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Kinetic Theory Class 11 Notes | CBSE Physics Chapter 12

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Kinetic Theory Class 11 Notes | CBSE Physics Chapter 12

Kinetic Theory is Chapter 12 of CBSE Class 11 Physics — the chapter that finally explains *why* gases behave the way they do. Instead of treating a gas as a mysterious fluid, kinetic theory pictures it as countless tiny molecules flying around, colliding, and bouncing off the walls. From that single picture you can derive the gas laws, define temperature, and even predict the speed of an air molecule in this room.

By the end of these notes you will be able to state the postulates of kinetic theory, derive the pressure of an ideal gas, link temperature to molecular motion, calculate RMS speed, apply the law of equipartition of energy, and find mean free path. This chapter carries roughly 4–5 marks in boards and is a reliable scorer in NEET and JEE, linking directly to Thermodynamics.

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Key Concepts

1. Molecular Nature of Matter

All matter is made up of tiny particles — atoms and molecules — that are in continuous motion. This idea, first proposed by John Dalton, explains why matter exists as solids, liquids, and gases depending on how strongly the molecules are held together.

In a **solid**, molecules are tightly packed and only vibrate. In a **liquid**, they are loosely bound and can slide past each other. In a **gas**, molecules are far apart, move freely at high speed, and the intermolecular forces are almost negligible.

Key idea: Gases are the simplest state to model because the molecules barely interact except during collisions — this is exactly why kinetic theory works best for gases.

2. Behaviour of Gases and the Gas Laws

Real gases at low pressure and high temperature behave almost like an **ideal gas**. The experimental gas laws describe how pressure (P), volume (V), and temperature (T) of a fixed amount of gas are related.

- **Boyle's Law** (constant T): $PV = \text{constant}$, so $P \propto 1/V$.
- **Charles's Law** (constant P): $V/T = \text{constant}$, so $V \propto T$.
- **Gay-Lussac's Law** (constant V): $P/T = \text{constant}$, so $P \propto T$.
- **Avogadro's Law**: equal volumes of all gases at the same T and P contain equal numbers of molecules.

[DIAGRAM: P–V curve for Boyle's law (a rectangular hyperbola) alongside the straight-line V–T graph for Charles's law passing through $-273.15\text{ }^{\circ}\text{C}$.]

3. Ideal Gas Equation

Combining the three gas laws with Avogadro's law gives the **ideal gas equation**, the master equation of this chapter.

$$PV = nRT$$

- **n** = number of moles, **R** = universal gas constant = $8.314\text{ J mol}^{-1}\text{ K}^{-1}$
- In terms of number of molecules N: **PV = NkT**, where k is the Boltzmann constant.
- **Boltzmann constant k** = $R/N_a = 1.38 \times 10^{-23}\text{ J K}^{-1}$

An **ideal gas** is one that obeys $PV = nRT$ at all pressures and temperatures. No real gas is perfectly ideal, but most gases come close at low density.

4. Postulates of Kinetic Theory of Gases

Kinetic theory rests on a set of simplifying assumptions about gas molecules. These postulates let us derive gas behaviour from pure mechanics.

- A gas consists of a very large number of identical molecules in random motion.
 - The size of a molecule is negligible compared to the average distance between molecules.
 - Molecules exert no force on each other except during collisions.
 - All collisions (molecule–molecule and molecule–wall) are perfectly **elastic**, so kinetic energy is conserved.
 - The time of a collision is negligible compared to the time between collisions.
 - Between collisions, molecules move in straight lines obeying Newton's laws.
-

5. Pressure of an Ideal Gas

Pressure arises because molecules continuously strike the walls of the container and transfer momentum. Applying Newton's laws to these collisions gives the central result of kinetic theory.

$$P = (1/3)(mN/V) \bar{v}^2 = (1/3)\rho \bar{v}^2$$

- m = mass of one molecule, N = number of molecules, V = volume, ρ = density.
- $\bar{v}^2$ is the mean of the squares of molecular speeds (mean-square speed).

This can also be written as $PV = (1/3)Nm \bar{v}^2$, directly connecting a macroscopic quantity (pressure) to microscopic motion.

6. Kinetic Interpretation of Temperature

Comparing $P = (1/3)(Nm/V)\bar{v}^2$ with $PV = NkT$ gives a remarkable result: temperature is a direct measure of the average kinetic energy of molecules.

$$(1/2)m \bar{v}^2 = (3/2)kT$$

- Average translational KE per molecule = $(3/2)kT$.
- Average KE per mole = $(3/2)RT$.
- This energy depends **only on temperature**, not on the nature, mass, or pressure of the gas.

Key idea: At absolute zero ($T = 0$ K), molecular translational motion would theoretically cease.

7. RMS Speed of Gas Molecules

The **root-mean-square (RMS) speed** is the square root of the mean-square speed — a useful single number for the "typical" molecular speed.

$$v_{\text{rms}} = \sqrt{\bar{v}^2} = \sqrt{(3kT/m)} = \sqrt{(3RT/M)}$$

- M = molar mass of the gas.
- $v_{\text{rms}} \propto \sqrt{T}$ — hotter gas means faster molecules.
- $v_{\text{rms}} \propto 1/\sqrt{M}$ — lighter molecules (like H_2) move faster than heavier ones (like O_2) at the same temperature.

Two other useful speeds: **average speed** $v_{\text{avg}} = \sqrt{(8RT/\pi M)}$ and **most probable speed** $v_{\text{mp}} = \sqrt{(2RT/M)}$. Their ratio is $v_{\text{mp}} : v_{\text{avg}} : v_{\text{rms}} = 1 : 1.128 : 1.224$.

8. Degrees of Freedom

The **degrees of freedom (f)** of a molecule is the number of independent ways it can store energy — that is, the number of independent coordinates needed to describe its motion.

Type of gas	Translational	Rotational	Total f (room temp)
Monatomic (He, Ar)	3	0	3
Diatomic (O ₂ , N ₂)	3	2	5
Triatomic (linear, CO ₂)	3	2	5
Triatomic (non-linear, H ₂ O)	3	3	6

At high temperatures, diatomic and polyatomic molecules also gain **vibrational** degrees of freedom, each contributing 2 to f.

9. Law of Equipartition of Energy

The **law of equipartition of energy** states that in thermal equilibrium, the total energy is shared equally among all degrees of freedom, and each degree of freedom contributes an average energy of **$(1/2)kT$** per molecule.

- Each translational and each rotational degree of freedom $\rightarrow (1/2)kT$ per molecule.
- Each vibrational mode $\rightarrow 2 \times (1/2)kT = kT$ (it has both kinetic and potential energy terms).

So total internal energy per mole = **$U = (f/2)RT$** , where f is the number of degrees of freedom.

10. Specific Heat Capacity of Gases

Using $U = (f/2)RT$, the molar specific heats of an ideal gas follow directly from the degrees of freedom.

- **At constant volume:** $C_v = (f/2)R$
- **At constant pressure:** $C_p = C_v + R = (f/2 + 1)R$ (Mayer's relation)
- **Ratio:** $\gamma = C_p/C_v = 1 + 2/f$

Gas type	f	C _v	C _p	γ
Monatomic	3	$(3/2)R$	$(5/2)R$	1.67
Diatomic	5	$(5/2)R$	$(7/2)R$	1.40
Polyatomic (non-linear)	6	$3R$	$4R$	1.33

11. Mean Free Path

The **mean free path (λ)** is the average distance a molecule travels between two successive collisions. The more crowded or larger the molecules, the shorter this distance.

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

- d = diameter of a molecule, n = number of molecules per unit volume.
- $\lambda \propto 1/n$, so $\lambda \propto T/P$ (using $n = P/kT$) — at higher temperature λ increases, at higher pressure λ decreases.
- For air at NTP, $\lambda \approx 10^{-7}$ m.

12. Avogadro's Number and Avogadro's Law

Avogadro's number ($N_a = 6.022 \times 10^{23} \text{ mol}^{-1}$) is the number of molecules in one mole of any substance.

Avogadro's law: equal volumes of all gases under the same conditions of temperature and pressure contain an equal number of molecules. At STP, one mole of any ideal gas occupies **22.4 litres**.

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	4–5 marks	Postulates, RMS speed, equipartition, specific heats
JEE Main / Advanced	1–2 questions	Pressure derivation, RMS speed, γ , mean free path
NEET	1–2 questions	Kinetic energy–temperature relation, RMS speed, degrees of freedom

[TABLE: Question-type split — VSA (1 mark): definitions, value of $R/k/N_a$; SA (2–3 marks): RMS speed numericals, $C_p - C_v = R$, degrees of freedom; LA (5 marks): derivation of pressure of an ideal gas, kinetic interpretation of temperature.]

Important Definitions

Term	Definition
Ideal gas	A gas that obeys $PV = nRT$ at all temperatures and pressures
Pressure (kinetic)	Force per unit area from molecular collisions: $P = (1/3)\rho v^2$
Mean-square speed	Average of the squares of molecular speeds, v^2
RMS speed	Square root of mean-square speed: $v_{\text{rms}} = \sqrt{3RT/M}$
Boltzmann constant	$k = R/N_a = 1.38 \times 10^{-23} \text{ J K}^{-1}$
Degrees of freedom	Number of independent ways a molecule can store energy
Equipartition of energy	Each degree of freedom carries average energy $(1/2)kT$ per molecule
Mean free path	Average distance between successive molecular collisions: $\lambda = 1/(\sqrt{2} \pi d^2 n)$
Avogadro's number	Number of molecules in one mole: $N_a = 6.022 \times 10^{23} \text{ mol}^{-1}$
Specific heat ratio	$\gamma = C_p/C_v = 1 + 2/f$

Solved Examples

Example 1

Calculate the RMS speed of oxygen molecules at 300 K. ($M = 32 \text{ g/mol} = 0.032 \text{ kg/mol}$, $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$)

Answer: $v_{\text{rms}} = \sqrt{3RT/M} = \sqrt{3 \times 8.314 \times 300 / 0.032} = \sqrt{233\,831} \approx \mathbf{484 \text{ m/s}}$.

Example 2

At what temperature will the RMS speed of hydrogen molecules be double its value at 300 K?

Answer: Since $v_{\text{rms}} \propto \sqrt{T}$, doubling the speed requires T to become 4 times. $T = 4 \times 300 = \mathbf{1200 \text{ K}}$.

Example 3

Find the average translational kinetic energy of a gas molecule at 27 °C. ($k = 1.38 \times 10^{-23} \text{ J K}^{-1}$)

Answer: $T = 27 + 273 = 300 \text{ K}$. $KE = (3/2)kT = (3/2)(1.38 \times 10^{-23})(300) = \mathbf{6.21 \times 10^{-21} \text{ J}}$.

Example 4

For a diatomic gas, find $C_{v,m}$, $C_{p,m}$, and γ . ($R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$)

Answer: $f = 5$, so $C_{v,m} = (5/2)R = 20.8 \text{ J mol}^{-1} \text{ K}^{-1}$, $C_{p,m} = (7/2)R = 29.1 \text{ J mol}^{-1} \text{ K}^{-1}$, and $\gamma = C_{p,m}/C_{v,m} = \mathbf{1.4}$.

Example 5

The RMS speed of a gas at 300 K is 500 m/s. What is its RMS speed at 1200 K?

Answer: $v_{\text{rms}} \propto \sqrt{T}$, so $v_2 = v_1\sqrt{(T_2/T_1)} = 500 \times \sqrt{(1200/300)} = 500 \times 2 = \mathbf{1000 \text{ m/s}}$.

Example 6

Calculate the number of molecules in 2 g of hydrogen gas. ($N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$, molar mass of $\text{H}_2 = 2 \text{ g/mol}$)

Answer: Moles $n = 2/2 = 1 \text{ mol}$. Number of molecules = $n \times N_A = 1 \times 6.022 \times 10^{23} = \mathbf{6.022 \times 10^{23} \text{ molecules}}$.

Important Questions for Board Exams

1-Mark Questions (VSA)

1. State the value and SI unit of Boltzmann's constant.
2. On what factor does the average kinetic energy of a gas molecule depend?
3. Why does a lighter gas diffuse faster than a heavier gas at the same temperature?
4. What are the degrees of freedom of a monatomic gas molecule?
5. Define mean free path.

2–3-Mark Questions (SA)

1. State the postulates of the kinetic theory of gases.
2. Derive the relation $C_{p,m} - C_{v,m} = R$ for an ideal gas.
3. Show that the average kinetic energy of a molecule is $(3/2)kT$ and is independent of the nature of the gas.
4. Explain how mean free path depends on temperature and pressure.

5-Mark Questions (LA)

1. Derive an expression for the pressure exerted by an ideal gas on the walls of its container using kinetic theory.
 2. State the law of equipartition of energy and use it to find $C_{v,m}$, $C_{p,m}$, and γ for monatomic and diatomic gases.
 3. Define RMS speed and derive $v_{\text{rms}} = \sqrt{(3RT/M)}$. Hence compare the RMS speeds of two gases of different molar masses at the same temperature.
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Quick Revision Points

- Ideal gas equation: $PV = nRT = NkT$; $k = R/N_a = 1.38 \times 10^{-23} \text{ J K}^{-1}$
- Pressure of ideal gas: $P = (1/3)\rho v^2 = (1/3)(Nm/V)v^2$
- Temperature: $(1/2)m v^2 = (3/2)kT \rightarrow$ average KE depends only on T
- RMS speed: $v_{\text{rms}} = \sqrt{(3RT/M)}$; $v_{\text{rms}} \propto \sqrt{T}$ and $\propto 1/\sqrt{M}$
- Speed ratio: $v_{\text{mp}} : v_{\text{avg}} : v_{\text{rms}} = 1 : 1.128 : 1.224$
- Degrees of freedom: monatomic 3, diatomic 5, non-linear triatomic 6
- Equipartition: each degree of freedom $\rightarrow (1/2)kT$ per molecule; $U = (f/2)RT$
- Specific heats: $C_v = (f/2)R$, $C_p = C_v + R$, $\gamma = 1 + 2/f$
- γ values: monatomic 1.67, diatomic 1.40, polyatomic 1.33
- Mean free path: $\lambda = 1/(\sqrt{2} \pi d^2 n)$; $\lambda \propto T/P$; $\approx 10^{-7} \text{ m}$ for air at NTP
- Avogadro's number $N_a = 6.022 \times 10^{23} \text{ mol}^{-1}$; 1 mole of gas at STP = 22.4 L

Next Chapter: Chapter 11 — Thermodynamics



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