KINETIC THEORY OF GASES

Gaseous state is simplest of all states of matter.

Laws of gaseous behavior are more uniform & better understood.

Boyle's law, Charles' law, Graham's law, Dalton's law & Avagadro's law are well-known laws. These are empirical generalizations based on experimental observations. There is no theoretical background to justify them.

- In 19th century, Kronig, Clausius, Maxwell & Boltzmann developed Kinetic Molecular Theory of Gases, which provided sound theoretical basis for various gas laws.

Postulates of Kinetic Theory of gases:

1. A gas consists of large number of minute particles called molecules.
2. Molecules are so small that their actual volume is negligible fraction of the total volume (or space) occupied by the gas.
2. Molecules are in state of constant rapid motion in all possible directions, colliding in random manner with one another & with walls of the vessel.
3. Molecular collisions are perfectly elastic so that there is no net loss of energy when gas molecules collide with one another or against walls of vessel.
4. Kinetic energy may be transferred from one molecule to another but it is not converted into any other form of energy.

4. There are no attractive forces between molecules or between molecules & walls of vessel in which gas is contained.
Molecules move completely independent of one another.
5. Pressure is due to bombardment of molecules of molecules on walls of vessel.
& Laws of Classical Mechanics, in particular Newton's IInd law of motion, are applicable to motion of gaseous molecules.

Kinetic theory of gases is valid for ideal gas only & applicable only approximately for real gas.

Few more points regarding Kinetic Theory of Gas:

- Elastic collision means conservation of momentum i.e. any change in particle linear momentum must result in an equal & opposite change in momentum of the container wall.
- A classical description of a single molecule collision with container wall was employed & then the result was scaled up to macroscopic proportions.
- Gives explanation of ideal gas PVT behaviour.
- Useful to understand quantitative property of gases & physical property of molecules.
- Container for derivation is considered as cube for simplification as volume of the container matters, not size of the container.
- Root mean square speed is only slightly different from a simple average speed, if individual contributions are bunched together.

Root mean square speed can be quite different from a simple average speed if the individual values are spread out.

- For molecular velocities root mean square speed is typically about 10% higher than simple average speed.

9 am

Kinetic Energy & Temperature:

10 am

Average translational kinetic energy of a
11 am molecule of gas of mass m & moving
with a speed c = $\varepsilon_k = \frac{1}{2} mc^2$

Noon

Kinetic gas equation:

1 pm

$$PV = \frac{1}{3} mNc^2$$

2 pm

$$= \frac{2}{2} \times \frac{1}{3} mNc^2$$

3 pm

$$= \frac{2}{3} N \left(\frac{1}{2} mc^2\right)$$

4 pm

$$= \frac{2}{3} N \varepsilon_k$$

5 pm ε_k = kinetic Energy of 1 molecule.

$E_k = N\varepsilon_k$ = kinetic Energy of 'N' molecules

6 pm

$$\therefore PV = \frac{2}{3} E_k \quad (1)$$

7 pm

Ideal gas equation : $PV = nRT$

for $n = 1$ moles

$$PV = RT \quad (2)$$

NOTES

from (1) & (2)

$$PV = RT = \frac{2}{3} E_k \quad \text{or} \quad RT = \frac{2}{3} E_k \quad (3)$$

$$\therefore E_k = \frac{3}{2} RT \quad (\text{for } n = 1 \text{ mole})$$