CONCEPT
MAP

GASEOUS STATE

Class XI

A substance exists in gaseous state if the thermal energy of molecules predominates over intermolecular forces, molecules are sufficiently apart from one another (intermolecular distances $10^{-7} - 10^{-5}$ cm), have large rotatory, vibratory and translatory motions and are most energetic.

Ideal gases

Gases which follow gas laws under all conditions of temperature and pressure.

Gas Laws

Boyle's law

$P \propto \frac{1}{V}$ (at constant T and n)

$\therefore PV = P_1V_1$ = constant $\Rightarrow \frac{P_1}{P_2} = \frac{V_2}{V_1}$

Avogadro's law

$V \propto n$ (at constant P and T)

$\rightarrow \frac{V}{n}$ = constant

Charles' law

$V \propto T$ (at constant P and n)

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

Gay-Lussac's law

$P \propto T$ (at constant V, n)

$\frac{P}{T}$ = constant $\Rightarrow \frac{P_1}{T_1} = \frac{P_2}{T_2}$

Ideal gas equation

$PV \propto nT$; $PV = nRT$

or $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ [Equation of state]

Dalton's law of partial pressures

$P_{total} = p_1 + p_2 + p_3 + p_4$ ...(at constant T, V)

where, $p_1 = x_1 P_{total}$

$p_2 = x_2 P_{total}$, ...

$P_{Dry\ gas} = P_{moist\ gas}$ – aqueous tension

Molecular velocities of gaseous molecules

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

– $U_{mp} : U_{av} : U_{rms} = 1 : 1.128 : 1.224$

Kinetic molecular theory of gases

- The actual volume of the molecules is negligible as compared to the space between them and hence they are considered as the point masses.
- Interaction between the particles is negligible.
- Particles of a gas are always in constant and random motion and the collisions between them are perfectly elastic.
- Pressure of the gas is due to the collisions of gaseous molecules with the walls of the container.

Kinetic gas equation

$PV = \frac{1}{3}mnu^2 \rightarrow u = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3RT}{M}}$; Average K.E./molecule $= \frac{3}{2}\frac{RT}{N} = \frac{3}{2}kT$

Real gases

Gases which do not follow gas laws under all conditions of temperature and pressure.

Van der Waal's equation

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

- a (atm L^2 mol^{-2}) is measure of intermolecular forces.
- b (L mol^{-1}) measure of effective size.

Compressibility factor, Z

- For ideal gas, compressibility factor i.e., $Z = PV/nRT = 1$
- For real gas, $Z = PV_{real}/nRT$, $Z \neq 1$. Hence two cases arise:
  - When $Z < 1$, it is a negative deviation i.e., easily compressible.
  - When $Z > 1$, it is a positive deviation i.e., difficult to compress.

Liquefaction of gases

It is carried out by increasing pressure or decreasing temperature.

Critical temperature (T_c)

It is the highest temperature above which a gas cannot be liquefied; $T_c = 8a/27Rb$

Critical volume (V_c)

Volume of one mole of a gas at critical temperature, $V_c = 3b$

Critical pressure (P_c)

Pressure of one mole a gas at its critical temperature: $P_c = a/27b^2$

CONCEPT
MAP
CLASS XI

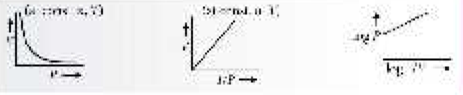
Gaseous State

A gas occupies a volume about 1600 times greater than that of an equal weight of liquid. If gas is compressed to 1/1600th 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 = constant $\Rightarrow P_1V_1 = P_2V_2$
- **Charles' law** : It states that at constant pressure (P), $V \propto T$ or (V/T) = 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) = constant
- **Gay Lussac's law** : It states that at constant volume (V), $P \propto T$ or (P/T) = constant.

Boyle's Law



Charles' Law

Ideal Gas Equation

$PV = nRT$ [R = gas constant]

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}
1.99 cal K^{-1} mol^{-1}

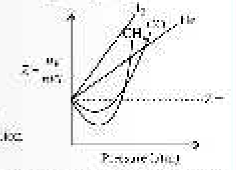
Deviation from Ideal Gas Behaviour (Real Gas)

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

Compressibility Factor

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

where Z is compressibility factor which measures the deviation from ideal behaviour.



- **Relation between critical constants and van der Waals' constants:** $V_c = 3b$, $P_c = \frac{a}{27b^2}$, $T_c = \frac{8a}{27Rb}$
- 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

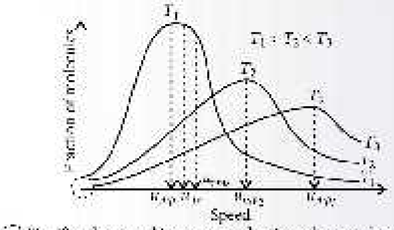
- All gases are made up of a very large number of minute particles like atoms or molecules.
- The molecules are in continuous random motion in all directions.
- Molecular collisions are perfectly elastic.
- The intermolecular forces of attraction are negligible.
- The average kinetic energy of a gas 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