CONCEPT MAP
CLASS XI

# Gaseous State

A gas occupies a volume about 1600 times greater than that of an equal weight of liquid. If gas is compressed to $1/1600^{th}$ of its volume above critical temperature (which is different for each gas) no phase change occurs and the resulting substance is a gas that is just as dense as a liquid.
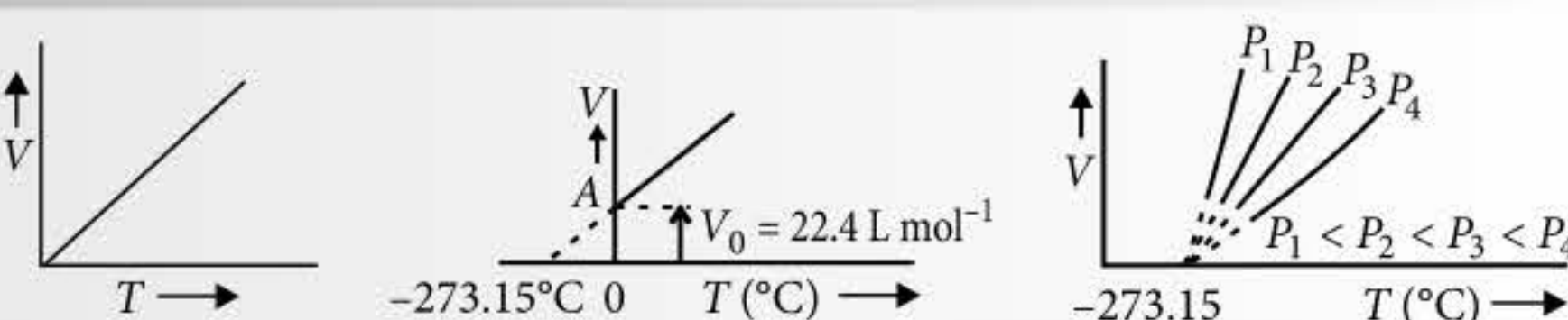
Gas Laws

- **Boyle's law :** It states that at a constant temperature (T), $P \propto (1/V)$ or $PV = \text{constant} \Rightarrow P_1 V_1 = P_2 V_2$
- **Charles' law :** It states that at constant pressure (P), $V \propto T$ or $(V/T) = \text{constant} \Rightarrow V_1/T_1 = V_2/T_2$
- **Avogadro law :** It states that at constant temperature (T) and pressure (P), $V \propto n$ or $(V/n) = \text{constant}$
- **Gay Lussac's law :** It states that at constant volume (V), $P \propto T$ or $(P/T) = \text{constant}$.

Boyle's Law
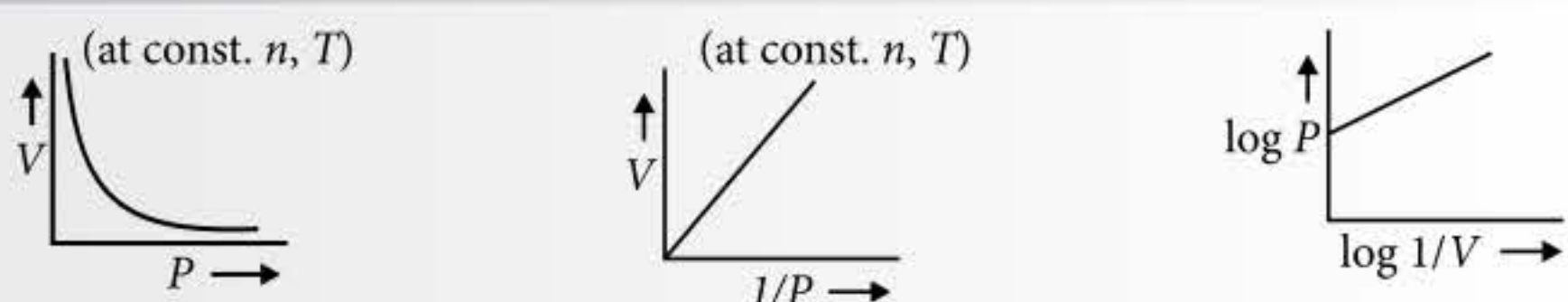
Charles' Law

Ideal Gas Equation

$PV = nRT$ [R = gas constant]

0.0821 L atm K^{-1} mol^{-1}		0.083 L bar K^{-1} mol^{-1}
8.314 J K^{-1} mol^{-1}	**R (gas constant)**	1.99 cal K^{-1} mol^{-1}

## Deviation from Ideal Gas Behaviour (Real Gas)

- Real gas follows gas laws only at high temperature and low pressure.
- It follows van der Waals' equation; $\left(P + \dfrac{an^2}{V^2} \right)(V - nb) = nRT$

Compressibility Factor

$Z = 1$ refers to ideal behaviour
$Z > 1$ shows positive deviation
$Z < 1$ shows negative deviation

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

where, Z is compressibility factor which measures the deviation from ideal behaviour.
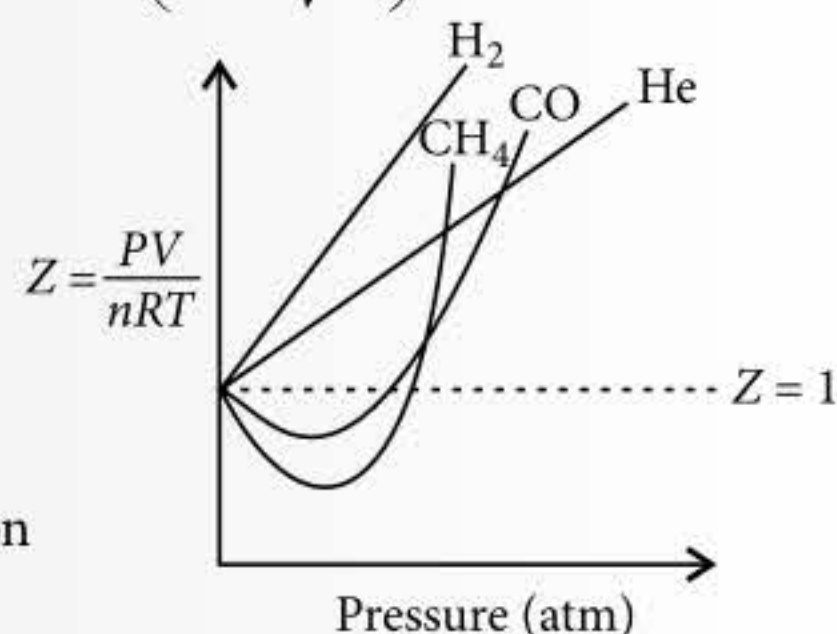
- **Relation between critical constants and van der Waals' constants :** $V_c = 3b$, $P_c = \dfrac{a}{27b^2}$, $T_c = \dfrac{8a}{27Rb}$
- As the value of a increases, the critical temperature (T_c) also increases. Higher the critical temperature of the gas, more easily it liquefies.
  The ease of liquefaction order : $Cl_2 > CO_2 > CH_4 > H_2 > He$

Graham's Law of Diffusion

$$\frac{r_1}{r_2} = \frac{v_1/t_1}{v_2/t_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

Dalton's Law of Partial Pressure

It states that "at a given temperature, the total pressure (P_{total}) exerted by two or more non-reacting gases occupying a definite volume is equal to the sum of partial pressures of the component gases".

$P_{total} = p_1 + p_2 + p_3 + ...$ (where p_1, p_2, p_3 are partial pressures of component gases)

Kinetic Molecular Theory of Gases

- All gases are made up of a very large number of minute particles like atoms or molecules.
- The molecules are in continuous random motion in all directions.
- Molecular collisions are perfectly elastic.
- The intermolecular forces of attraction are negligible.
- The average kinetic energy of a gas molecule is directly proportional to the absolute temperature.
- Kinetic gas equation : $PV = 1/3\, mnu^2$

Molecular Speed

- Root mean square velocity : $$u_{rms} = \sqrt{\frac{3RT}{M}}$$
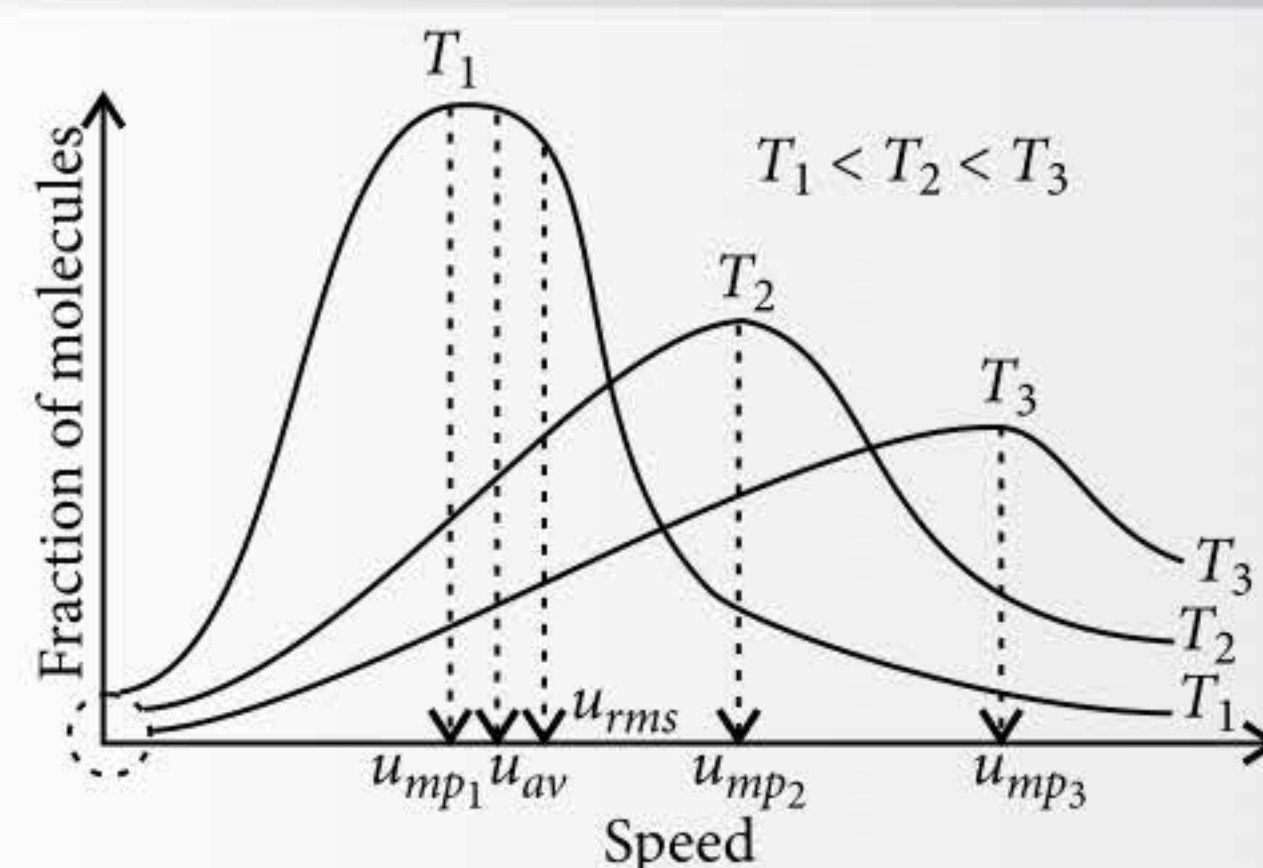- Average velocity : $$u_{av} = \sqrt{\frac{8RT}{\pi M}}$$
- Most probable speed : $$u_{mp} = \sqrt{\frac{2RT}{M}}$$

$u_{mp} : u_{av} : u_{rms} = 1 : 1.128 : 1.224$

Maxwell–Boltzmann Distribution Curve

Signifies that speed is never zero but it tends towards zero.