KINETIC THEORY OF GASES

KTG
↳ a model of molecular behaviour which confirms the observed behaviour of Ideal gas → follows $PV = nRT$

ASSUMPTIONS FOR IDEAL GAS.

- molecules move randomly in all direct
- size of molecule << avg. distance betn molecules.
- Molecules do not exert any force on each other and on walls except collisions.
- All collisions are elastic.
- molecules obey NLM.
- density is same throughout in steady state.

• Assumptions are true for High temperature & low pressure.

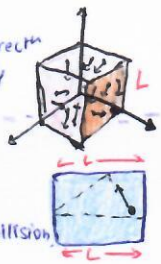
MEAN FREE PATH (λ)
↳ avg. distance traversed by a molecule with constant velocity betn two succesive collisions.

$\therefore \lambda \propto \frac{1}{\rho}$; $\lambda \propto \frac{1}{d^2}$ (ρ: density, d: diameter)

$$\lambda = \frac{1}{\sqrt{2}\pi d^2 (N/V)}$$

N = no. of molecules, V = volume.

PRESSURE OF AN IDEAL GAS
↳ exerted due to the collision of molecules on walls of container.



→ If a molecules collides in x directn there's no effect on velocity of y and z directn.

$P_i = mv_x \quad P_f = -mv_x$

$\Delta P = -mv_x - mv_x = -2mv_x$

time taken betn consecutive collision $\Delta t = \frac{2L}{v_x}$

$\text{Avg. } F = \frac{\Delta P}{\Delta t}$

$\Delta F = \frac{2mv_x}{2L/v_x} = \frac{mv_x^2}{L}$

$v_x = v_y = v_z = \frac{v}{3}$

$F_{avg} = \frac{m(v_{x_1}^2 + v_{2x}^2 + v_{3x}^2 + \dots)}{L}$

$\therefore P = \frac{F_{avg}}{A} = \frac{m(v_{1x}^2 + v_{2x}^2 + v_{3x}^2 + \dots)}{L \times L^2}$

$P = \frac{1}{3}\frac{m}{L^3}\Sigma v^2 = \frac{1}{3}\frac{mN}{L^3}\frac{\Sigma v^2}{N} \quad (\bar{v}^2, \rho)$

$P = \frac{1}{3}\rho \bar{v}^2$ → meansquare speed.

ROOT MEAN SQUARE SPEED

Square : $v_1^2 + v_2^2 + v_3^2 + v_4^2 + \dots v_N^2$

mean : $\frac{v_1^2 + v_2^2 + v_3^2 + \dots + v_N^2}{N}$

Root : $\sqrt{\frac{v_1^2 + v_2^2 + v_3^2 + \dots v_n^2}{N}}$

$\therefore v_{rms} = \sqrt{\frac{\Sigma v_i^2}{N}}$ $\quad v_{rms} = v_{rms}^2 = \frac{\Sigma v_i^2}{N}$

$v_{rms} = \sqrt{\frac{3P}{\rho}} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3nRT}{M}} = \sqrt{\frac{3RT}{M_0}}$ (M_0: molar mass)

$v_{rms} = \sqrt{\frac{3RT}{M_0}}$

KINETIC INTERPRETATN OF TEMP.
$P = \frac{1}{3}\rho v_{rms}^2 = \frac{1}{3}\frac{M}{V}v_{rms}^2 \rightarrow PV = \frac{2}{3}\left[\frac{1}{2}Mv_{rms}^2\right]$

$\frac{1}{2}Mv_{rms}^2$ is avg. translatn K.E of one mole (only energy in molec)

$PV = \frac{2E}{3} \rightarrow nRT = \frac{2E}{3} = \frac{nN_AR}{N_A}T$

$\frac{R}{N_A} = k_B$ (Boltzmann constant)

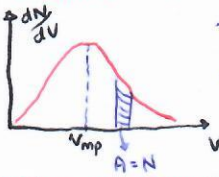
$\therefore Nk_BT = \frac{2E}{3}$

$E = \frac{3}{2}k_BT \times N$ (N: no. of molecules, T: temp)

Energy per molecule $= \frac{E}{N} = \frac{3}{2}k_BT$

$\therefore K.E \propto T$ and does not depend on P, V or nature of gas.

DISTRIBUTION OF MOLECULAR SPEEDS



→ At any temp., molecular speed varies from 0 to ∞

→ Number of molecules is area under curve.

→ The speed at which dN/dV is maximum is called most probable speed.

→ given by Maxwell

TYPES OF MOLECULAR SPEEDS
Root Mean Square Speed.

$v_{rms} = \sqrt{\frac{v_1^2 + v_2^2 + \dots + v_n^2}{n}}$ $\quad v_{rms} = \sqrt{\frac{3RT}{M_0}} = \sqrt{\frac{3kT}{m}}$

Average speed.

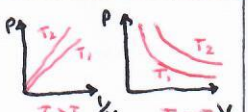
$v_{avg} = \frac{v_1 + v_2 + \dots + v_n}{n}$ $\quad v_{avg} = \sqrt{\frac{8}{\pi}\frac{RT}{M_0}} = 0.92\, v_{rms}$

Most Probable Speed.

$v_{mp} = \sqrt{\frac{2RT}{M_0}} = 0.816\, v_{rms}$

$v_{rms} > v_{avg} > v_{mp}$

BOYLE'S LAW
$P \propto \frac{1}{V}$ (temp. is constant) $\quad PV = \text{constant}$



CHARLE'S LAW
$V \propto T$ (Pressure constant.) $\quad \frac{V}{T} = \text{constant}$

GAY-LUSAAC'S LAW/PRESSURE LAW
$P \propto T$ (volume constant) $\quad \frac{P}{T} = \text{constant}$

IDEAL GAS LAW
$PV = nRT$

P = pressure, V = volume, n = no. of moles, R = universal gas constant, T = temperature (K)

$R = 8.314$ J/mol K

GRAHAM'S LAW OF DIFFUSION
$P_1 = \frac{1}{3}\rho_1 v_1^2 \quad P_2 = \frac{1}{2}\rho_2 v_2^2$

When reached in steady state. $P_1 = P_2$

$\frac{v_1}{v_2} = \sqrt{\frac{\rho_2}{\rho_1}} = \frac{r_1}{r_2}$ (r: rate of diffusn)

DALTON'S LAW OF PARTIAL PRESSURE
Total pressure exerted by no. of non-reactn gases when taken together in a container is equal to sum of partial pressu

$P_1 + P_2 + P_3 + \dots = P$

DEGREE OF FREEDOM
→ it is the total no. of possible motions a gas molecule can have.

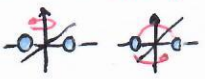
→ includes translational, rotational, vibrational

Translational dof

$\frac{1}{2}mv_x^2, \frac{1}{2}mv_y^2, \frac{1}{2}mv_z^2$

$f = 3$

Rotational dof



$\frac{1}{2}I_x\omega_x^2, \frac{1}{2}I_y\omega_y^2, \frac{1}{2}I_z\omega_z^2$

$f = 3$

Vibrational dof.

→ act at very high temp.

→ just like a spring.

MONOATOMIC GAS.


→ negligible in size (just like ideal gas)

→ no rotatn dof.

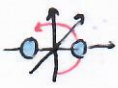
e.g Ne, Ar, He, etc.

$f = 3 + 0 = 3$

DIATOMIC GAS
→ dumbell shaped. e.g H_2, N_2, O_2, etc.

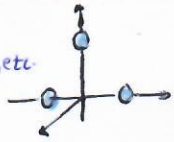
→ At room temp. $f = 3 + 2 = 5$ (Trans, Rot)

→ At high temp. $f = 3 + 2 + 2 = 7$ (Trans, Rot, Vib)



TRIATOMIC OR POLYATOMIC GAS
→ non linear molecule.

$f = 3 + 3 = 6$ (Trans, Rot) e.g H_2O, SO_2, etc.



LAW OF EQUIPARTITION OF ENERGY.
"The total energy of ideal gas molecule is distributed equally among of all its dof"

Energy associated with each dof $= \frac{1}{2}kT$

For 1 molecule, $E = \frac{1}{2}kT$ $\quad$ For 1 mole, $E = \frac{1}{2}RT$

INTERNAL ENERGY OF A GAS.
From law of Equipartitn of energy.

$E = f/_2 kT$ → For one molecule.

$E = f/_2 kT \times n \times N_A \Rightarrow U = f/_2 nRT$ (for gas.)

→ Internal energy of gas depends only on absolute temperature.

For Monoatomic Gas :→ $\frac{3}{2}nRT$

For Diatomic Gas (Room Temp) :- $\frac{5}{2}nRT$

For Diatomic Gas (High Temp.) : $7/_2 nRT$

EQUIVALENT DEGREE OF GASEOUS MIXTURE.
$f_{eq} = \frac{f_1 n_1 + f_2 n_2 + f_3 n_3 + \dots + f_n n_n}{n_1 + n_2 + n_3 + \dots n_N}$