$$

$$= \frac{\text{Change in vol.}}{\text{Time}}.$$

$$= \frac{\text{no. of Moles of gas effused}}{\text{Time}}.$$

On) NH_3 | N_2

$V = 100ml$ | $V = 60ml$

$t = 32.5 sec$ | $t = ?$.

$$R \propto \frac{1}{\sqrt{M}}$$

$$\frac{\Delta V}{t} \propto \frac{1}{\sqrt{M}}$$

$$\frac{V_{NH_3}}{V_{N_2}} \times \frac{t_{N_2}}{t_{NH_3}} = \sqrt{\frac{M_{N_2}}{M_{NH_3}}}$$

$$\frac{100}{60} \times \frac{t_{N_2}}{32.5} = \sqrt{\frac{28}{17}}$$

$\Rightarrow$

(10) Diff. types of Molecular Speed

Most probable Velocity –

$$V_{mp} = \sqrt{\frac{2RT}{M}}$$

Root Mean Square Velocity :

$$V_{rms} = \sqrt{\frac{3RT}{M}}$$

Avg. Vel :

$$V_{avg.} = \sqrt{\frac{8RT}{\pi M}}$$

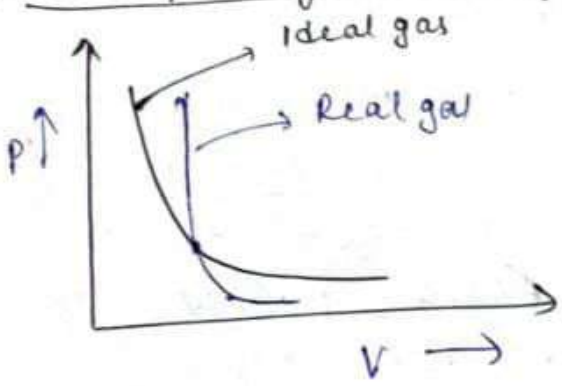
On) A balloon filled with Ethylene is pricked with needle & quickly dropped in tank of Hydrogen under identical cond^n then balloon will –

a) Shrink b) Enlarge

c) Collapse d) Remain unchanged

$R \propto \frac{1}{\sqrt{M}}$

(11) Concept of Real gas :



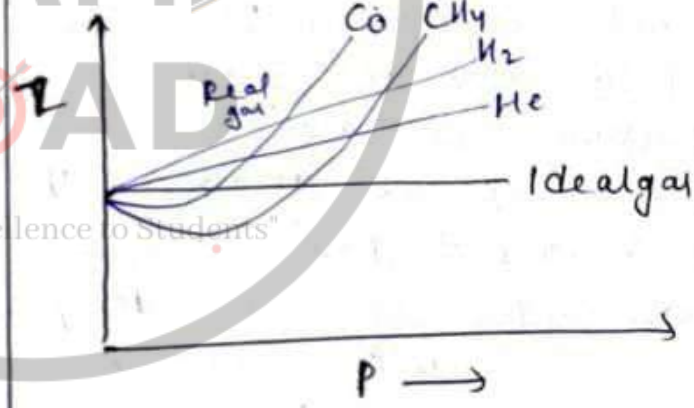
Compressibility factor (Z)

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

For Ideal gas $Z = 1$

Non Real gas $Z \neq 1$

$Z > 1$ → Real gas showing the deviation & vice versa.



Real gas do not obey Ideal gas :
due to two faulty assumpt of K.T.G :

i) There is no force of attrac

ii) Vol. of molecules of gas is negligible small in comparision to the space occupied by gas.

$P_{correct} = \underset{\text{obs.}}{P} + \frac{n^2a}{V^2}$

a → Measure Intermolecular F.O.A.

$V_{correct} = V - nb$

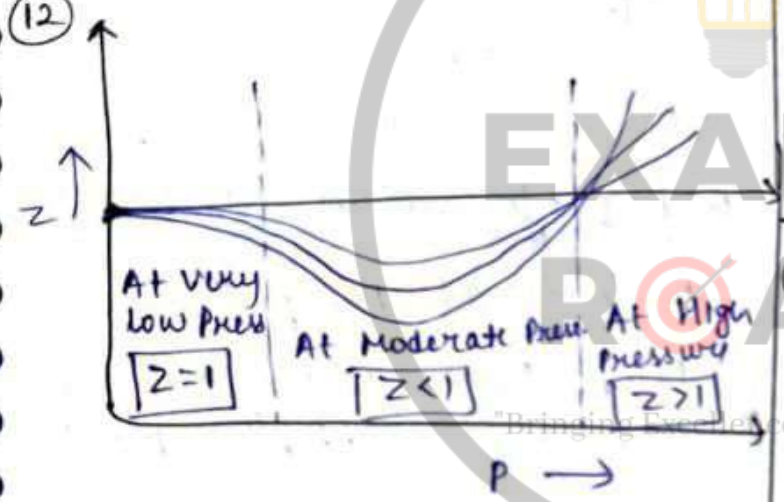
b → measure size of Molecule

$PV = nRT$

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

↳ Real gas eqn.

(12)



Case I :- At Very Low Pressure

$PV = nRT$

At Very low Pressure Real gas obey Ideal gas eqn.

Case II :- At Moderate Pressure

$\left(P + \frac{n^2a}{V^2}\right)(V - \underset{x}{nb}) = nRT$

$\frac{PV}{nRT} + \frac{n^2a}{VnRT} = 1$

$Z = 1 - \frac{na}{}$ (Z<1)

Case III :- At High Pressure

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

$\frac{PV}{nRT} - \frac{Pnb}{nRT} = 1$

$Z = 1 + \frac{Pb}{RT}$ (Z>1)

Case IV :- At High Temp.

$V \propto T\uparrow\uparrow$

Vol. → ↑↑

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

$PV = nRT$

∴ At Very low Pressure & High temp. Real gas obey Ideal gas eqn.

(13) Exceptional behaviour of H_2 & He

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

$Z > 1$ Always.

∴ a is too small

(14) Application of Compressibility factor :-

1. $Z = 1$ → It means gas is ideal at all temp & pressure.

2. $Z > 1$ → $PV > PV_{ideal}$

All gas at high pressure are less compressible than Ideal gas.

3. $Z < 1$ → $PV_{real} < PV_{ideal}$

These gases are more compressible than Ideal gas.

$$Z \propto \frac{1}{\text{Compressible}}$$

$$Z = \frac{V_{real}}{V_{ideal}}$$

$Z = 1$

$V_{real} = V_{ideal}$

$V_{real} = V_{ideal}$ = 22.4 L for 1 mole of gas at STP

$Z > 1$

$\frac{V_{real}}{V_{ideal}} > 1$

$V_{real} > V_{ideal}$

$V_{real} > 22.4\ L$

Less Compressible

$Z < 1$

$\frac{V_{real}}{V_{ideal}} < 1$

$V_{real} < V_{ideal}$

More Compressible

$V_{real} < 22.4\ L$

(15) Value of R :-

R → 0.0821 L atm $mol^{-1} K^{-1}$
→ 0.083 L bar $mol^{-1} K^{-1}$
→ 2 Cal $mol^{-1} K^{-1}$
→ 8.314 J $mol^{-1} K^{-1}$

(16) $K.E_{avg} = \frac{3}{2} RT$

(17) Boyles Temp. (T_b)

Temp. at which real gas obey / behave Ideally. (large temp.)

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

(18) Liquification of Gases

I Critical Temp. (T_c)

Temp. upto which gas can be Liquify applying pressure

$T_{gas} < T_c$

⇓

Liquification of Gas Possible.

$T_{gas} > T_c$

⇓

No chance of Liquification

II Critical Pressure (P_c)

Press. req. to liquify gas at Critical temp. is called P_c.

III Critical Vol. (V_c)

Vol. occupied by 1 mol of gas at ... press.

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

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

$$V_c = 3b$$

$$\frac{P_c V_c}{T_c} = \frac{3}{8}R$$

Critical Compressibility factor (Z_c)

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

$n = 1$

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

$$Z_c \Rightarrow \frac{3}{8}$$