CHAPTER 12: KINETIC THEORY

Class 11 CBSE Physics

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📚 12.1 INTRODUCTION

Kinetic theory explains the behaviour of gases based on the idea that the gas consists of rapidly moving atoms or molecules. The kinetic theory was developed in the nineteenth century by **Maxwell, Boltzmann and others**.

🌟 Historical Background

Boyle discovered the law named after him in **1661**. Boyle, Newton and several others tried to explain the behaviour of gases by considering that gases are made up of tiny atomic particles.

Key Achievements of Kinetic Theory:

- ✓ Gives molecular interpretation of pressure and temperature
- ✓ Consistent with gas laws and Avogadro's hypothesis
- ✓ Correctly explains specific heat capacities of many gases
- ✓ Relates measurable properties (viscosity, conduction, diffusion) with molecular parameters
- ✓ Yields estimates of molecular sizes and masses

📚 12.2 MOLECULAR NATURE OF MATTER

🏛 Richard Feynman's Atomic Hypothesis

"All things are made of atoms - little particles that move around in perpetual motion, attracting each other when they are a little distance apart, but repelling upon being squeezed into one another."

Richard Feynman, one of the great physicists of 20th century, considers the discovery that **"Matter is made up of atoms"** to be a very significant one.

IN Atomic Hypothesis in Ancient India and Greece

India - Vaiseshika School (6th century B.C.)

Founded by Kanada:

- **Paramanu** (Sanskrit for smallest particle) - eternal, indivisible, infinitesimal
- **Four kinds of atoms:** Bhoomi (Earth), Ap (water), Tejas (fire), Vayu (air)
- **Akasa** (space) - continuous and inert, no atomic structure
- Atoms combine to form molecules:
  - Dvyanuka (diatomic molecule) - 2 atoms
  - Tryanuka (triatomic molecule) - 3 atoms
- Atomic size estimate in **Lalitavistara** (2nd century B.C.) ≈ 10^{-10} m (close to modern estimate!)

Greece - Democritus (4th century B.C.)

'Atom' means 'indivisible' in Greek

- Water atoms: **smooth and round** → unable to hook, flows easily
- Earth atoms: **rough and jagged** → hold together, form hard substances
- Fire atoms: **thorny** → cause painful burns

🔬 John Dalton's Atomic Theory (~200 years ago)

Dalton proposed atomic theory to explain:

1. **Law of Definite Proportions:** Any compound has fixed proportion by mass of its constituents
2. **Law of Multiple Proportions:** When two elements form more than one compound, for fixed mass of one element, masses of other element are in ratio of small integers

Key Concepts:
• Atoms of one element are identical but differ from other elements
• Small number of atoms combine to form molecules
• Molecules constitute matter

📏 Molecular Sizes and Distances

Property	Value	Description
Atomic Size	~10^{-10} m (1 Å)	Size of an atom
Spacing in Solids/Liquids	~2 Å	Atoms tightly packed
Spacing in Gases	Tens of angstroms	Atoms far apart
Mean Free Path (gases)	~1000s of Å	Average distance between collisions

💡 Important Points:

- In **solids & liquids:** Atoms closely packed, interatomic forces important (long range attraction + short range repulsion)
- In **gases:** Atoms much freer, can travel long distances without colliding
- Gas appears static but has **dynamic equilibrium** - molecules constantly colliding and changing speeds

📚 12.3 BEHAVIOUR OF GASES

Properties of gases are easier to understand than solids and liquids because molecules are far apart and mutual interactions are negligible except during collisions.

🎯 Ideal Gas Equation

$$PV = \mu RT = k_B NT$$

Where:

P = Pressure, V = Volume, T = Absolute Temperature

μ = Number of moles, N = Number of molecules

R = Universal gas constant = 8.314 J mol^{-1} K^{-1}

k_B = Boltzmann constant = 1.38 × 10^{-23} J K^{-1}

🔑 Key Relations:

1. Number of Moles:

$$\mu = M/M_0 = N/N_A$$

M = Mass of gas, M_0 = Molar mass

N_A = Avogadro's number = 6.02×10^{23}

2. Alternative Forms:

$$PV = k_B NT \quad \text{or} \quad P = k_B nT$$

where n = number density (molecules per unit volume)

3. Density Form:

$$P = (\rho RT)/M_0$$

where ρ = mass density

📊 Avogadro's Hypothesis

Equal volumes of all gases at equal temperature and pressure have the same number of molecules.

Standard Conditions (STP):

- Temperature: 273 K (0°C)
- Pressure: 1 atm
- **Molar Volume: 22.4 litres**
- Contains $N_A = 6.02 \times 10^{23}$ molecules

📐 Gas Laws

Law	Statement	Mathematical Form
Boyle's Law	At constant temperature, pressure varies inversely with volume	**PV = constant** (T, μ fixed)
Charles' Law	At constant pressure, volume is proportional to absolute temperature	**$V \propto T$** (P, μ fixed)
Dalton's Law	Total pressure of mixture equals sum of partial pressures	**$P = P_1 + P_2 + \ldots$**

> ***Note:*** *Real gases satisfy ideal gas equation only approximately, more accurately at* ***low pressures*** *and* ***high temperatures****.*

📝 Example 12.1: Molecular Volume Fraction in Water Vapor

Given:

- Density of water = 1000 kg m^{-3}
- Density of water vapour at 100°C and 1 atm = 0.6 kg m^{-3}

Question: Estimate the ratio of molecular volume to total volume in water vapour.

Solution:

Step 1: Volume ratio = Density ratio (for same mass)

Volume of vapour/Volume of liquid = 1000/0.6 = $1/(6 \times 10^{-4})$

Step 2: In liquid state, molecular volume fraction ≈ 1

In vapour state, volume increased by factor $1/(6 \times 10^{-4})$

Answer: Fractional molecular volume in vapour = $\mathbf{6 \times 10^{-4}}$

📝 Example 12.2: Volume of Water Molecule

Given: Data from Example 12.1

Solution:

Step 1: Density of water molecules ≈ density of bulk water = 1000 kg m^{-3}

Step 2: Mass of 1 mole of water = 18 g = 0.018 kg

1 mole contains $N_A = 6 \times 10^{23}$ molecules

Step 3: Mass of one water molecule:

$m = 0.018/(6 \times 10^{23}) = 3 \times 10^{-26}$ kg

Step 4: Volume of water molecule:

$V = m/\rho = (3 \times 10^{-26})/(1000) =$ **3×10^{-29} m^3**

Step 5: If $V = (4/3)\pi r^3$

Radius ≈ 2×10^{-10} m = 2 Å

📚 12.4 KINETIC THEORY OF AN IDEAL GAS

🎯 Basic Assumptions

1. Gas consists of large number of molecules in **incessant random motion**
2. Molecules move in **straight lines** according to Newton's first law
3. Average distance between molecules >> molecular size
4. **Intermolecular forces negligible** except during collisions
5. All collisions are **perfectly elastic** (KE conserved)
6. Molecules collide with walls and change momentum

📊 Derivation: Pressure of Ideal Gas

Consider a cubic container of side l

Step 1: Molecule with velocity (v_x, v_y, v_z) hits wall parallel to yz-plane

After elastic collision: velocity becomes ($-v_x$, v_y, v_z)

Step 2: Change in momentum of molecule:

$\Delta p = -mv_x - (mv_x) = -2mv_x$

Step 3: Momentum imparted to wall = $\mathbf{2mv_x}$

Step 4: Molecules hitting wall in time Δt:

All molecules within distance $v_x\Delta t$ from wall

Volume = $A \times v_x\Delta t$ (where $A = l^2$ is wall area)

Number of molecules = $n \times A \times v_x\Delta t$ (n = number density)

Only half moving toward wall: $\mathbf{(1/2) \times n \times A \times v_x\Delta t}$

Step 5: Total momentum transferred:

$Q = (2mv_x) \times (1/2) \times n \times A \times v_x\Delta t = \mathbf{nm\langle v_x^2\rangle A\Delta t}$

Step 6: Pressure = Force/Area = $(Q/\Delta t)/A$

$$P = nm\langle v_x^2\rangle$$

Step 7: For isotropic gas (no preferred direction):

$\langle v_x^2\rangle = \langle v_y^2\rangle = \langle v_z^2\rangle = (1/3)\langle v^2\rangle$

Final Result:

$$P = (1/3)nm\langle v^2 \rangle$$

or equivalently

$$PV = (1/3)Nm\langle v^2 \rangle$$

where N = total number of molecules

🌡️ Kinetic Interpretation of Temperature

Step 1: From kinetic theory: $PV = (2/3)N \times (1/2)m\langle v^2 \rangle$

Step 2: Total translational kinetic energy:

$$E = N \times (1/2)m\langle v^2 \rangle$$

Therefore: $\mathbf{PV = (2/3)E}$

Step 3: Combining with ideal gas equation $PV = Nk_BT$:

$$E = (3/2)Nk_BT$$

Average KE per molecule $= (1/2)m\langle v^2 \rangle = (3/2)k_BT$

🌟 Key Insights:

- ✓ Temperature is a measure of average kinetic energy of molecules
- ✓ Average KE is **independent** of nature of gas (monatomic, diatomic, etc.)
- ✓ Depends only on **absolute temperature T**
- ✓ Internal energy of ideal gas depends only on temperature

📈 Root Mean Square Speed

$$\mathbf{v_{rms} = \sqrt{\langle v^2 \rangle} = \sqrt{(3k_BT/m)} = \sqrt{(3RT/M_0)}}$$

At same temperature, lighter molecules have greater rms speed!

📝 Example: RMS Speed of Nitrogen at 300 K

Given: T = 300 K, M_0 = 28 u for N_2

Step 1: $m = M_0/N_A = 28/(6.02 \times 10^{23}) = 4.65 \times 10^{-26}$ kg

Step 2: $\langle v^2 \rangle = 3k_BT/m = 3 \times (1.38 \times 10^{-23}) \times 300/(4.65 \times 10^{-26})$

$\langle v^2 \rangle = (516)^2\ m^2s^{-2}$

Answer: v_{rms} = 516 m/s (comparable to speed of sound!)

📚 12.5 LAW OF EQUIPARTITION OF ENERGY

🎯 Degrees of Freedom

Degree of Freedom: Number of independent coordinates needed to specify position/configuration

Motion Type	Degrees of Freedom	Description
Linear (1D)	1	Motion along a line
Planar (2D)	2	Motion in a plane
Spatial (3D)	3	Motion in space

⚡ Energy Distribution

Kinetic Energy of Single Molecule:

$$\varepsilon_t = (1/2)mv_x^2 + (1/2)mv_y^2 + (1/2)mv_z^2$$

Average energy per degree of freedom = $(1/2)k_BT$

🔬 Molecular Energy Modes

Gas Type	Translational	Rotational	Vibrational	Total DOF
Monatomic (Ar, He)	3	0	0	**3**

Diatomic (O_2, N_2) - rigid	3	2	0	**5**
Diatomic - with vibration	3	2	2	**7**
Polyatomic	3	3	f modes	**6 + f**

💡 Law of Equipartition of Energy

"In thermal equilibrium at temperature T, energy is equally distributed in all modes, with each mode having average energy (1/2)k_BT"

Important Points:

- Each **translational** DOF contributes: (1/2)k_BT
- Each **rotational** DOF contributes: (1/2)k_BT
- Each **vibrational** mode contributes: k_BT (KE + PE)

12.6 SPECIFIC HEAT CAPACITY

Monatomic Gases

Degrees of freedom: 3 (translational only)

Internal energy per mole:

$U = (3/2)N_A k_B T = \mathbf{(3/2)RT}$

Molar specific heat at constant volume:

$$C_v = dU/dT = (3/2)R$$

Using $C_p - C_v = R$:

$$C_p = (5/2)R$$

Ratio of specific heats:

$$\gamma = C_p/C_v = 5/3 = 1.67$$

🔥 Diatomic Gases (Rigid Rotator)

Degrees of freedom: 3 (translational) + 2 (rotational) = **5**

Internal energy per mole:

$U = (5/2)N_A k_B T = \mathbf{(5/2)RT}$

$$C_v = (5/2)R$$
$$C_p = (7/2)R$$
$$\gamma = 7/5 = 1.40$$

🔥 Diatomic Gases (With Vibration)

Degrees of freedom: 3 (trans) + 2 (rot) + 2 (vib) = **7**

Internal energy per mole:

$U = (7/2)N_A k_B T = \mathbf{(7/2)RT}$

$$C_v = (7/2)R$$
$$C_p = (9/2)R$$
$$\gamma = 9/7 = 1.29$$

🔥 Polyatomic Gases

With f vibrational modes:

$$C_v = (3 + f)R$$
$$C_p = (4 + f)R$$
$$\gamma = (4 + f)/(3 + f)$$

📊 Summary Table

Gas Type	C_v (J mol⁻¹ K⁻¹)	C_p (J mol⁻¹ K⁻¹)	γ	Example
Monatomic	12.5	20.8	1.67	He, Ar, Ne
Diatomic (rigid)	20.8	29.1	1.40	H_2, N_2, O_2
Triatomic	24.93	33.24	1.33	H_2O, CO_2

> ***Note:*** *R = 8.31 J mol⁻¹ K⁻¹ is universal gas constant*

📚 12.7 MEAN FREE PATH

Mean Free Path (λ): Average distance a molecule travels between successive collisions

Why molecules don't disperse immediately?

Despite high speeds (~500 m/s), molecules undergo frequent collisions, causing their paths to constantly deflect.

📐 Derivation of Mean Free Path

Consider:

- Molecules as spheres of diameter **d**
- Number density = **n** (molecules per unit volume)
- Average speed = ⟨**v**⟩

Step 1: In time Δt, molecule sweeps volume = $\pi d^2 \langle v \rangle \Delta t$

Step 2: Number of collisions = $n\pi d^2 \langle v \rangle \Delta t$

Step 3: Time between collisions:

$\tau = 1/(n\pi d^2 \langle v \rangle)$

Step 4: Mean free path = $\langle v \rangle \times \tau$

More exact treatment (considering relative velocity):

$$\lambda = 1/(\sqrt{2}\, \pi n d^2)$$

📝 Example: Mean Free Path in Air at STP

Given:

- $T = 300$ K
- $\langle v \rangle = 485$ m/s for nitrogen
- $d = 2 \times 10^{-10}$ m

Step 1: At STP, $n = (0.02 \times 10^{23})/(22.4 \times 10^{-3}) = 2.7 \times 10^{25}\ m^{-3}$

Step 2: $\tau = 1/(\sqrt{2} \times \pi \times n \times d^2 \times \langle v \rangle)$

$\tau = 1/(\sqrt{2} \times \pi \times 2.7 \times 10^{25} \times (2 \times 10^{-10})^2 \times 485)$

$\tau \approx 6.1 \times 10^{-10}$ s

Step 3: $\lambda = \langle v \rangle \times \tau = 485 \times 6.1 \times 10^{-10}$

$\lambda \approx 2.9 \times 10^{-7}$ m ≈ 1500d

Conclusion: Mean free path is about 1500 times the molecular diameter!

💡 Key Points about Mean Free Path:

- ✓ Depends **inversely** on number density (n) and molecular size (d)
- ✓ In highly evacuated tube, λ can be as large as tube length
- ✓ Large λ leads to typical gaseous behavior
- ✓ Gases cannot be confined without a container

🎯 SUMMARY

📌 Key Formulas to Remember

Concept	Formula
Ideal Gas Equation	$PV = \mu RT = k_B NT$
Pressure (Kinetic Theory)	$P = (1/3)nm\langle v^2\rangle$
Average KE per molecule	$(1/2)m\langle v^2\rangle = (3/2)k_B T$
RMS Speed	$v_{rms} = \sqrt{(3k_B T/m)} = \sqrt{(3RT/M_0)}$
Internal Energy	$U = (f/2)\mu RT$, where f = degrees of freedom
Mean Free Path	$\lambda = 1/(\sqrt{2}\, \pi n d^2)$
$C_p - C_v$	R (for any ideal gas)

📊 Specific Heat Capacities

Gas Type	DOF	C_v	C_p	γ
Monatomic	3	(3/2)R	(5/2)R	5/3
Diatomic (rigid)	5	(5/2)R	(7/2)R	7/5
Diatomic (vibrating)	7	(7/2)R	(9/2)R	9/7
Polyatomic	6+f	(3+f)R	(4+f)R	(4+f)/(3+f)

🌟 Important Constants

- **R** = 8.314 J mol^{-1} K^{-1} (Universal gas constant)
- **k_B** = 1.38 × 10^{-23} J K^{-1} (Boltzmann constant)
- **N_A** = 6.02 × 10^{23} (Avogadro's number)
- **Molar volume at STP** = 22.4 litres
- **1 atm** = 1.01 × 10^5 Pa

🎓 IMPORTANT EXAM QUESTIONS

Section A: Multiple Choice Questions (1 mark each)

Q1. The rms speed of gas molecules is:

(a) $\sqrt{(RT/M_0)}$
(b) $\sqrt{(3RT/M_0)}$ ✓
(c) $\sqrt{(2RT/M_0)}$
(d) $\sqrt{(RT/3M_0)}$

Answer: (b)

Explanation: From kinetic theory, $(1/2)m\langle v^2\rangle = (3/2)k_BT$, solving gives $v_{rms} = \sqrt{(3k_BT/m)} = \sqrt{(3RT/M_0)}$

Q2. For a monatomic ideal gas, the ratio $\gamma = C_p/C_v$ is:

(a) 7/5
(b) 9/7
(c) 5/3 ✓
(d) 4/3

Answer: (c)

Explanation: Monatomic gas has 3 translational DOF. $C_v = (3/2)R$, $C_p = (5/2)R$, $\gamma = 5/3$

Q3. Mean free path is inversely proportional to:

(a) Temperature
(b) Pressure ✓
(c) Volume
(d) Speed

Answer: (b)

Explanation: $\lambda = 1/(\sqrt{2}\,\pi n d^2)$. Since $n \propto P$ (at constant T), $\lambda \propto 1/P$

Q4. At what temperature will oxygen molecules have same rms speed as helium at 300 K? (M_{He} = 4, M_{O_2} = 32)

(a) 1200 K
(b) 2400 K ✓
(c) 300 K
(d) 600 K

Answer: (b)

Explanation: $v_{rms} = \sqrt{3RT/M_0}$, so T/M_0 must be same. (300/4) = (T/32), T = 2400 K

Section B: Short Answer Questions (2-3 marks)

Q5. State law of equipartition of energy. How many degrees of freedom does a diatomic molecule have?

Answer: Law of equipartition of energy states that in thermal equilibrium at temperature T, the total energy is equally distributed in all energy modes, with each mode having average energy $(1/2)k_BT$.

A diatomic molecule has:

- 3 translational DOF
- 2 rotational DOF
- Total = **5 DOF** (as rigid rotator)
- With vibration: **7 DOF**

Q6. Why does mean free path increase with decrease in pressure?

Answer: Mean free path $\lambda = 1/(\sqrt{2}\ \pi n d^2)$, where n is number density. As pressure decreases (at constant temperature), number density n decreases according to ideal gas law ($P = nk_BT$). Since $\lambda \propto 1/n$, when n decreases, λ increases. Lower pressure means molecules are farther apart, so they travel longer distances between collisions.

Q7. Calculate molar specific heat at constant volume for oxygen gas.

Answer: Oxygen (O_2) is diatomic. As rigid rotator:

- Degrees of freedom = 5
- $C_V = (f/2)R = (5/2)R$
- $C_V = (5/2) \times 8.314$
- **$C_V = 20.8\ J\ mol^{-1}\ K^{-1}$**

Section C: Long Answer Questions (5 marks)

Q8. Derive an expression for pressure exerted by an ideal gas from kinetic theory. Hence obtain the relation between average kinetic energy and temperature.

Answer: [Complete derivation as shown in section 12.4]

Key Steps:

1. Consider molecule hitting wall
2. Calculate momentum change = $2mv_x$
3. Find number of molecules hitting wall in time Δt
4. Calculate total momentum transfer
5. Derive $P = (1/3)nm\langle v^2 \rangle$
6. Compare with ideal gas equation to get $(1/2)m\langle v^2 \rangle = (3/2)k_B T$

Q9. Using law of equipartition of energy, derive expressions for molar specific heats of monatomic and diatomic gases.

Answer: [Complete derivation as shown in section 12.6]

🏆 FINAL SUCCESS TIPS

🌟 Master These Concepts:

- ✨ Kinetic theory connects macroscopic (P, V, T) and microscopic (molecular motion)
- ✨ Temperature is measure of average kinetic energy - fundamental insight!
- ✨ Law of equipartition - each DOF contributes $(1/2)k_BT$
- ✨ Mean free path derivation is frequently asked
- ✨ Specific heat calculations for different gases
- ✨ Remember: $C_p - C_v = R$ for ALL ideal gases
- ✨ Practice numerical problems on rms speed
- ✨ Understand difference between monatomic, diatomic, polyatomic gases

💪 KEY TO SUCCESS:

"Master the Derivations - They Form the Foundation of Modern Physics!"

📚 Must Practice:

- ☑ Pressure derivation from kinetic theory (5 marks)
- ☑ 10 numerical problems on rms speed
- ☑ Specific heat calculations for all gas types
- ☑ Mean free path problems
- ☑ Degree of freedom concept thoroughly
- ☑ All NCERT examples and exercises

📚 Study Material Information

This comprehensive study material on **Kinetic Theory** has been prepared following the latest CBSE curriculum and examination pattern for Class 11 Physics. The content includes detailed explanations of molecular nature of matter, kinetic theory of gases, law of equipartition of energy, specific heat capacities, mean free path, important derivations from NCERT (pressure from kinetic theory, RMS speed, specific heats), and practice questions aligned with current board exam format to help students achieve excellence.

Key Features of This Material:

- Complete chapter coverage with conceptual clarity
- Historical perspective (Kanada, Democritus, Dalton, Maxwell, Boltzmann)
- Step-by-step derivations with clear explanations
- Kinetic interpretation of temperature and pressure
- Law of equipartition of energy with applications
- Comprehensive specific heat calculations
- Mean free path concept and calculations
- Solved examples from NCERT
- Multiple choice and descriptive practice questions
- Formula summary and important constants
- Exam-focused tips and strategies

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Disclaimer: This material is prepared as a comprehensive study aid for Class 11 students. While every effort has been made to ensure accuracy and alignment with CBSE curriculum, students are advised to refer to their NCERT textbooks and official CBSE guidelines for examination preparation. This material covers Chapter 12: Kinetic Theory from Class 11 Physics NCERT textbook (Reprint 2025-26).