

## States of Matter

Page No.

Date

- Matter - having mass and occupying space.

① Solid state

② Liquid state ✓

③ Gaseous state. ✓

- The states of matter depends on two factors

① Intermolecular Forces. (Vanderwaal's forces)

② Thermal energy. (Measure of KE)

Vibration & rotation of molecules

• Vanderwaal's forces:

① London Forces / Dispersion Forces — (non polar molecules -  $\text{CO}_2$ ,  $\text{Cl}_2$ )

② Dipole - Dipole interaction. — (polar molecules -  $\text{HCl}$ ,  $\text{SO}_2$ ,  $\text{H}_2\text{S}$ )

③ Dipole - Induced dipole. (Polar - non polar molecule -  $\text{HCl} + \text{Cl}_2$ )

• Interaction energy :

$$\textcircled{1} \quad \text{London} \propto \frac{1}{r^6}$$

$$\textcircled{2} \quad \text{Dipole - induced dipole} \propto \frac{1}{r^6}$$

$$\textcircled{3} \quad \text{Dipole-Dipole} \propto \frac{1}{r^3} \quad (\text{stationary}) \propto \frac{1}{r^6} \quad (\text{rotating})$$

• Thermal energy : (K.E)

T.E  $\propto T$

Due to motion.

- The states of matter is the result of balance between Intermolecular forces and Thermal energy.

- If T.E is dominant,  
Solid  $\rightarrow$  Liquid  $\rightarrow$  Gas.

- If IMF is dominant,  
Gas  $\rightarrow$  Liquid  $\rightarrow$  Solid.

- Gas laws: Gas laws inter relate the measurable properties of gases. (P, V, T and n)

### ① Boyle's law

n, T constant.

$$V \propto 1 \quad \text{cancel out no absoprtion system to isolate out}$$

$$P \cdot V = \text{constant} \quad (PV = K \text{ depends on } n \text{ and } T)$$

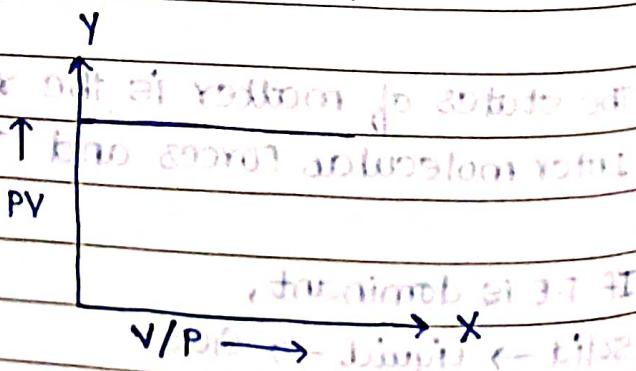
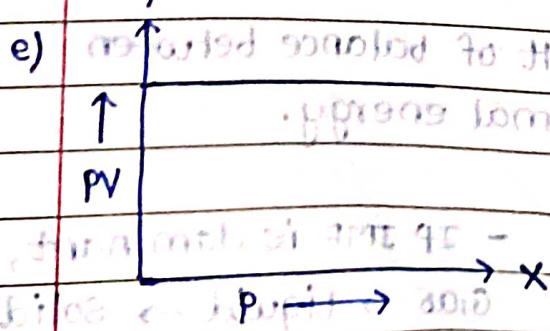
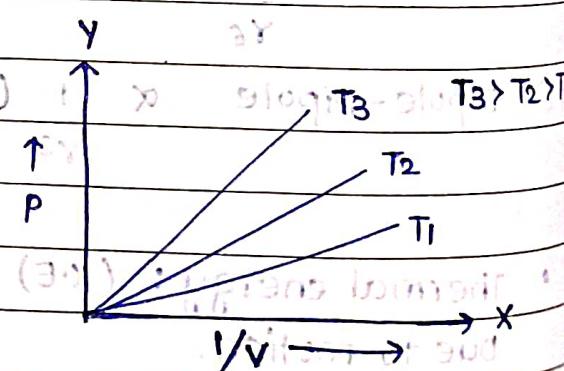
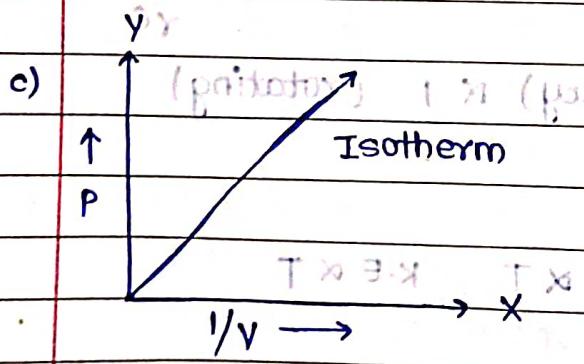
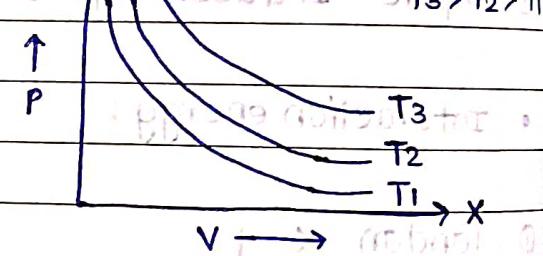
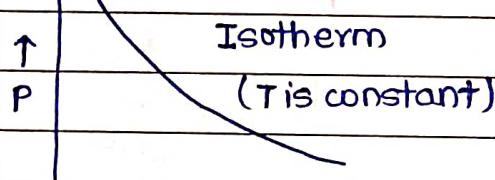
$$P_1 V_1 = P_2 V_2$$

also on units of P and V.

(s) ① Graphs: (Boyle's law) - direct proportionality

(real,  $\text{CH}_4$  - extension analog) --- (ideal gas - ideal gas)

(s) ② (T) - (Boyle's law analog) - b) (ideal gas) -  $T_3 > T_2 > T_1$



② Charles' law:

$P, n$  constant.

$$V \propto T$$

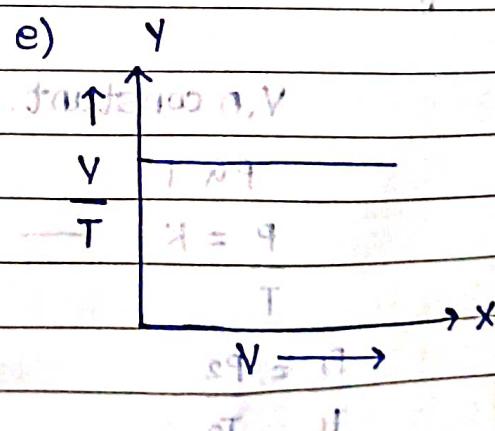
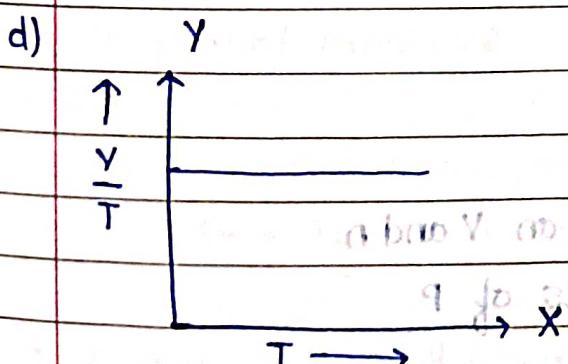
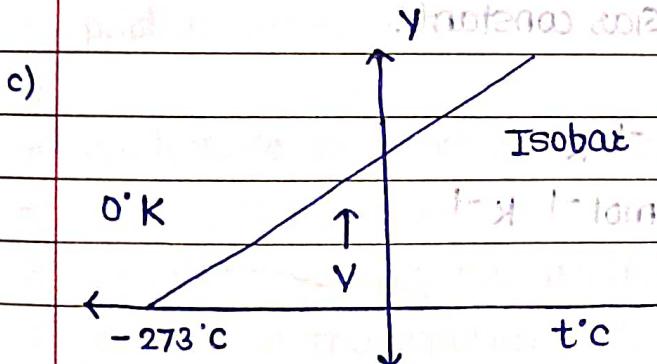
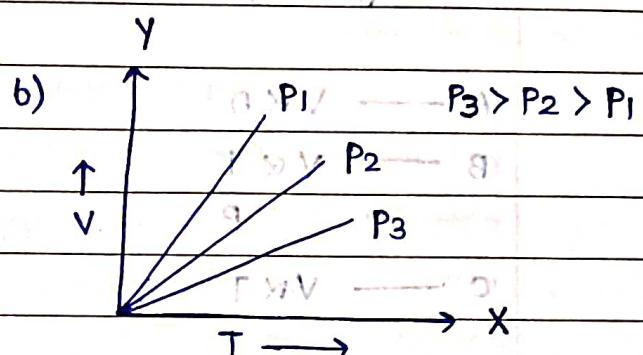
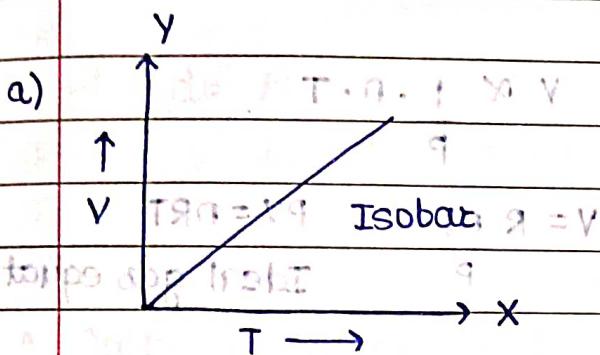
$$\frac{V}{T} = K$$

and units of volume.

$$V_1 = V_2$$

Isobari  $T_1 < T_2 \rightarrow T$  must be taken in Kelvin only.

• Graphs:



③ Avogadro's law:

P, T constant

$V \propto n$

$$\frac{V}{n} = K \quad \text{depends on } P \text{ and } T$$

$\frac{V_1}{n_1} = \frac{V_2}{n_2}$  ————— above and units of V. —————  $K = V$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{law of fraction law}$$

$n_1 \quad n_2$

$$sV = IV$$

The gas which strictly obeys ABC is known as Ideal gas.

- Ideal gas equation:

$$A \longrightarrow V \propto n$$

$$B \longrightarrow V \propto I$$

$$C \longrightarrow V \propto T$$

$$\therefore V \propto 1 \cdot n \cdot T$$

$$V = R n T \quad PV = nRT$$

Ideal gas equation

R = Universal Gas constant:

$$R = 8.314 \text{ Joule mol}^{-1} \text{ K}^{-1}$$

$$R = 0.0821 \text{ lit atm mol}^{-1} \text{ K}^{-1}$$

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

- Gay Lussac's law:

V, n constant.

$$P \propto T$$

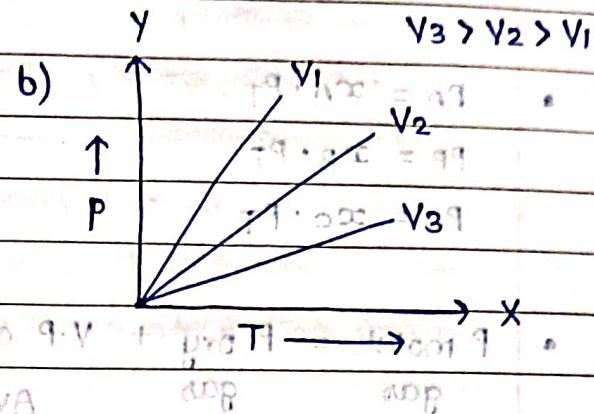
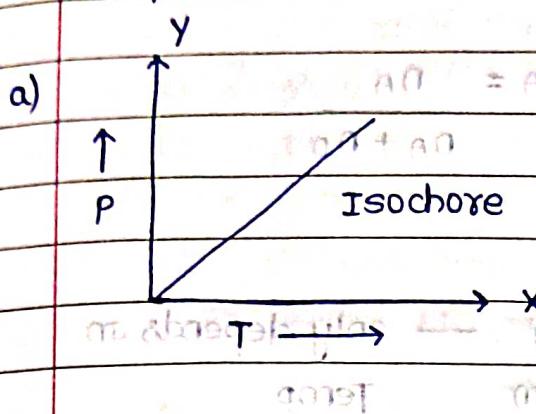
$$\frac{P}{T} = K \quad \text{depends on } V \text{ and } n$$

and units of P

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

$T_1 \quad T_2$

- Graphs:



- Density of ideal gas:

- combined gas law:

$$d = \frac{PM}{RT} \text{ or } d = \frac{m}{V} \text{ or } PV = nRT \text{ or } P = \frac{nRT}{V}$$

finding no. of moles to get  $n = \frac{PV}{RT}$  from  $PV = nRT$  then  $n = \frac{P_1V_1}{RT_1}$  and  $n = \frac{P_2V_2}{RT_2}$  for nitrogen

$$\text{If } n_1 = n_2 \Rightarrow T \text{ has to be same}$$

$$\text{Assume to get same no. of moles: } P_1V_1 = P_2V_2 \text{ or } P_1V_1/T_1 = P_2V_2/T_2$$

- Dalton's law of partial pressure: "partial pressure" is dependent on concentration

- applicable to mixture of non-reacting gases

- The total pressure exerted by a gaseous state mixture is equal to sum of the partial pressures of all the gases present in the mixture.

- The pressure exerted by an individual gas in the mixture is partial pressure.

$$P_T = P_A + P_B + P_C$$

- On applying ideal gas equation,

$$\frac{P_T}{V} = \frac{n_T RT}{V} = P_T$$

$$\frac{P_A}{V} = \frac{n_A RT}{V} = P_A$$

$$\frac{P_B}{V} = \frac{n_B RT}{V} = P_B$$

$$\frac{P_C}{V} = \frac{n_C RT}{V} = P_C$$

$$\bullet \quad n_T = n_A + n_B + n_C$$

~~N < V < T~~

$$\bullet \quad P_A = \chi_A \cdot P_T$$

$$P_B = \chi_B \cdot P_T$$

$$P_C = \chi_C \cdot P_T$$

$$\chi_A = n_A$$

$$n_A + n_B$$

$$n_A + n_D$$

- $P_{\text{moist}} = P_{\text{dry}} + V \cdot P_{\text{of } H_2O}$  or — only depends on gas gas Avg. Tension Temp

- Graham's law :

The rate of diffusion or effusion of a gas is inversely proportional to molecular mass / density of gas at constant P and T.

- Diffusion - The spontaneous intermixing of gases.

- effusion - The diffusion of a gas from high P to low P through a 'orifice' : ~~inversely~~ ~~proportional~~ to density

$$\gamma \propto \frac{1}{M} \quad \text{or} \quad \gamma \propto \frac{1}{\rho} \quad \text{or} \quad \text{Rate of diffusion} \propto \frac{1}{\sqrt{M}}$$

$$\text{Rate of diffusion} \propto \frac{1}{\sqrt{M}} \quad \text{or} \quad \text{Rate of effusion} \propto \frac{1}{\sqrt{M}}$$

$$\gamma_1 = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{\rho_2}{\rho_1}} \quad (P, T) = \text{constant}$$

$$\gamma_2 = \sqrt{\frac{M_1}{M_2}} = \sqrt{\frac{\rho_1}{\rho_2}}$$

① If P constant

$$\gamma \propto \sqrt{T} \quad \gamma_1 = \sqrt{\frac{T_1}{T_2}} \times \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{T_1 \times M_2}{T_2 \times M_1}} = \sqrt{\frac{T_1 \times \rho_2}{T_2 \times \rho_1}}$$

$$\gamma_1 = \sqrt{\frac{T_1}{T_2}} \times \sqrt{\frac{\rho_2}{\rho_1}} = \sqrt{\frac{T_1 \times \rho_2}{T_2 \times \rho_1}}$$

$$T_1 + T_2 + T_3 = T_f$$

② If T constant

$$\gamma \propto P \quad \gamma_1 = \frac{P_1}{P_2} \times \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{P_1}{P_2}} \times \sqrt{\frac{\rho_2}{\rho_1}} = \sqrt{\frac{P_1 \times \rho_2}{P_2 \times \rho_1}}$$

$$\frac{T_1 \times \rho_1}{V_1} = P_1$$

$$\frac{T_2 \times \rho_2}{V_2} = P_2$$

$$\frac{T_3 \times \rho_3}{V_3} = P_3$$

$$\frac{T_f \times \rho_f}{V_f} = P_f$$

$$\gamma = \frac{\text{vol. of gas diffused / effused}}{\text{time taken}} = \frac{V}{t}$$

$$\frac{\gamma_1}{\gamma_2} = \frac{(V_1 \times t_2)}{t_1 \times V_2}$$

$K_f = \frac{2}{3} K_b T$

$$\gamma = \frac{\text{distance travelled by gas}}{\text{time taken}} = d$$

$$\frac{\gamma_1}{\gamma_2} = \frac{d_1 \times t_2}{t_1 \times d_2}$$

$$\gamma = \frac{\text{no. of moles of gas diffused / effused}}{\text{time taken}} = n$$

$$\frac{\gamma_1}{\gamma_2} = \frac{n_1 \times t_2}{t_1 \times n_2}$$

$$\frac{\gamma_1}{\gamma_2} = \frac{1}{1} \quad \text{or} \quad \frac{1}{1} = \frac{1}{1}$$

- **KTG1:**

A gas consists of large number of tiny identical particles called atoms/ molecules.

There is no force of attraction b/w the particles.

- The vol. occupied by gas molecules is negligible as compared to total volume of gas.
- The particles are always in constant and random motion.
- Particles move in all possible directions in straight lines. During their motion, they collide with other and with walls of container.
- Collisions of gas molecules are perfectly elastic.

- $K.E$  and speed always changes.

$$\approx 10^{-9} \text{ sec}$$

- Avg  $K.E \propto$  Absolute Temperature,

- $PV = \frac{1}{3} mn c^2$  (kinetic gas equation)

$$P \cdot K.E = \frac{3}{2} RT \quad (\text{Average K.E of 1 mol. of gas})$$

in first unit

$$K.E = \frac{3}{2} nRT \quad (\text{Average K.E of } n \text{ mol. of gas})$$

in second unit

$$K.E = \frac{3}{2} k_B T \quad (\text{Average K.E of one gas molecule})$$

in third unit

$k_B = \text{Boltzmann constant} = 1.38 \times 10^{-23} \text{ J/K}$

- Maxwell - Boltzmann velocity distribution curve:

- The molecules possessing very low velocities and very high fraction of velocities are found to be less.
  - VMP - velocity possessed by maximum fraction of the molecules at temperature  $T$ .
- 

$$V_{MP} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2PV}{M}} = \sqrt{\frac{2P}{\rho M}} \quad (\text{For 1 mole of ideal gas})$$

$$V_{Avg} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8PV}{\pi M}} = \sqrt{\frac{8P}{\rho \pi S}} \quad (\text{For 1 mole of ideal gas})$$

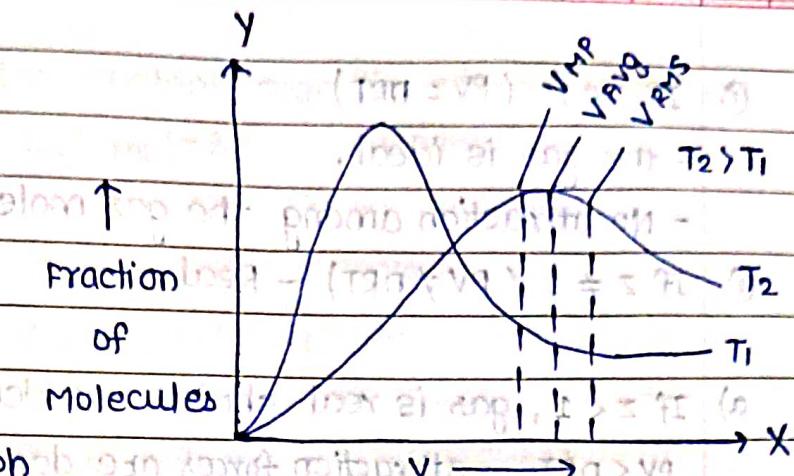
$$V_{RMS} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3P}{\rho M}} \quad (\text{For 1 mole of ideal gas})$$

- On Increasing 'T' graph shifts to right side.
- Peak becomes broader.

- Molecules possessing higher velocities increases.

- 'M' increases of velocity decreases. just molecules remains same.

- Area under the graph (Total no. of molecules)
- Density will be highest at High pressure and low temperature
- $V_{MP} < V_{AVG} < V_{RMS}$



- Deviation of a gas from ideal behaviour:

- | Ideal gas                                       | Non-ideal (Real) gas                                 |
|---|--|
| - Do obeys $KTG \propto PV$                     | - Does not obey $KTG \propto PV$                     |
| - Vol. occupied by gas molecules is negligible  | - Volume occupied by gas molecules is significant.   |
| - No attraction forces among the gas molecules. | - There is attraction or repulsion force.            |
| - obeys ideal gas law                           | - obeys van der waal's gas law                       |
| $PV = nRT$                                      | $\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT$ |
| - Very low P and very high temp.                | $f_{AD} = 9A + bV^2$                                 |

- Compressibility factor ( $z$ ):

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

- It gives the nature of gas.

① If  $z = 1$  ( $PV = nRT$ )

- the gas is ideal.

- No attraction among the gas molecules.

② If  $z \neq 1$  ( $PV \neq nRT$ ) - Real gas

a) If  $z < 1$ , gas is real shows -ve deviation from ideal behaviour.

$PV < nRT$  - attraction forces are dominating.

(attraction force dominates)

b) If  $z > 1$ , gas is real shows +ve deviation from ideal behaviour.

$PV > nRT$  - Repulsion forces are dominating.

$AV > PV > nRT$

• Vanderwaal's Gas equation :

- Derived from ideal gas equation by correcting P and V terms

$P_{\text{real}} = \frac{P}{1 - \frac{a}{V} - \frac{b}{P}}$

Vanderwaal's Gas equation,  $(P + \Delta P)(V - \Delta V) = nRT$

$\Delta P$  - pressure correction

$\Delta V$  - volume correction

① Pressure correction

Due to molecular attractions,

$$T_{\text{Real}} = (n - V) (P_{\text{real}} + \Delta P) < P_{\text{ideal}}$$

$$P_{\text{ideal}} = P_{\text{real}} + \Delta P$$

$\Delta P$  - the pressure reduced due to molecular attractions

$$\Delta P = \frac{an^2}{V^2}$$

$\Delta P \propto$  colliding molecules.

$\propto$  molecules attracting the colliding molecules

$$\Delta P \propto \frac{n^2}{V^2}$$

$$\Delta P = \frac{an^2}{V^2}$$

$a$  = vanderwaal's constant signifies attraction forces.

units of  $a$  = atm  $\text{lit}^2 \text{mol}^{-2}$

## ② Volume correction

$$V_{\text{gas}} = V_{\text{container}} - nb$$

$$TQ = (d-v)T$$

$$TQ = (v)(b+T)$$

$nb$  - vol. occupied by gas molecules is significant  
(excluded volume)

$b$  - vanderwaal's constant

signifies size of molecules.

Units :  $\text{lit mol}^{-1}$

Now, Vanderwaal's Gas equation

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

- $b$  = excluded volume per mol.

$$\text{excluded vol due to 1 molecule} = 4 \times 4 \times \pi r^3$$

$$\text{volume of each gas molecule} = \frac{4}{3} \pi r^3$$

$$\text{Volume of 1 mol of molecules} = 4 \times \frac{4}{3} \pi r^3 \times N_A = b$$

- For gases like:  $\text{CO}_2, \text{CH}_4, \text{SO}_2$  :-

- size is more

-  $a$  is more (attraction)

- For 1 mol,

~~At low P deviation is zero~~ ~~at high P = 0~~

- $V$  is very high  $\rightarrow V \approx V_{ideal}$   $\rightarrow$   $a$  is very low
- $b$  is negligible.
- $a$  is not negligible.
- $a$  is negligible.

 $V^2$ 

$$PV = \text{constant} \cdot V = \frac{V^2}{RT} \cdot PV$$

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

$$PV = 1 - a$$

$$\frac{RT}{V} = VRT$$

$$\therefore Z < 1$$

$$Z = 1 - a$$

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

$$1 - \frac{a}{V} \approx 1 + \frac{RT}{PV}$$

- Real, +ve deviation.

- Real, -ve deviation

$$T(a) = (V_b - V) / (V + V_b)$$

- For gases like  $H_2$ ,  $He$  and  $N_2$  :-

- size is very less.

-  $a$  is low.

$$RT \times P \times P = \text{constant} \rightarrow \text{at low P deviation is zero}$$

Low P

High P

-  $V$  is very high

-  $b$  is negligible

-  $a$  is negligible

$V^2$

$PV = RT$

$PV = 1$

$RT$

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

-  $a$  is negligible.

-  $b$  is not negligible.

-  $P(V-b) = RT$

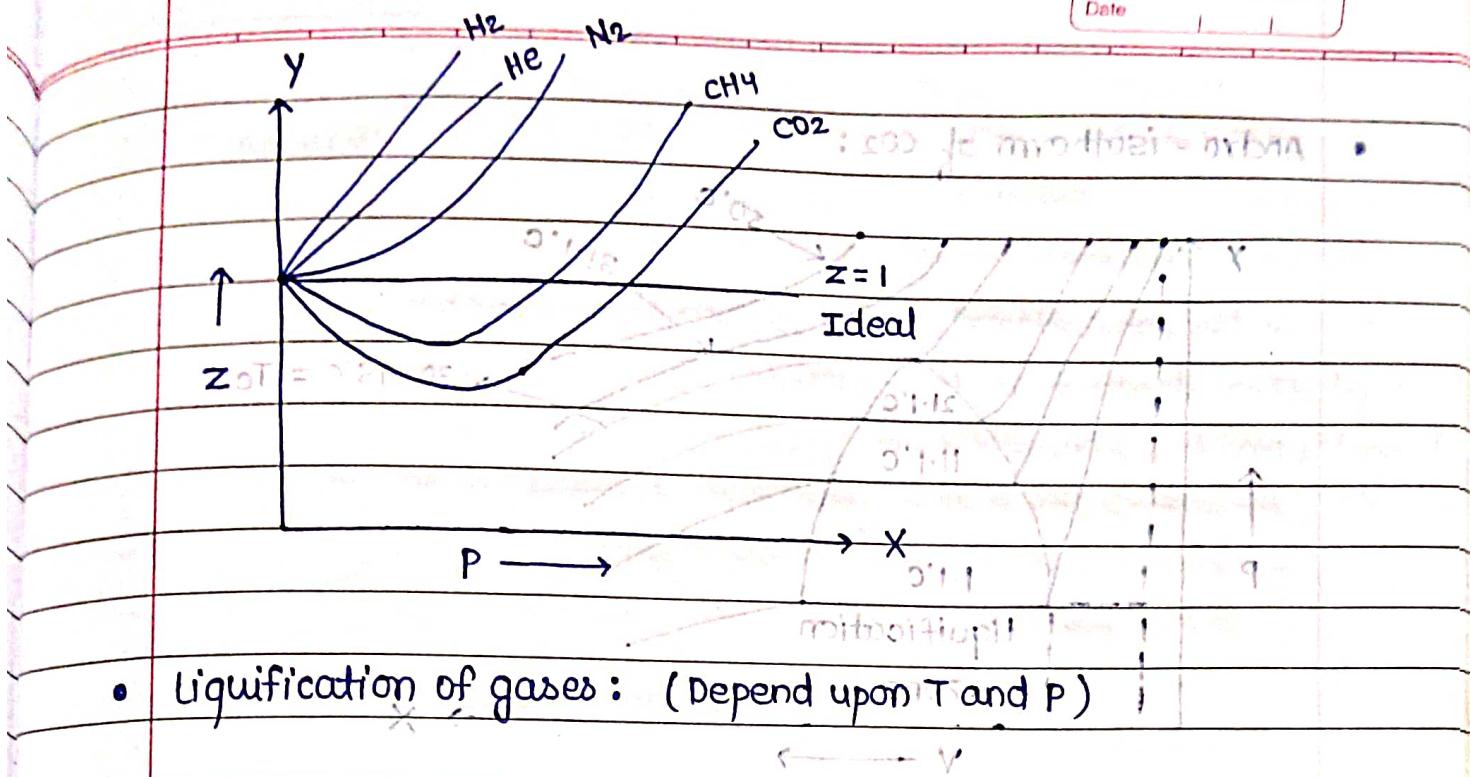
-  $PV = 1 + Pb$

$RT$

$RT$

- ( $\because Z > 1$ )

+ve deviation.



- Liquification of gases: (Depend upon  $T$  and  $P$ )

① critical temp ( $T_c$ ) - The temp above which a gas can't be liquified whatever the pressure is applied.

$$T_c = \frac{8a}{27Rb} \quad (\text{Vanderwaal's constant})$$

$$T_c \propto a$$

Liquified easily if 'a' is more.

Higher the 'a' value ( $T_c$  value), Higher the tendency to liquify

② critical pressure ( $P_c$ ):

The pressure applied at  $T_c$  to liquify the gas.

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

$$P_c = \frac{a}{27b^2} \cdot P_c = \text{Liq. point barometer}$$

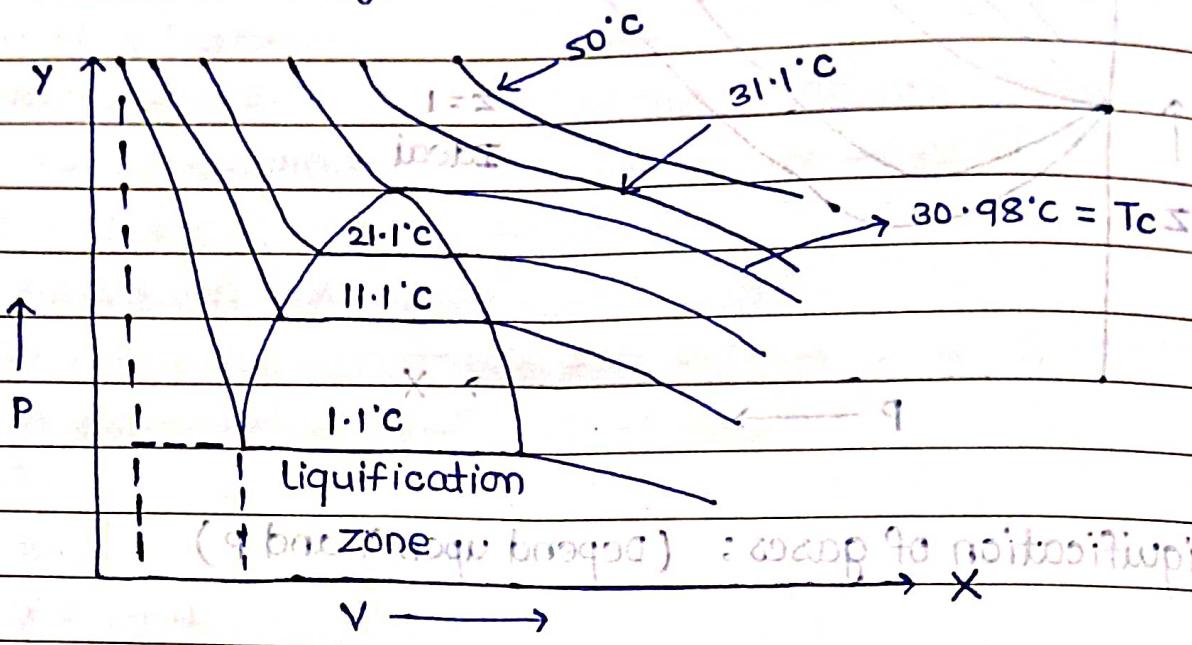
③ critical volume ( $V_c$ ):

The volume occupied at  $T_c$  and  $P_c$

$$V_c = 3b$$

$- z < 1$  at critical condition for 1 mol of gas.  $\left( z = \frac{3}{8} \right)$

- Andro - isotherm of  $\text{CO}_2$  :



- Vapour pressure:
- Intensive property (mass independent)
- depends on  $T$ ,  $a$  &  $b$

$$V \cdot P \propto T$$

- Boiling point  $V \cdot P \propto 1$

B.pt at 1 bar =  $T_b$

- Boyle's Temperature  $T_b = -\frac{a}{R}$

- It is the temperature at which real gases obey ideal gas laws over an appreciable range of pressure.

- Standard Boiling point :  $99.6^\circ\text{C}$

Temperature at which the vapour pressure of the liquid is 1 bar

- Normal boiling point :  $100^\circ\text{C}$

Temperature at which the vapour pressure of the liquid is 1 atm

$$(8 = 1) \quad \Delta H = 1 \text{ atm} \times 10^5 \text{ N/m}^2$$