Theory of gases

Kinetic Theory of Gases

☆ BEHAVIOUR OF GASES

Date

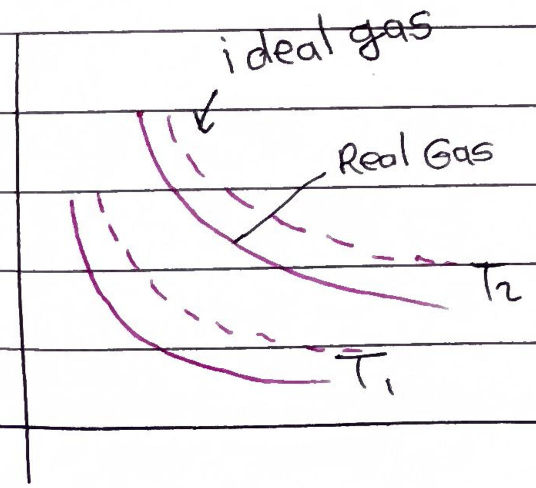
* Boyle's law

For a fixed mass of perfect (ideal) gas, Pressure of a gas is inversely proportion to the volume of gas provided the temperature remains constant (isothermal)

$$P \propto \frac{1}{V} = \frac{K}{V} \Rightarrow PV = K = \text{constant}$$

$$\boxed{P_1V_1 = P_2V_2}$$

Given by Robert boyle in 1662



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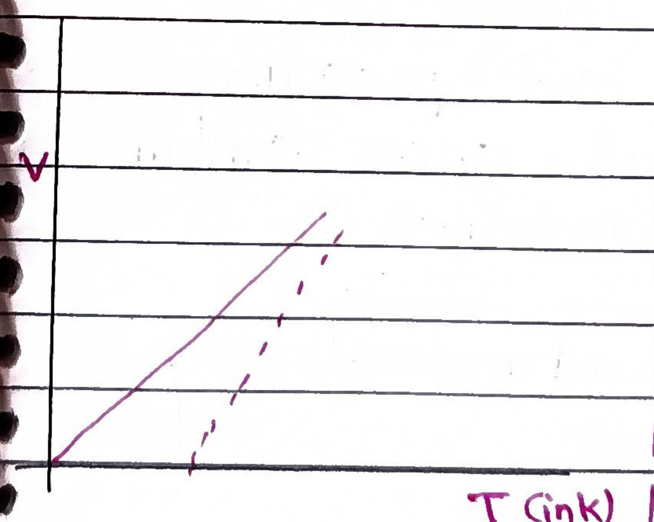
$(T_2 > T_1)$

$$\text{Slope} = \frac{dP}{dV} = \frac{-P}{V}$$

* (Real Gasbehave as ideal gas at high temp and Low pressure)

Charle's Law :

- Given by Alexander Charles in 1787
- For a fixed mass of ideal gas, Volume of gas is directly proportio to the absolute temp. of gas, provided the pressure remains Constant (isobaric)



$$V \propto T \rightarrow V = RT \Rightarrow$$

$$K = \frac{V}{T} = \boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

$$\text{Slope} = \frac{dV}{dT} = \frac{nR}{P}$$

$$\left(\text{Slope} \propto \frac{1}{P}\right)$$

Date

Behaviour of Gases

→ If Temp. is in degree Celcius

for each degree rise in temp in Celcius, the volume of given mass of gas increases by $\frac{1}{273}$ of its Volume at 0°C

$$V_{t°C} = V_0 + V_0 \times \frac{t}{273}$$ $\boxed{V_0 = V_{0°C}}$

Avagadro's Hypothesis ÷

At constant temp. and pressure equal Volume of all gases contain equal number of molecules

$\boxed{V \propto \text{no. of molecules}}$

Date

Ideal Gas equation - Equation of State

$$\boxed{PV = nRT}$$

- 1atm = 1bar ≈ 10^5 Pa ≈ $10^5 \frac{N}{m^2}$ ≈ 760 mmHg ≈ 760 torr
- 1L ≈ $10^{-3}m^2 = 10^3 cm^3 = 1 dm^2$

$$PV = K.NT$$

↓

Boltzman's constant $(K = 1.38 \times 10^{-23}$ J/K

N = no. of molecules

Date

Density of a gas $(d = \frac{m}{v})$

Ideal gas eq.n $PV = nRT$

$$PV = \frac{m}{M} RT$$

$$\frac{m}{V} = \frac{PM}{RT}$$

$$\boxed{d = \frac{PM}{RT}} \quad \ldots (M = m \times N_A)$$

$$d = \frac{Pm \times Na}{RT} = \frac{Pm}{\left(\frac{R}{Na}\right) T} = \boxed{d = \frac{Pm}{KT}}$$

$$\left(\frac{R}{Na} = K\right)$$

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Ideal gas =

The gas which obeys the ideal gas equation ($PV = nRT$) at all temp and pressure.

Assumption -

(1) The size of the gas molecules is negligibly small.

2) There is no force of attraction amongst the molecules of the gas.

3) Molecules are in random motion
- Molecules are always moving in random directions with different speeds.
- They continuously collide with each other and with the container wall.

Date

Kinetic Theory of Ideal gas

i) Gas is made up of atoms and molecules.

ii) Molecules of same gases are idential in all respect.

iii) Molecules are constanlly in random motion along straight line (only translation motion) (DOF $(f)=3$)

iv) All the collisions of gas molecules among themselves and wall of container is elastic in nature

$$\left[\frac{\vec{P_i}}{KE_i} = \frac{\vec{P_f}}{E_f f}\right]$$

v) No intermolecular force

vi) Elastic collisions only

Date

vii) Gravity in neglected

viii) Molecules of gas are tiny rigid shpores of negligible volume

Date

Pressure Derivation from Molecular Motion

Let's derive the expression for Pressure using kinetic theory.

Consider:

- A gas in a cubic container of side L
- Let a single molecule of mass m move with velocity components (v_x, v_y, v_z)

Change in momentum after one wall collision

$$\Delta P = -2mv_x \quad \text{(elastic collision)}$$

time between two such collisions with same wall

$$\Delta t = \frac{2L}{v_x}$$

Force on the wall

$$F = \frac{\Delta P}{\Delta t} = \frac{2mv_x}{2L} = \frac{mv_x^2}{L}$$

For N molecules, average over all

$$P = \frac{F}{A} = \frac{1}{3}\frac{Nm\overline{v^2}}{V}$$

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Pressure of an Ideal gas.

Final formula:

$$P = \frac{1}{3}\frac{nm\overline{v^2}}{V}$$

V = tov Volume of gases

P = Pressure

m = mass of 1 molecule

n = no. of molecule

$\overline{v}$ = Root mean square velocity

This can be also written as.

$$\boxed{P = \frac{1}{3}d\overline{v^2}}$$

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Root mean Square Velocity ($\bar{U}$)

$$\bar{U} = \sqrt{\frac{v_1^2 + v_2^2 + \dots + v_n^2}{n}}$$

RMS Velocity for gas molecules

$$\bar{U} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3KT}{m}}$$

$R = 8.314$ J/mol·K

$T =$ in kelvin

$M =$ Molar Mass in kg

$m =$ mass of 1 molecules in kg

In terms of P and density

$$d = \frac{PM}{RT} \quad \frac{RT}{M} = \frac{P}{d} \quad \bar{U} = \sqrt{\frac{3RT}{M}} = \boxed{\bar{U} = \sqrt{\frac{3P}{d}}}$$

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Molecular Velocities =

1. Average velocity

$$V_{avg} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{M}} = \sqrt{\frac{8P}{\pi d}}$$

$$= \frac{V_1 + V_2 + \dots + V_n}{n}$$

2. Root mean square velocity

$$\bar{V} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3KT}{m}} = \sqrt{\frac{3P}{d}}$$

3. Most Probable velocity

$$V = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2KT}{m}} = \sqrt{\frac{2P}{d}}$$

Date

Mean Free Path

The average distance travelled by a molecule betn two succesive collisions is known as mean free path

$$\lambda m = \frac{\lambda_1 + \lambda_2 + \dots + \lambda n}{n}$$

$$\boxed{\lambda m = \frac{1}{\sqrt{2} n \pi d^2}}$$

λm = mean free path
n = no. of molecules per unit Volume
d = diameter of molecule.

Date

Specific heat Capacity

1) Monatomic gases

Monoatomic gases are gases whose molecules consist of only one atom. These gases exhibit only translational motion and hence have 3 degree of fredom

ex. Helium(He), Neon(Ne), argon.

* Degrees of Freedom (f)

$f = 3$ (only translational - x, y, z direction)

* Kinetic Energy of one Molecule:

$$\varepsilon = \frac{3}{2} K_B T$$

- K_B = Boltzmann constant
- T = absolute tem.

Internal Energy (U):

For one mole of gas / For N molecule

$$U = \frac{3}{2} RT \qquad U = \frac{3}{2} N K_B T$$

Date

2) Diatomic Gases

3) Polyatomic gases

Date

Specific Heat of Solids

Near the room temp. the molar specific heat of most of the solids at constant volume is equal to 3R or 6 cal/mol K or 25 J/mol.K

This statement is known as Dulong and Petit's law

Debye Temp –

T at which the molar specific heat of a solid at constant volume becomes equal to 3R

The specific heat of water is nearly 75 J/mol K ≈ 9R

Date

2. Dulong and Petit's law - Based on K.T

It is the classical prediction ~~of~~ for Specific heat of solids:

$$\boxed{C = 3R}$$

- Valid