KINETIC THEORY OF GASES:

1. Kinetic molecular theory of gases was proposed by Maxwell, Boltzmen, Calssius.
2. Kinetic molecular theory is applicable only to ideal gases.
3. Every gas contains a large number of tiny particles called molecules.
4. The actual volume of the molecules is negligible when compared with the volume occupied gas.
5. There are no intermolecular attractions or repulsion between the gas molecules. So ideal gases cannot be liquefied.
6. The molecules move randomly and straight in all directions with different velocities.
7. The molecules collide among themselves and also with the walls of the container.
8. The molecular collisions are perfectly elastic.
9. Molecular collisions are unaffected by gravity .
10. The pressure exerted by the gas is due to the collisions of molecules on the walls of the container. There is no loss of energy in these collisions, $\therefore$ collisions are said to be elastic.
11. The average kinetic energy of the gas molecules is proportional to the absolute temperature of the gas.

Validity of kinetic theory:

- Kinetic theory holds good at low pressure and high temperatures and fails at high pressure and low temperatures.
  i) Actual volumes of gas molecules are negligible at low P and high T but considerable at high 'P' and low 'T'
  ii) Gases are not liquefiable at very low P and high T but they can be liquefied at high 'P' low T.
  Thus kinetic theory is applicable for ideal gases and not applicable for real gases.

Kinetic gas Equation:

Based on the assumptions of kinetic theory of gases, kinetic gas equation is derived.

Kinetic gas equation PV = $\frac{1}{3}$mnc^2

m = mass

n = number of molecules

c = RMS velocity.

The RMS velocity is the root of mean of squares of individual velocities of gas molecules at a given temperature.

RMS velocity is the true average velocity because it avoids the possibility of negative or zero velocity for gas molecules.

$$C = \sqrt{\frac{c_1^2+c_2^2+c_3^2+----+c_n^2}{n}}$$

c_1, c_2, c_3 ...... c_n are the individual velocities of 'n' molecules.

Deduction of gas laws from kinetic gas equation:

1. Boyle's law: PV = $\frac{2}{3}$KT (at constant 'T')
2. Charles' law: $\frac{V}{T} = \frac{2}{3}\frac{K}{P}$ (at constant 'P')
3. Avogadro's law: $n_1 = n_2$ or
   $\frac{\frac{1}{3}m_1 n_1 c_1^2}{\frac{1}{2}m_1 c_1^2} = \frac{\frac{1}{3}m_2 n_2 c_2^2}{\frac{1}{2}m_2 c_2^2}$ (at - constant P and T)
4. Dalton's law of partial pressures:
   P = $\frac{1}{3}\frac{m_1 n_1 c_1^2}{v} + \frac{1}{3}\frac{m_2 n_2 c_2^2}{v}$ or P = $P_1 + P_2$
5. Graham's law of diffusion:
   C = $\sqrt{\frac{3p}{d}}$ or r $\alpha \frac{1}{\sqrt{d}}$ (at constant P)

Kinetic energy:

For 'n' moles of gas, kinetic energy $E_k = \frac{3}{2}$ nRT

For 1 mole of gas, kinetic energy $E_k = \frac{3}{2}$RT

For 1 molecule of gas, kinetic energy

$$E_k = \frac{3}{2}\frac{RT}{N} \text{ or } E_k = \frac{3}{2}KT$$

'K' is Boltzman constant

'K' is defined as the gas constant per molecule i.e.

$$k = \frac{R}{N}$$

k = 1.38×10^{-16} erg.$\bar{k}$. molecule□

k = 1.38×10^{-23} Joule. $\bar{k}$. molecule□

= 3.3×10^{-24} Cal. $\bar{k}$. molecule□

- Average kinetic energy of any gas is directly proportional to the absolute temperature and independent of nature of gas. This is called Maxwell's generalization.
- Two different gases at same temperature will posses same average K.E.
- If two different gases present at same temperature are mixed with each other there will not be any rise in temperature.