BRUSH UP YOUR CONCEPTS

Class XI

This specially designed column will help you to brush up your concepts by practicing questions. You can mail us your queries and doubts related to this topic at editor@mtg.in. The queries will be entertained by the author.*

STATES OF MATTER

GASEOUS STATE

Forces

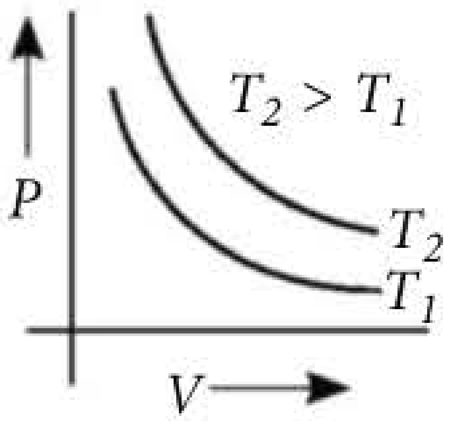
- Molecules of gases may have three types of weak van der Waals' forces :
  - Induced dipole-induced dipole interactions also called as London forces or dispersion forces, as in O_2, CH_4, CO_2, etc.
  - Dipole-dipole interactions as in HCl gas.
  - Dipole-induced dipole interactions as in Xe and H_2O during dissolution of Xe.
- In general, interactional energy $\propto \dfrac{1}{[r\,(\text{distance})]^6}$
- For stationary units having dipole-dipole interaction, the interactional energy $\propto \dfrac{1}{r^3}$

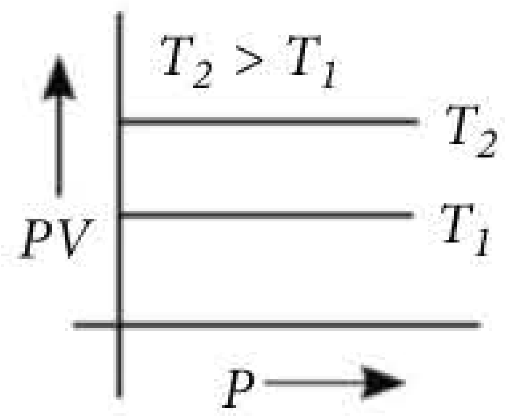
Laws

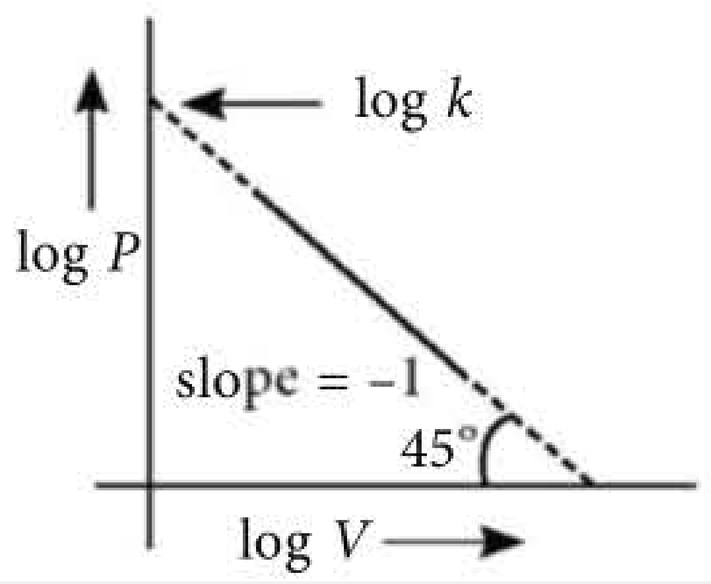
- **Boyle's law :** At constant temperature, for a given mass of a gas

$$P_1V_1 = P_2V_2 = k \quad \text{or} \quad \frac{P_1}{d_1} = \frac{P_2}{d_2} = k'$$

$$PV = k \;\Rightarrow\; P = \frac{k}{V} \;\Rightarrow\; \log P = (-1)\log V + \log k$$

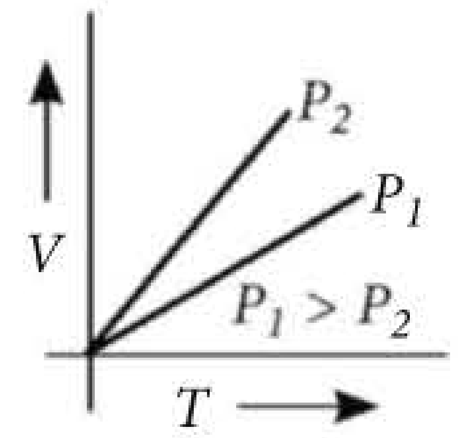


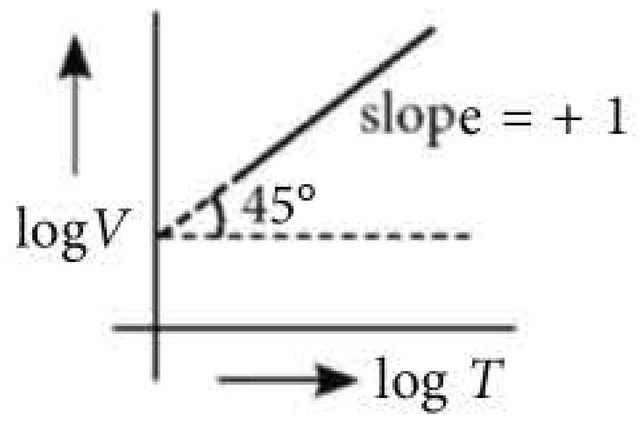




- **Charles' law :** At constant pressure, for a given mass of a gas

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} = k \quad \text{or} \quad T_1d_1 = T_2d_2 = k'$$

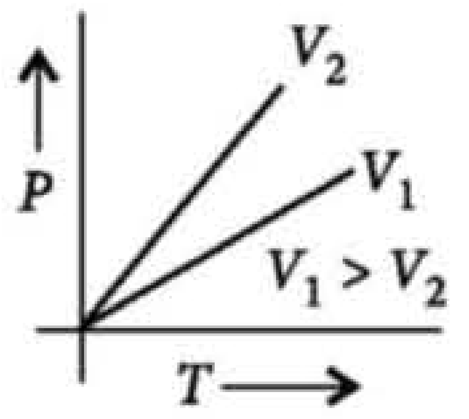


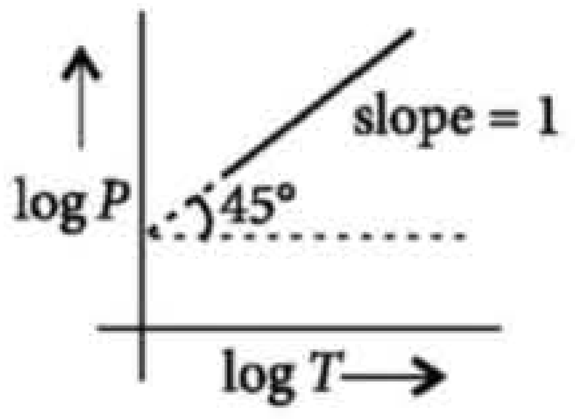


- **Gay-Lussac's or Amonton's law :** At constant volume, for a given volume of a gas,

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

$$\frac{P}{T} = k \;\Rightarrow\; \log P = \log T + \log k$$





- **Dalton's Law of Partial Pressures :** Total pressure of a mixture of non-reacting gases, in a chamber, at a specific temperature is the sum of their partial pressures.

$$P_{\text{Total}} = p_1 + p_2 + ... = \sum_{i=1}^{n} p_i$$

$$P_{\text{total}} = \frac{n_1RT}{V} + \frac{n_2RT}{V} + ... = (\Sigma \text{ moles})\,\frac{RT}{V}$$

$$x_i = \frac{n_i}{n_{\text{total}}} = \frac{p_i}{P_{\text{total}}} \;\Rightarrow\; p_i = x_iP_{\text{total}}$$

$$P_{\text{dry gas}} = P_{\text{wet gas}} - \text{Aq. tension}$$

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*By **R.C. Grover**, having 45+ years of experience in teaching chemistry.

Ideal Gas Equation

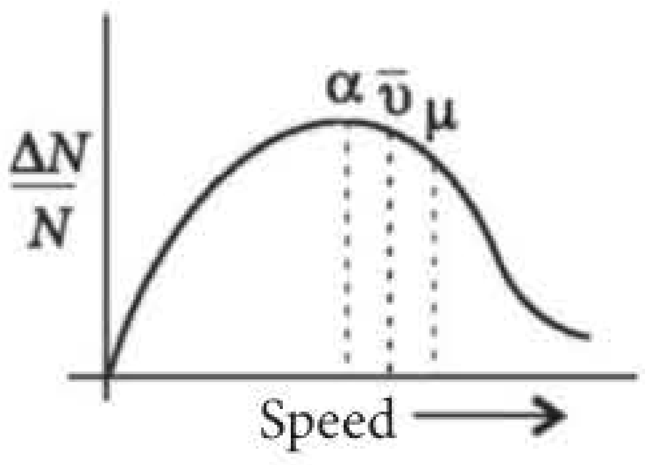
- $PV = nRT \Rightarrow R$ (molar gas constant) $= \dfrac{PV}{nT}$

  $= P - V$ work K^{-1} mol^{-1}

  $$PV = \frac{\text{mass}}{\text{molar mass }(M)} \times RT \Rightarrow PM = dRT$$

Molecular Speeds

- **Most probable speed** (α or u_{mp} or c^*) **:** It is speed of maximum fraction $\dfrac{\Delta N}{N}$ of molecules at a given temperature.



  $$\alpha = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2PV}{M}} = \sqrt{\frac{2P}{d}}$$

- **Average speed** ($\bar{\upsilon}$ or u_{av} or $\bar{c}$) **:** It is the average of speeds of different molecules of a gas at a given temperature.

  $$\bar{\upsilon} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8PV}{\pi M}} = \sqrt{\frac{8P}{\pi d}} = \frac{n_1v_1 + n_2v_2 + ...}{n_1 + n_2 + ...}$$

- **Root mean square speed** (μ or u_{rms} or c) **:** It is the square root of mean of squares of speeds of different molecules of a gas at a given temperature.

  $$\mu = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3P}{d}} = \sqrt{\frac{n_1v_1^2 + n_2v_2^2 + ...}{n_1 + n_2 + ...}}$$

- **Relations**

  $c^* : \bar{c} : c$ or $\alpha : \bar{\upsilon} : \mu = 1 : 1.128 : 1.224$

Kinetic Gas Equation

- For N_A molecules, each of mass m of a gas

  $$PV = \frac{1}{3}mN_A\,\mu^2 = \frac{1}{3}M\mu^2 \quad \text{... (i)}$$

  $$K.E. = \frac{1}{2}M\mu^2 = \frac{3}{2} \times \frac{1}{3}M\mu^2 = \frac{3}{2}PV \text{ [From eqn. (i)]}$$

  $$= \frac{3}{2}RT$$

  $$\Rightarrow K.E. \propto T; \ i.e., \ \frac{1}{2}M\mu^2 \propto T \Rightarrow \mu \propto \sqrt{T}$$

  $K.E.$ of one molecule $= \dfrac{3}{2}\dfrac{RT}{N_A} = \dfrac{3}{2}kT$

  $k = \dfrac{R}{N_A} = 1.38 \times 10^{-23}$ J K^{-1} molecule^{-1}

  [k is Boltzmann constant]

Compressibility Factor

- Ideal gas obeys all gas laws and postulates of kinetic theory of gases.

  $PV = nRT$

  Compressibility factor, $Z = \dfrac{PV}{nRT} = 1$, for ideal gas.

  For real gas, if $Z > 1$, $V > 22.4$ L at STP

  if $Z < 1$, $V < 22.4$ L at STP

  $$Z = \frac{V_{real}}{\left(\frac{nRT}{P}\right)} = \frac{V_{real}}{V_{ideal}}$$

van der Waals' Equation for Real Gas

- For 1 mole : $\left(P + \dfrac{a}{V^2}\right)(V - b) = RT$
- For n moles : $\left(P + \dfrac{an^2}{V^2}\right)(V - nb) = nRT$

  $\dfrac{an^2}{V^2}$ is pressure correction.

  Units of a = N m^4 mol^{-2} or atm L^2 mol^{-2}.

  nb is volume correction (co-volume).

  Units of b = m^3 mol^{-1} or L mol^{-1}
- Smaller the values of 'a' and 'b' closer is the gas to ideal gas behaviour.
- Higher the value of a and lower the value of b easier is the liquefaction of the gas.
- At high temperature or very low pressure, a and b both are neglected.

  $PV = RT$
- At high pressure, b cannot be neglected.

  $P(V - b) = RT \Rightarrow PV - bP = RT$

  $$\Rightarrow Z = \frac{PV}{RT} = 1 + \frac{bP}{RT}, \ i.e.\ Z > 1$$
- At moderate pressure, b can be neglected.

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

  $$\Rightarrow Z = \frac{PV}{RT} = 1 - \frac{a}{VRT}, \ i.e., Z < 1$$

Critical Constants

- **Critical temperature (T_c) :** It is the temperature above which a gas cannot be liquefied whatever the pressure be. T_c of CO_2 = 303.98 K.

  The decreasing order of critical temperatures of various gases is

  H_2O (559.73 K) > NH_3 (405.5 K) > O_2 (154.4 K) > N_2 (126 K)
- **Critical pressure (P_c) :** It is the pressure required to liquefy a gas at T_c. P_c (CO_2) = 73 atm.
- **Critical volume (V_c) :** It is the volume of 1 mole of a gas at T_c and P_c. V_c(CO_2) = 95.6 cm^3 mol^{-1}.