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Introduction

Overview

- Gases consist of rapidly moving atoms or molecules.
- Kinetic theory explains gas behavior by neglecting inter-atomic forces.
- Developed in 19th century by Maxwell, Boltzmann, and others.
- Explains molecular basis of pressure and temperature.
- Consistent with gas laws and Avogadro's hypothesis.
- Explains specific heat capacities of gases.

Solved Examples

- **Ex 1:** Calculate pressure from molecular collisions.
 - Use formula: $\text{Pressure} = \text{Force}/\text{Area} = (\text{Change in momentum} \times \text{collisions per second})/\text{Area}$.
 - Apply kinetic theory assumptions for ideal gas.

Practice Set

- Q1: Kinetic theory assumes gas molecules are:
 - A) Stationary
 - B) Rapidly moving
 - C) Strongly interacting
 - D) At rest
- Q2: Kinetic theory was developed by:
 - A) Newton and Galileo
 - B) Maxwell and Boltzmann
 - C) Dalton and Avogadro
 - D) Einstein and Planck
- Q3: Kinetic theory explains:
 - A) Molecular basis of pressure and temperature
 - B) Only chemical reactions
 - C) Solid state properties
 - D) Nuclear forces

Answer Key

- Q1: B - Gas molecules are rapidly moving particles.
- Q2: B - Maxwell and Boltzmann developed kinetic theory.
- Q3: A - It explains molecular interpretation of pressure and temperature.

Memory Aids

- Gas molecules = Rapid motion + negligible forces.
- Maxwell & Boltzmann = Kinetic theory pioneers.
- Pressure & temperature = Molecular collisions.

Molecular Nature of Matter

Key Concepts

- **Feynman's atomic hypothesis:** Atoms are tiny particles in constant motion, attracting at a distance, repelling when close.
- **Dalton's Atomic Theory:** Atoms combine in fixed mass ratios obeying simple laws.
- **First law:** Fixed mass proportion in compounds.
- **Second law:** Masses of elements in multiple compounds are in small integer ratios.
- **Gay Lussac's law:** Volumes of gases combine in small integer ratios.
- **Avogadro's hypothesis:** Equal volumes of gases at same T and P contain equal molecules.
- **Mean free path:** Average distance a molecule travels without collision.

Solved Examples

- **Ex 1:** Calculate mean free path.
 - Use formula: $\lambda = \frac{1}{\sqrt{2}\pi d^2 \frac{N}{V}}$
 - Where d = molecular diameter, N/V = number density.

Practice Set

- Q1: According to Feynman, atoms:
 - A) Are static
 - B) Move perpetually
 - C) Do not attract
 - D) Are large particles
- Q2: Dalton's first law states:
 - A) Variable mass ratios in compounds
 - B) Fixed mass proportion in compounds
 - C) Gases combine by volume
 - D) Molecules are indivisible
- Q3: Mean free path is:
 - A) Time between collisions
 - B) Average distance between collisions
 - C) Molecular speed
 - D) Volume of gas

Answer Key

- Q1: B - Atoms move in perpetual motion.
- Q2: B - Fixed mass proportion in compounds.
- Q3: B - Mean free path is average distance traveled without collision.

Memory Aids

- Feynman: Atoms always moving, attract & repel.
- Dalton: Fixed mass ratios in compounds.
- Mean free path = distance between collisions.

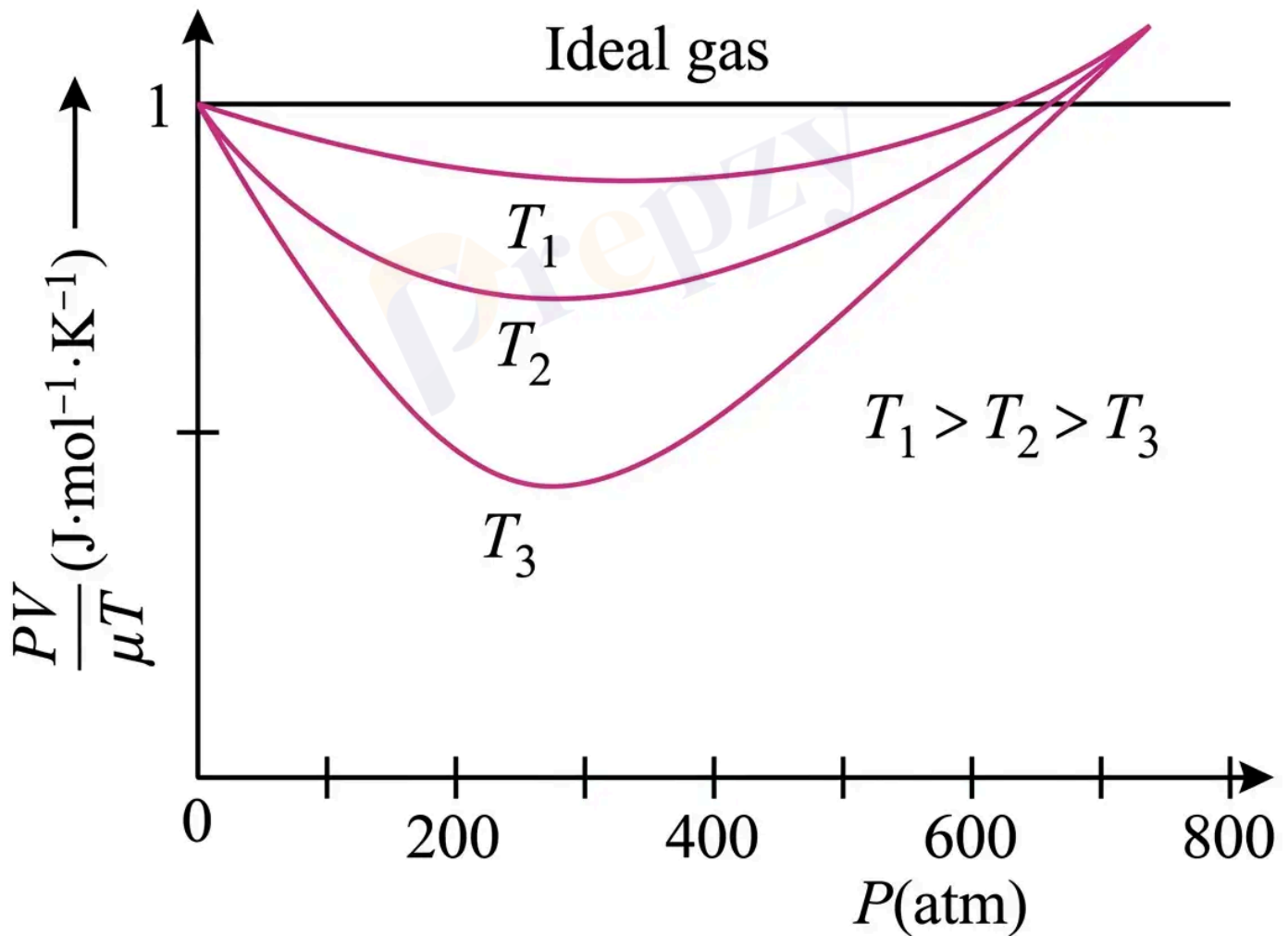
Behaviour of Gases

Gas Behavior at Low Temperature and High Pressure

- For gases at low temperature and high pressure, $PV = KT$
- Constant $K = Nk_B$, where k_B is Boltzmann constant (1.38×10^{-23} J/K)
- N is the number of molecules in the gas
- At same P, V, T , number of molecules N is same for all gases
- Number of molecules in 22.4 litres of any gas = Avogadro number $N_A = 6.02 \times 10^{23}$

Perfect Gas Equation

- Equation: $PV = \mu RT$ where μ = number of moles, $R = N_A k_B = 8.314$ J/mol·K
- Relation: $\mu = \frac{M}{M_0} = \frac{N}{N_A}$, where M = mass, M_0 = molar mass
- Alternate form: $PV = \mu RT = k_B NT$ or $P = k_B nT$, where n = number density
- Real gases approach ideal gas behavior at low pressure and high temperature
- Pressure relation: $P = \frac{M}{M_0} \frac{RT}{V} = \frac{\rho RT}{M_0}$, where ρ = mass density
- Ideal gas perfectly satisfies $PV = \mu RT$ at all pressures and temperatures (theoretical model)



Solved Examples

- Ex 1:** Calculate pressure P if $N = 2 \times 10^{23}$, $T = 300\text{K}$, and number density $n = 5 \times 10^{25} \text{ m}^{-3}$ using $P = k_B nT$.
 - $P = 1.38 \times 10^{-23} \times 5 \times 10^{25} \times 300 = 2070 \text{ Pa}$

- **Ex 2:** Find volume V of 1 mole of gas at $P = 1 \text{ atm}$, $T = 273\text{K}$ using $PV = \mu RT$.
 - $V = \frac{\mu RT}{P} = \frac{1 \times 8.314 \times 273}{1.013 \times 10^5} = 0.0224 \text{ m}^3$ or 22.4 litres

Practice Set

- Q1: What is the value of Boltzmann constant k_B ?
 - A. $1.38 \times 10^{-23} \text{ J/K}$
 - B. $6.02 \times 10^{23} \text{ mol}^{-1}$
 - C. $8.314 \text{ J/mol}\cdot\text{K}$
 - D. 22.4 litres
- Q2: For an ideal gas, which equation is correct?
 - A. $PV = nRT$
 - B. $PV = \mu RT$
 - C. $P = k_B nT$
 - D. All of the above
- Q3: Number of molecules in 22.4 litres of any gas at STP is?
 - A. 6.02×10^{23}
 - B. 1.38×10^{-23}
 - C. 8.314
 - D. 22.4

Answer Key

- Q1: A - Boltzmann constant $k_B = 1.38 \times 10^{-23} \text{ J/K}$ is universal for gases.
- Q2: D - All equations represent ideal gas relations in different forms.
- Q3: A - Avogadro number 6.02×10^{23} molecules in 22.4 litres at STP.

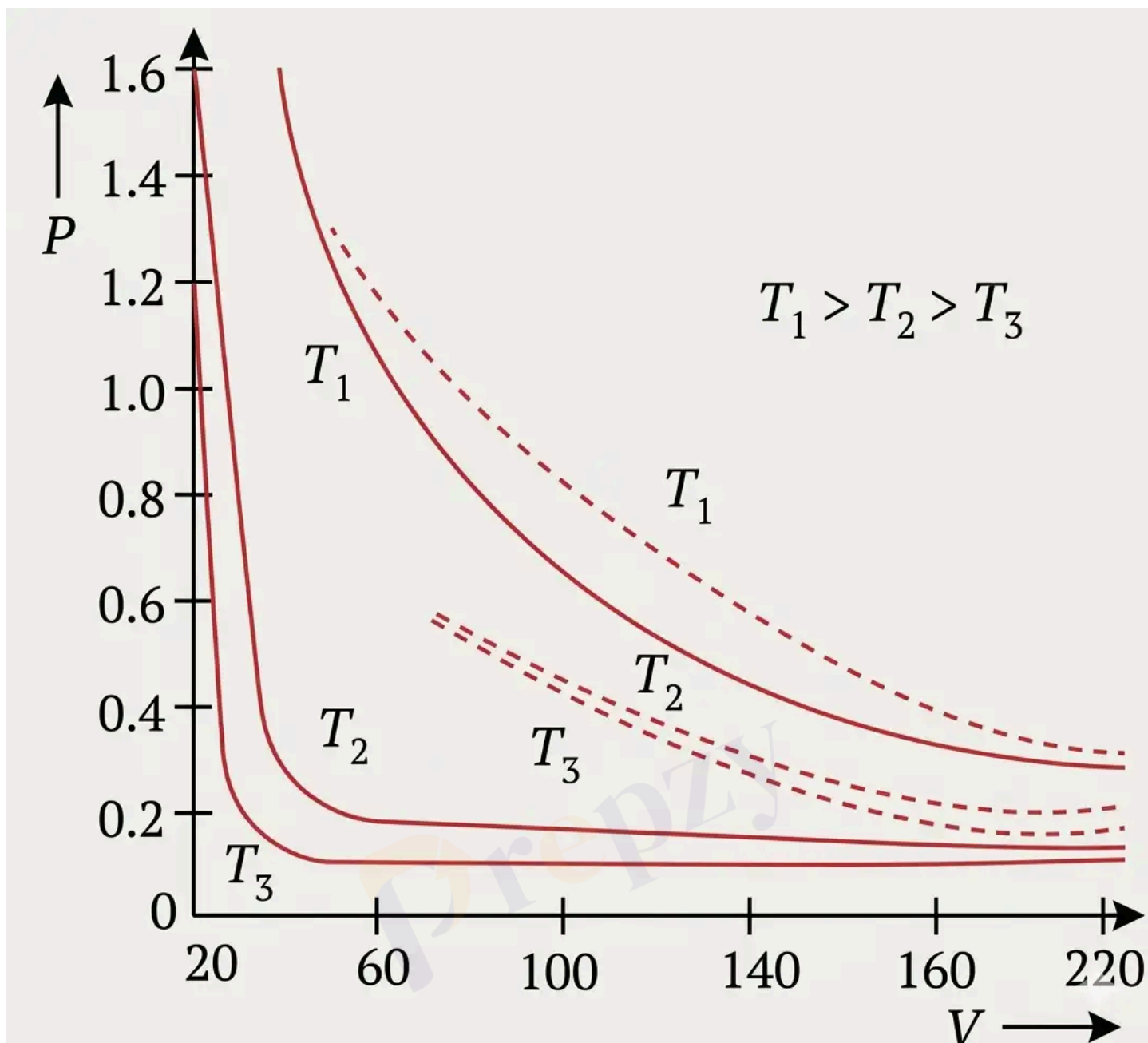
Memory Aids

- $PV = \mu RT$ is the perfect gas equation.
- Boltzmann constant $k_B = 1.38 \times 10^{-23} \text{ J/K}$.
- Avogadro number $N_A = 6.02 \times 10^{23}$.
- Real gases behave ideally at low pressure, high temperature.

Boyle's Law

Pressure-Volume Relationship

- At constant temperature (μ, T), $PV = \text{constant}$
- Pressure varies inversely with volume: $P \propto \frac{1}{V}$
- Experimental P-V curves for steam match Boyle's law at high temperature and low pressure



Solved Examples

- **Ex 1:** If volume halves at constant temperature, what happens to pressure?
 - Pressure doubles ($P \propto \frac{1}{V}$)
- **Ex 2:** Calculate new pressure if initial $P_1 = 2$ atm, $V_1 = 4$ L, and volume changes to $V_2 = 2$ L.
 - $P_2 = \frac{P_1 V_1}{V_2} = \frac{2 \times 4}{2} = 4$ atm

Practice Set

- Q1: Boyle's law states that at constant temperature, pressure is?
 - A. Directly proportional to volume
 - B. Inversely proportional to volume
 - C. Independent of volume
 - D. Proportional to temperature
- Q2: If volume is doubled at constant temperature, pressure will?
 - A. Double
 - B. Halve
 - C. Remain same

- D. Become zero
- Q3: For a gas at constant temperature, if pressure is 3 atm and volume is 2 L, what is volume when pressure is 6 atm?
 - A. 1 L
 - B. 4 L
 - C. 3 L
 - D. 6 L

Answer Key

- Q1: B - Pressure varies inversely with volume at constant temperature.
- Q2: B - Doubling volume halves the pressure.
- Q3: A - Using $P_1V_1 = P_2V_2$, $V_2 = 1$ L.

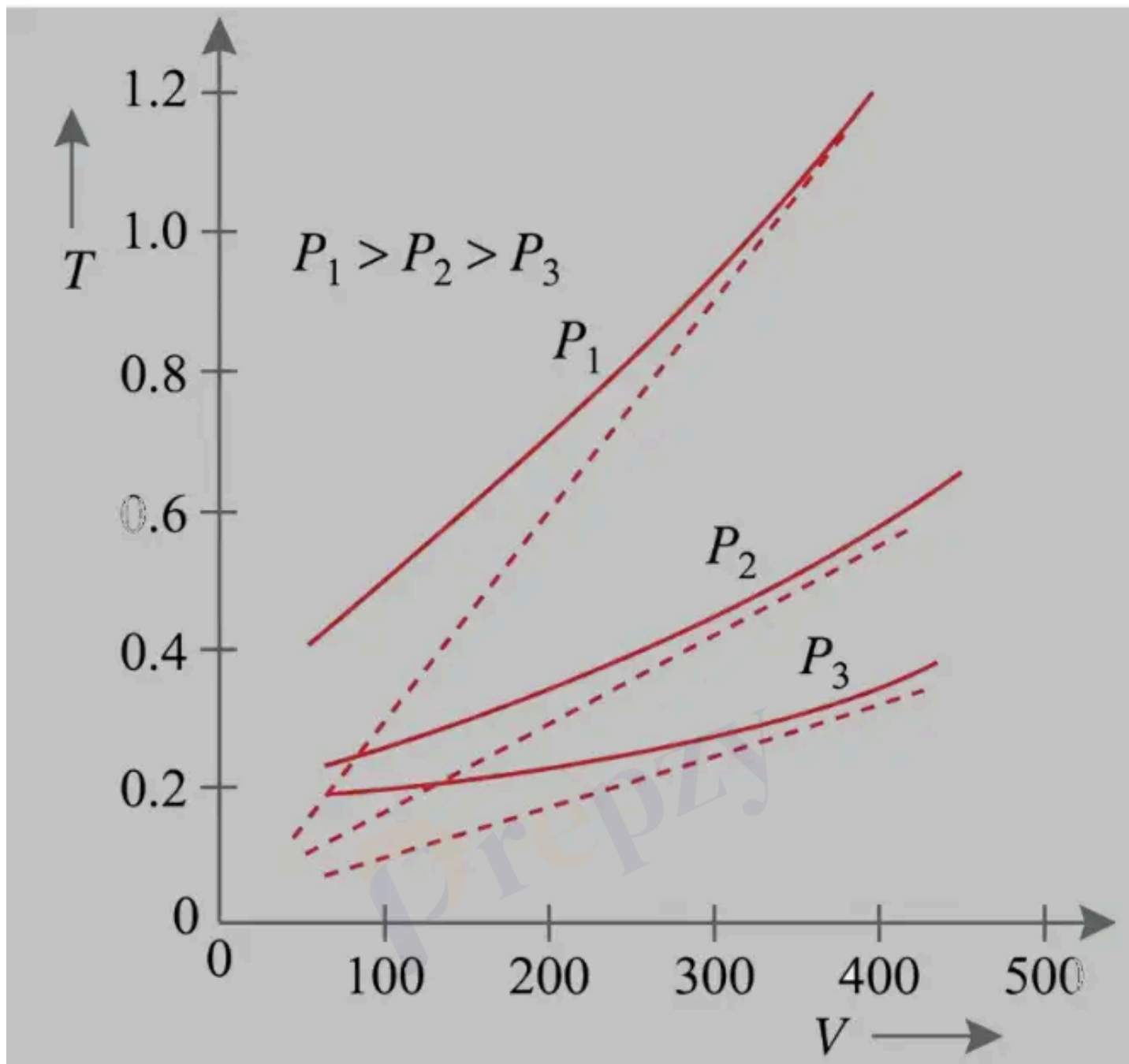
Memory Aids

- Boyle's law: $PV = \text{constant}$
- Pressure and volume are inversely related at constant temperature

Charles' Law

Volume-Temperature Relationship

- At constant pressure, volume V is proportional to absolute temperature T : $V \propto T$
- Experimental T-V curves for CO₂ match Charles' law at high temperature and low pressure



Solved Examples

- **Ex 1:** If volume is 2 L at 300 K, find volume at 600 K at constant pressure.
 - $V_2 = V_1 \times \frac{T_2}{T_1} = 2 \times \frac{600}{300} = 4 \text{ L}$
- **Ex 2:** Calculate temperature if volume changes from 1 L to 3 L at constant pressure and initial temperature 273 K.
 - $T_2 = T_1 \times \frac{V_2}{V_1} = 273 \times \frac{3}{1} = 819 \text{ K}$

Practice Set

- Q1: Charles' law states volume is proportional to?
 - A. Pressure
 - B. Temperature (absolute)
 - C. Number of moles
 - D. Mass
- Q2: If temperature doubles, volume will?
 - A. Halve
 - B. Double

- C. Remain same
- D. Become zero
- Q3: Volume of gas is 3 L at 300 K. Find volume at 450 K at constant pressure.
 - A. 4.5 L
 - B. 2 L
 - C. 3 L
 - D. 6 L

Answer Key

- Q1: B - Volume is proportional to absolute temperature at constant pressure.
- Q2: B - Doubling temperature doubles volume.
- Q3: A - $V_2 = 3 \times \frac{450}{300} = 4.5$ L.

Memory Aids

- Charles' law: $V \propto T$ at constant pressure
- Volume increases with temperature

Dalton's Law of Partial Pressures

Partial Pressure Concept

- Total pressure of gas mixture = sum of partial pressures of each gas
- At constant T and P , $P_{total} = P_1 + P_2 + P_3 + \dots$
- Partial pressure of i^{th} gas: $P_i = \mu_i RT/V$
- Total pressure: $P_{total} = (RT/V)(\mu_1 + \mu_2 + \dots) = (RT/V)\mu_{total}$
- Partial pressure relation: $P_i = P_{total} \times \frac{\mu_i}{\mu_{total}}$, where mole fraction $\frac{\mu_i}{\mu_{total}}$

Solved Examples

- **Ex 1:** Calculate total pressure if $\mu_1 = 1$ mol, $\mu_2 = 2$ mol, $R = 8.314$ J/mol·K, $T = 300$ K, $V = 0.01$ m³.
 - $P_{total} = \frac{RT}{V}(\mu_1 + \mu_2) = \frac{8.314 \times 300}{0.01} \times 3 = 7.48 \times 10^5$ Pa
- **Ex 2:** Find partial pressure of gas 1 if total pressure is 7.48×10^5 Pa and mole fraction $\frac{\mu_1}{\mu_{total}} = \frac{1}{3}$.
 - $P_1 = P_{total} \times \frac{1}{3} = 2.49 \times 10^5$ Pa

Practice Set

- Q1: Dalton's law states total pressure is?
 - A. Product of partial pressures
 - B. Sum of partial pressures
 - C. Difference of partial pressures
 - D. Ratio of partial pressures
- Q2: Partial pressure of a gas is proportional to?
 - A. Volume
 - B. Mole fraction
 - C. Temperature
 - D. Mass
- Q3: If total pressure is 1 atm and mole fraction of gas A is 0.4, partial pressure of A is?
 - A. 0.4 atm
 - B. 0.6 atm
 - C. 1 atm

- D. 0.8 atm

Answer Key

- Q1: B - Total pressure equals sum of partial pressures of gases.
- Q2: B - Partial pressure proportional to mole fraction.
- Q3: A - Partial pressure = total pressure $\times$ mole fraction = 0.4 atm.

Memory Aids

- Dalton's law: $P_{total} = \sum P_i$
- Partial pressure $P_i = P_{total} \times \text{mole fraction}$

Law of Equipartition of Energy

Concept

- In thermal equilibrium, total energy is equally divided among all energy modes.
- Each mode has average energy $\frac{1}{2}k_B T$, where k_B is Boltzmann constant and T is temperature in Kelvin.
- Each translational and rotational degree of freedom contributes $\frac{1}{2}k_B T$ energy.
- Each vibrational frequency contributes $2 \times \frac{1}{2}k_B T = k_B T$ due to kinetic and potential energy.

Monoatomic Molecules

- Translational kinetic energy per molecule: $E = \frac{1}{2}mv_x^2 + \frac{1}{2}mv_y^2 + \frac{1}{2}mv_z^2$
- Or simplified: $E = \frac{3}{2}k_B T$

Diatomic Molecules (No Vibration)

- Kinetic energy includes translational and rotational parts:
- $E = E_t + E_r = \frac{1}{2}m(v_x^2 + v_y^2 + v_z^2) + \frac{1}{2}I_1\omega_1^2 + \frac{1}{2}I_2\omega_2^2 = \frac{5}{2}k_B T$
- Assumes no vibrational energy.

Diatomic Molecules (With Vibration)

- Vibrational energy term: $E_v = \frac{1}{2}m\left(\frac{dy}{dt}\right)^2 + \frac{1}{2}ky^2$
- Total energy: $E = E_t + E_r + E_v$

Solved Examples

- **Ex 1:** Calculate average KE of a monoatomic gas molecule at temperature $T = 300K$.
 - Use $E = \frac{3}{2}k_B T$
 - Substitute $k_B = 1.38 \times 10^{-23} J/K$
 - $E = \frac{3}{2} \times 1.38 \times 10^{-23} \times 300 = 6.21 \times 10^{-21} J$
- **Ex 2:** Find total energy of a diatomic molecule (no vibration) at $400K$.
 - Use $E = \frac{5}{2}k_B T$
 - $E = \frac{5}{2} \times 1.38 \times 10^{-23} \times 400 = 1.38 \times 10^{-20} J$

Practice Set

- **Q1:** What is the average energy per degree of freedom at temperature T ?
 - A) $k_B T$
 - B) $\frac{1}{2} k_B T$
 - C) $2k_B T$
 - D) $\frac{3}{2} k_B T$
- **Q2:** The kinetic energy of a monoatomic gas molecule at temperature 300 K is approximately:
 - A) $1.38 \times 10^{-21} \text{ J}$
 - B) $6.21 \times 10^{-21} \text{ J}$
 - C) $4.14 \times 10^{-21} \text{ J}$
 - D) $9.0 \times 10^{-21} \text{ J}$
- **Q3:** For a diatomic molecule with vibration considered, total energy is:
 - A) $\frac{5}{2} k_B T$
 - B) $3k_B T$
 - C) $\frac{7}{2} k_B T$
 - D) $4k_B T$

Answer Key

- **Q1:** B) $\frac{1}{2} k_B T$ - Each degree of freedom contributes half $k_B T$ energy.
- **Q2:** B) $6.21 \times 10^{-21} \text{ J}$ - Using $\frac{3}{2} k_B T$ at 300 K .
- **Q3:** D) $4k_B T$ - Vibration adds $k_B T$ to $\frac{5}{2} k_B T$ making total $4k_B T$.

Memory Aids

- Energy per mode = $\frac{1}{2} k_B T$
- Monoatomic KE = $\frac{3}{2} k_B T$
- Diatomic (no vibration) KE = $\frac{5}{2} k_B T$
- Vibration adds $k_B T$ energy

Specific Heat Capacity

Key Concepts

- Specific heat capacity is the amount of heat required to raise the temperature of a substance by 1 K .
- For gases, specific heats at constant volume (C_V) and constant pressure (C_p) are important.
- Relation for ideal gases: $C_p - C_V = R$ (gas constant).

Solved Examples

- **Ex 1:** Calculate C_p if $C_V = \frac{3}{2} R$.
 - Use $C_p = C_V + R = \frac{3}{2} R + R = \frac{5}{2} R$.

Practice Set

- **Q1:** What is the relation between C_p and C_V for an ideal gas?
 - A) $C_p = C_V$
 - B) $C_p - C_V = R$

- C) $C_p + C_v = R$
- D) $C_p < C_v$
- Q2: If $C_v = \frac{5}{2}R$, find C_p .
 - A) $\frac{7}{2}R$
 - B) $\frac{5}{2}R$
 - C) $\frac{3}{2}R$
 - D) $\frac{9}{2}R$
- Q3: For an ideal gas, if $C_p = 20.8 \text{ J/mol}\cdot\text{K}$ and $R = 8.3 \text{ J/mol}\cdot\text{K}$, find C_v .
 - A) $12.5 \text{ J/mol}\cdot\text{K}$
 - B) $28.9 \text{ J/mol}\cdot\text{K}$
 - C) $20.8 \text{ J/mol}\cdot\text{K}$
 - D) $8.3 \text{ J/mol}\cdot\text{K}$

Answer Key

- Q1: B - By definition for ideal gases, $C_p - C_v = R$.
- Q2: A - $C_p = C_v + R = \frac{5}{2}R + R = \frac{7}{2}R$.
- Q3: A - $C_v = C_p - R = 20.8 - 8.3 = 12.5 \text{ J/mol}\cdot\text{K}$.

Memory Aids

- $C_p - C_v = R$ for ideal gases.
- Remember $C_v = \frac{f}{2}R$ where f is degrees of freedom.
- Ratio $\gamma = \frac{C_p}{C_v}$ is key for gas behavior.

Monoatomic Gases

Key Concepts

- Degrees of freedom, $f = 3$.
- Total internal energy: $U = \frac{3}{2}N_A k_B T = \frac{3}{2}RT$.
- Specific heat at constant volume: $C_v = \frac{dU}{dT} = \frac{3}{2}R$.
- Specific heat at constant pressure: $C_p = C_v + R = \frac{5}{2}R$.
- Ratio of specific heats: $\gamma = \frac{C_p}{C_v} = \frac{5}{3}$.

Solved Examples

- **Ex 1:** Calculate γ for monoatomic gas.
 - $\gamma = \frac{5}{3} = 1.67$.

Practice Set

- Q1: Degrees of freedom for monoatomic gases?
 - A) 2
 - B) 3
 - C) 5
 - D) 6
- Q2: Value of C_v for monoatomic gases?
 - A) $\frac{3}{2}R$
 - B) $\frac{5}{2}R$

- C) R
- D) $3R$
- Q3: Calculate C_p if $C_v = \frac{3}{2}R$.
 - A) $\frac{3}{2}R$
 - B) $\frac{5}{2}R$
 - C) $2R$
 - D) $3R$

Answer Key

- Q1: B - Monoatomic gases have 3 degrees of freedom (translational).
- Q2: A - $C_v = \frac{3}{2}R$ for monoatomic gases.
- Q3: B - $C_p = C_v + R = \frac{5}{2}R$.

Memory Aids

- Monoatomic gases: $f = 3$, $C_v = \frac{3}{2}R$, $\gamma = \frac{5}{3}$.
- Use $\gamma = \frac{C_p}{C_v}$ to check values.

Diatomic Gases

Key Concepts

- Degrees of freedom, $f = 5$ (3 translational + 2 rotational).
- Total internal energy: $U = \frac{5}{2}RT$.
- Specific heat at constant volume: $C_v = \frac{5}{2}R$.
- Specific heat at constant pressure: $C_p = C_v + R = \frac{7}{2}R$.
- Ratio of specific heats: $\gamma = \frac{7}{5} = 1.4$.
- If vibrational mode is active, $U = \frac{7}{2}RT$, $C_v = \frac{7}{2}R$, $C_p = \frac{9}{2}R$, $\gamma = \frac{9}{7}$.

Solved Examples

- **Ex 1:** Calculate γ for diatomic gas without vibration.
 - $\gamma = \frac{7}{5} = 1.4$.
- **Ex 2:** Calculate C_p if $C_v = \frac{7}{2}R$.
 - $C_p = C_v + R = \frac{9}{2}R$.

Practice Set

- Q1: Degrees of freedom for diatomic gases?
 - A) 3
 - B) 5
 - C) 6
 - D) 7
- Q2: Value of γ for diatomic gases without vibration?
 - A) 1.2
 - B) 1.4
 - C) 1.67
 - D) 1.5
- Q3: If vibrational mode is active, what is C_v ?
 - A) $\frac{5}{2}R$

- B) $\frac{7}{2}R$
- C) $\frac{9}{2}R$
- D) $4R$

Answer Key

- Q1: B - Diatomic gases have 5 degrees of freedom (3 translational + 2 rotational).
- Q2: B - $\gamma = \frac{7}{5} = 1.4$ without vibration.
- Q3: B - With vibration, $C_v = \frac{7}{2}R$.

Memory Aids

- Diatomic gases: $f = 5$, $\gamma = \frac{7}{5}$.
- Vibrational mode adds 1 to f and increases C_v and C_p .

Polyatomic Gases

Key Concepts

- Total internal energy: $U = [(3/2)k_B T + (3/2)k_B T + f k_B T] N_A$.
- Specific heat at constant volume: $C_v = (3 + f)R$.
- Specific heat at constant pressure: $C_p = (4 + f)R$.
- Ratio of specific heats: $\gamma = \frac{C_p}{C_v} = \frac{4+f}{3+f}$.

Solved Examples

- **Ex 1:** Calculate γ if $f = 3$.
 - $\gamma = \frac{4+3}{3+3} = \frac{7}{6} = 1.17$.

Practice Set

- Q1: Expression for C_v in polyatomic gases?
 - A) $(3 + f)R$
 - B) $(4 + f)R$
 - C) fR
 - D) $3R$
- Q2: Expression for γ in polyatomic gases?
 - A) $\frac{3+f}{4+f}$
 - B) $\frac{4+f}{3+f}$
 - C) $\frac{f}{3}$
 - D) $\frac{3}{f}$
- Q3: If $f = 2$, find γ .
 - A) 1.2
 - B) 1.25
 - C) 1.3
 - D) 1.4

Answer Key

- Q1: A - $C_v = (3 + f)R$ by definition.
- Q2: B - $\gamma = \frac{4+f}{3+f}$.
- Q3: B - $\gamma = \frac{4+2}{3+2} = \frac{6}{5} = 1.25$.

Memory Aids

- Polyatomic gases: $C_v = (3 + f)R$, $\gamma = \frac{4+f}{3+f}$.
- Remember to add 1 to numerator and denominator for C_p and ratio.

Specific Heat Capacity of Solids

Key Concepts

- Consider a solid of N atoms vibrating about mean positions.
- Average energy per 1D oscillator: $k_B T$.
- For 3D, average energy: $3k_B T$.
- For 1 mole, total energy: $U = 3RT$.
- Specific heat capacity: $C = \frac{\Delta Q}{\Delta T} = \frac{\Delta U}{\Delta T} = 3R$ (no work done).

Solved Examples

- **Ex 1:** Calculate specific heat capacity of a solid.
 - $C = 3R = 3 \times 8.314 = 24.94 \text{ J/mol}\cdot\text{K}$.

Practice Set

- Q1: Average energy of 1D oscillator?
 - A) $\frac{1}{2}k_B T$
 - B) $k_B T$
 - C) $2k_B T$
 - D) $3k_B T$
- Q2: Total energy for 3D oscillators per mole?
 - A) RT
 - B) $3RT$
 - C) $\frac{3}{2}RT$
 - D) $6RT$
- Q3: Specific heat capacity of solids?
 - A) R
 - B) $2R$
 - C) $3R$
 - D) $4R$

Answer Key

- Q1: B - Average energy per 1D oscillator is $k_B T$.
- Q2: B - For 3D, total energy is $3k_B T N_A = 3RT$.
- Q3: C - Specific heat capacity of solids is $3R$.

Memory Aids

- Specific heat of solids = $3R$ (Dulong-Petit law).
- Energy per oscillator = $k_B T$ in 1D.

Mean Free Path

Concept and Formula

- Mean free path (λ) is the average distance a gas molecule travels between collisions.
- In time t , molecule travels distance vt and collides with molecules in a cylindrical volume $V = \pi d^2 vt$, where d is molecular diameter.
- Number density $n = \frac{N}{V}$, where N is number of molecules in volume V .
- Number of collisions $N = n\pi d^2 vt$.
- Mean free path formula:

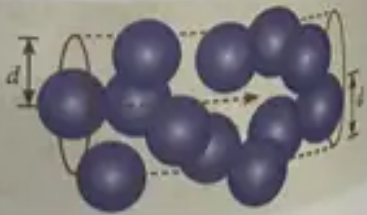
$$\lambda = \frac{vt}{N} = \frac{1}{n\pi d^2}$$

- Exact formula considering all molecular motion:

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

- Effective collision area $A = \pi d^2$.

• The average distance travelled between successive collisions of molecules of a gas is called as mean free path (λ).
Let in time interval t , the molecule moves a distance vt and collides with every molecule in the cylindrical volumes, $V = \pi d^2 vt$ and number of molecules in cylindrical volume be N .

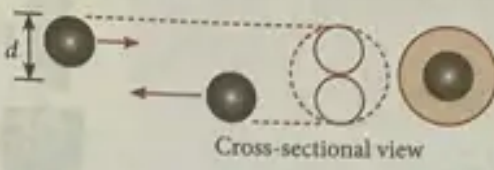


$\therefore$ Number of molecules per unit volume be $n = \frac{N}{V} \Rightarrow N = n\pi d^2 vt$

The mean free path is the total distance divided by the number of collisions.

$$\lambda = \frac{vt}{n\pi d^2 vt} = \frac{1}{n\pi d^2}$$

By considering motion of all the molecules the exact expression is $\lambda = \frac{1}{\sqrt{2}n\pi d^2}$



Cross-sectional view

The effective collision area is $A = \pi d^2$

Solved Examples

- **Ex 1:** Calculate λ if $n = 2 \times 10^{25} \text{ m}^{-3}$ and $d = 3 \times 10^{-10} \text{ m}$.
 - Use $\lambda = \frac{1}{\sqrt{2}n\pi d^2}$
 - $\lambda = \frac{1}{\sqrt{2} \times 2 \times 10^{25} \times \pi \times (3 \times 10^{-10})^2} \approx 4.2 \times 10^{-8} \text{ m}$

Practice Set

- Q1: Mean free path depends inversely on
 - A) Number density n
 - B) Molecular diameter d

- C) Temperature
- D) Pressure
- Q2: If molecular diameter doubles, mean free path
 - A) Doubles
 - B) Halves
 - C) Quadruples
 - D) Quartered
- Q3: For $n = 1 \times 10^{25} \text{ m}^{-3}$, $d = 2 \times 10^{-10} \text{ m}$, calculate λ (use $\pi = 3.14$)
 - A) $7.1 \times 10^{-8} \text{ m}$
 - B) $1.1 \times 10^{-7} \text{ m}$
 - C) $5.6 \times 10^{-8} \text{ m}$
 - D) $3.2 \times 10^{-8} \text{ m}$

Answer Key

- Q1: A - Mean free path $\lambda = \frac{1}{\sqrt{2}n\pi d^2}$ inversely depends on number density n .
- Q2: D - $\lambda \propto \frac{1}{d^2}$, doubling d quarters λ .
- Q3: A - $\lambda = \frac{1}{\sqrt{2} \times 1 \times 10^{25} \times 3.14 \times (2 \times 10^{-10})^2} \approx 7.1 \times 10^{-8} \text{ m}$.

Memory Aids

- Mean free path λ inversely $\propto nd^2$.
- Exact formula includes $\sqrt{2}$ factor.
- Think "collision cylinder" volume $\pi d^2 vt$.

Avogadro's Number

Definition and Value

- Number of units (atoms, molecules, ions, electrons) in one mole of substance.
- Value: $6.022140857 \times 10^{23}$ units/mole.

Solved Examples

- **Ex 1:** Calculate number of molecules in 2 moles of oxygen.
 - Number = moles $\times$ Avogadro's number = $2 \times 6.022 \times 10^{23} = 1.2044 \times 10^{24}$

Practice Set

- Q1: Avogadro's number represents
 - A) Number of atoms in 1 gram of substance
 - B) Number of units in 1 mole
 - C) Number of molecules in 1 liter
 - D) Number of ions in 1 kg
- Q2: Number of molecules in 3 moles of water is
 - A) 1.8×10^{23}
 - B) 1.8×10^{24}
 - C) 6.022×10^{23}
 - D) 3×10^{23}
- Q3: Avogadro's number is approximately
 - A) 6.02×10^{22}

- B) 6.02×10^{23}
- C) 6.02×10^{24}
- D) 6.02×10^{25}

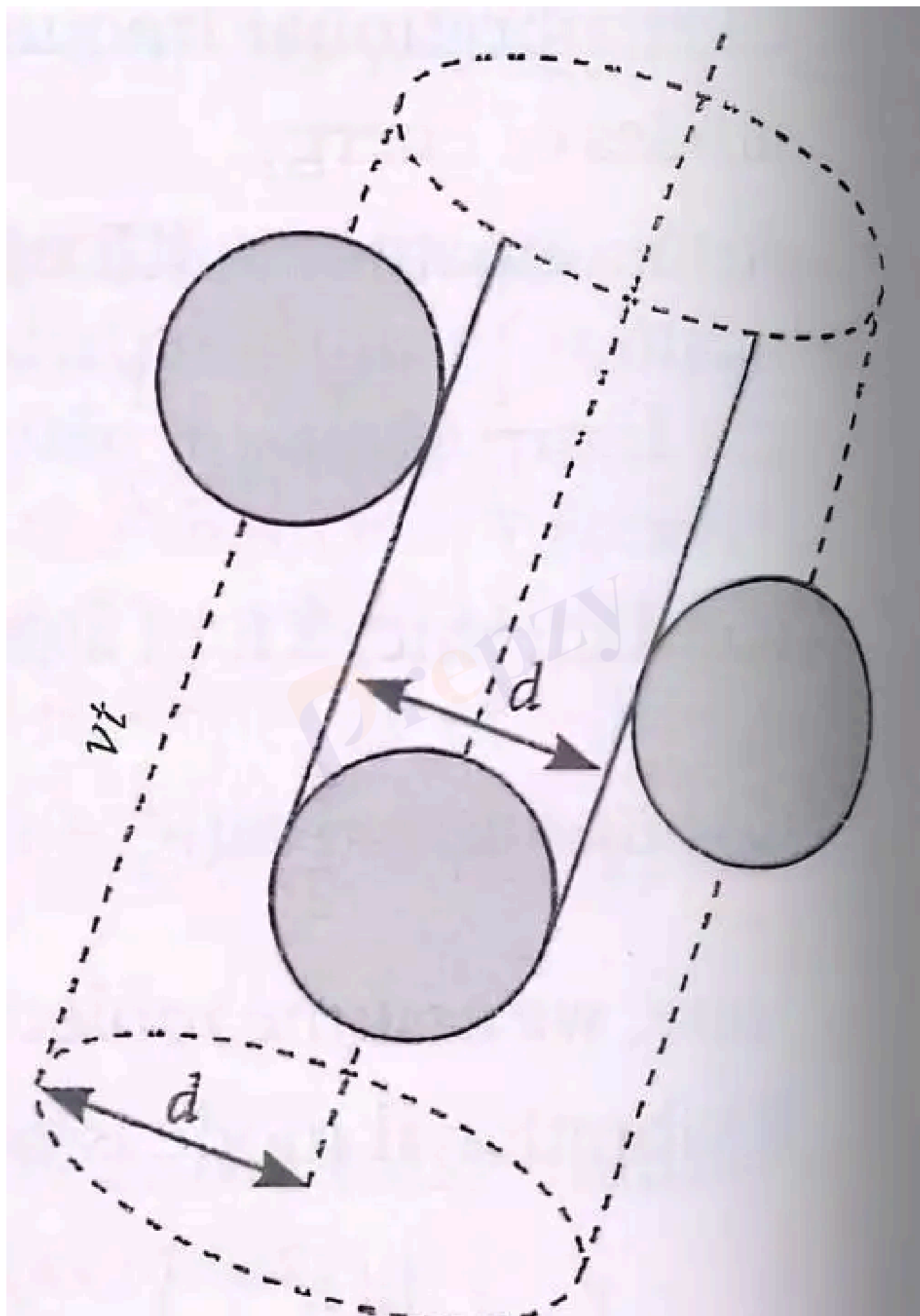
Answer Key

- Q1: B - Avogadro's number is units per mole.
- Q2: B - $3 \times 6.022 \times 10^{23} = 1.8 \times 10^{24}$.
- Q3: B - Standard value is 6.022×10^{23} .

Memory Aids

- Avogadro's number $\approx 6.02 \times 10^{23}$.
- Represents units per mole.
- Used to convert moles to number of particles.

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