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CHAPTER NOTES

States of Matter Class 11 Notes | CBSE Chemistry Chapter

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States of Matter Class 11 Notes | CBSE Chemistry Chapter

States of Matter is one of the most rewarding chapters in CBSE Class 11 Chemistry — it explains *why* a gas fills its container, why a balloon expands in the sun, and why water forms droplets. It bridges everyday observation and the physics of molecules, using a handful of beautiful gas laws given by Boyle, Charles, Avogadro and others. Master this chapter and a whole family of JEE and NEET problems on gases and liquids becomes routine.

By the end of these notes you will be able to apply every gas law, use the ideal gas equation and Dalton's law confidently, derive results from kinetic molecular theory, explain real-gas deviations using the van der Waals equation and compressibility factor, and reason about vapour pressure, surface tension and viscosity. This is an exam-complete, NCERT-accurate chapter that feeds directly into Thermodynamics, Equilibrium and Solutions.

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Key Concepts

1. Intermolecular Forces

Intermolecular forces are the forces of attraction and repulsion between interacting particles (atoms or molecules). They decide whether a substance is a gas, liquid or solid at a given temperature. They do **not** include the covalent bonds *within* a molecule.

These attractive forces are collectively called **van der Waals forces**, named after the scientist who first explained real-gas behaviour.

Types of Intermolecular Forces

- **Dispersion (London) forces:** weak forces between instantaneous and induced dipoles; present in all atoms and molecules, even non-polar ones like Cl_2 and noble gases.
- **Dipole–dipole forces:** act between molecules having permanent dipoles, e.g. HCl . Stronger than dispersion forces.

- **Dipole–induced dipole forces:** a permanent dipole induces a dipole in a nearby non-polar molecule.
- **Hydrogen bonding:** a strong attraction when H is bonded to highly electronegative N, O or F (e.g. in H_2O , HF , NH_3). It explains the high boiling point of water.

Key idea: London force energy is proportional to $1/r^6$, so it falls off very rapidly with distance.

2. Thermal Energy

Thermal energy is the energy of a body arising from the motion of its particles. It is directly proportional to the absolute temperature and is a measure of the average kinetic energy of the particles.

The physical state of matter is a balance between two opposing tendencies: **intermolecular forces** (which try to keep particles together) and **thermal energy** (which tends to keep them apart and moving).

- Intermolecular forces dominate $\rightarrow$ solid.
 - Thermal energy dominates $\rightarrow$ gas.
 - The two are comparable $\rightarrow$ liquid.
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3. The Gas Laws

The behaviour of gases is summarised by simple relationships between pressure (p), volume (V), temperature (T) and amount (n).

Boyle's Law (pressure–volume relationship)

At constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure.

$p \propto 1/V$, so $pV = \text{constant}$, and $p_1V_1 = p_2V_2$

[DIAGRAM: A p – V graph showing a rectangular hyperbola at constant T (isotherm); higher T curves lie farther from the origin.]

Charles's Law (temperature–volume relationship)

At constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature.

$V \propto T$, so $V/T = \text{constant}$, and $V_1/T_1 = V_2/T_2$

This law leads to the concept of **absolute zero**: at -273.15°C (0 K), the volume of an ideal gas would theoretically become zero.

Gay-Lussac's Law (pressure–temperature relationship)

At constant volume, the pressure of a fixed mass of gas is directly proportional to its absolute temperature.

$p \propto T$, so $p/T = \text{constant}$, and $p_1/T_1 = p_2/T_2$

Avogadro's Law (volume–amount relationship)

At the same temperature and pressure, equal volumes of all gases contain equal numbers of molecules.

$V \propto n$ at constant T and p . One mole of any gas at STP (273.15 K, 1 bar) occupies **22.711 L** (≈ 22.4 L at 1 atm).

4. Ideal Gas Equation

Combining Boyle's, Charles's and Avogadro's laws gives the **ideal gas equation**, the central equation of this chapter.

$$pV = nRT$$

- **R** is the universal gas constant = $8.314 \text{ J K}^{-1} \text{ mol}^{-1} = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}$.
- It is an **equation of state** — it links all four variables for a given amount of gas.
- Since $n = \text{mass}/M$ and density $d = \text{mass}/V$, it can be written as **$pM = dRT$** , useful for finding molar mass.

For a fixed mass of gas, combining the laws gives the **combined gas law**: $p_1V_1/T_1 = p_2V_2/T_2$.

5. Dalton's Law of Partial Pressures

At constant temperature, the total pressure exerted by a mixture of two or more non-reacting gases is the sum of the partial pressures of the individual gases.

$$p_{\text{total}} = p_1 + p_2 + p_3 + \dots$$

- The **partial pressure** of a gas equals its mole fraction times the total pressure: **$p_1 = x_1 \times p_{\text{total}}$** .
 - When a gas is collected over water, **$p_{\text{dry gas}} = p_{\text{total}} - \text{aqueous tension}$** (the vapour pressure of water).
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6. Kinetic Molecular Theory of Gases

The **kinetic molecular theory (KMT)** explains the gas laws in terms of the motion of molecules.

Main Postulates

- Gases consist of a large number of tiny particles; their actual volume is negligible compared with the volume of the container.
- There is no force of attraction between the particles of an ideal gas.
- Particles are in constant, random motion, colliding with each other and the walls.
- Collisions are perfectly **elastic** — total kinetic energy is conserved.
- The average kinetic energy of the particles is directly proportional to the absolute temperature.

The theory gives the key result **$pV = \frac{1}{3} mN u^2(\text{rms})$** , and the average kinetic energy of one mole is **$(3/2)RT$** .

7. Maxwell Distribution of Molecular Speeds

At any temperature, molecules in a gas do not all move at the same speed. The **Maxwell–Boltzmann distribution** shows the fraction of molecules having each speed.

[DIAGRAM: A Maxwell speed-distribution curve — fraction of molecules on the y-axis versus speed on the x-axis; as temperature rises the peak shifts right and flattens.]

Three Important Speeds

- **Most probable speed (u_{mp}):** the speed possessed by the largest number of molecules = $\sqrt{(2RT/M)}$.
- **Average speed (u_{av}):** = $\sqrt{(8RT/\pi M)}$.
- **Root mean square speed (u_{rms}):** = $\sqrt{(3RT/M)}$.

Their ratio is fixed: $u_{mp} : u_{av} : u_{rms} = 1 : 1.128 : 1.224$. The rms speed is the largest of the three.

8. Behaviour of Real Gases — Deviation from Ideal Behaviour

Real gases obey the ideal gas equation only at **low pressure** and **high temperature**. At high pressure and low temperature they deviate because two ideal-gas assumptions break down.

- **Molecular volume is not negligible** at high pressure — molecules occupy real space.
 - **Intermolecular attractions exist** — they reduce the pressure a real gas exerts on the walls.
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9. van der Waals Equation

To correct the ideal gas equation for real gases, van der Waals introduced two correction terms — one for attractive forces (a) and one for molecular volume (b).

$$(p + \frac{an^2}{V^2})(V - nb) = nRT$$

- **a** is a measure of the strength of intermolecular attraction; larger a means easier liquefaction.
 - **b** (the excluded volume) is a measure of the actual size of the molecules.
 - Units: a in $\text{atm L}^2 \text{mol}^{-2}$; b in L mol^{-1} .
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10. Compressibility Factor (Z)

The **compressibility factor** measures how far a real gas departs from ideal behaviour.

$$Z = \frac{pV}{nRT}$$

- For an **ideal gas**, **Z = 1** at all temperatures and pressures.
- **Z < 1** (at moderate pressure): attractive forces dominate; the gas is more compressible than ideal.

- **$Z > 1$** (at high pressure): molecular volume dominates; the gas is less compressible than ideal.

The temperature at which a real gas behaves ideally over a range of pressure is its **Boyle temperature**.

11. Liquefaction of Gases and Critical Constants

A gas can be liquefied by lowering temperature and increasing pressure — but only below a certain temperature called the **critical temperature**.

- **Critical temperature (T_c)**: the highest temperature at which a gas can be liquefied by pressure alone. Above T_c no amount of pressure can liquefy it.
- **Critical pressure (p_c)**: the pressure required to liquefy a gas at its critical temperature.
- **Critical volume (V_c)**: the molar volume at the critical point.

At and above T_c the substance exists as a **supercritical fluid**, with no distinct gas–liquid boundary. The critical constants relate to van der Waals constants: $T_c = 8a/27Rb$ and $p_c = a/27b^2$.

12. The Liquid State

In liquids, intermolecular forces are strong enough to hold molecules together but not to fix them in place, so liquids have a definite volume but no definite shape.

Vapour Pressure

The **vapour pressure** of a liquid is the pressure exerted by its vapour in equilibrium with the liquid at a given temperature. It increases with temperature. A liquid **boils** when its vapour pressure equals the external (atmospheric) pressure.

Surface Tension

Surface tension is the force acting per unit length along the surface of a liquid, arising because surface molecules are pulled inward. It is why drops are spherical (minimum surface area) and why water rises in a capillary. SI unit: N m^{-1} . It **decreases** with rise in temperature.

Viscosity

Viscosity is the resistance to flow of a liquid, caused by internal friction between layers moving at different speeds. SI unit: N s m^{-2} (poiseuille). Viscosity **decreases** with increasing temperature, which is why honey flows faster when warmed.

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	4–6 marks	Gas laws, ideal gas equation, Dalton's law, liquid-state properties
JEE Main / Advanced	1–2 questions	van der Waals equation, compressibility factor Z , KMT, molecular speeds
NEET	1–2 questions	Gas laws, ideal gas numericals, real-gas deviation, critical constants

[TABLE: Question-type split — VSA (1 mark): definitions & laws; SA (2–3 marks): gas-law numericals, Dalton's law, surface tension/viscosity; LA (5 marks): KMT postulates, van der Waals equation, deviations from ideal behaviour.]

Important Definitions

Term	Definition
Boyle's law	At constant T , $p \propto 1/V$; $pV = \text{constant}$
Charles's law	At constant p , $V \propto T$; $V/T = \text{constant}$
Gay-Lussac's law	At constant V , $p \propto T$; $p/T = \text{constant}$
Avogadro's law	Equal volumes of gases at same T and p contain equal numbers of molecules
Ideal gas equation	$pV = nRT$ — the equation of state of an ideal gas
Dalton's law	Total pressure = sum of partial pressures: $p = p_1 + p_2 + \dots$
Compressibility factor	$Z = pV/nRT$; $Z = 1$ for an ideal gas
Critical temperature	Highest temperature at which a gas can be liquefied by pressure alone
Vapour pressure	Pressure of vapour in equilibrium with its liquid at a given temperature
Surface tension	Force per unit length acting along a liquid surface (N m^{-1})
Viscosity	Resistance to flow due to internal friction between liquid layers

Solved Examples

Example 1

A gas occupies 2 L at 1 atm. What volume will it occupy at 4 atm, at constant temperature?

Answer: By Boyle's law, $p_1V_1 = p_2V_2$. $V_2 = (1 \times 2)/4 = \mathbf{0.5\text{ L}}$.

Example 2

A gas has a volume of 300 mL at 27 °C. Find its volume at 127 °C at constant pressure.

Answer: $T_1 = 300\text{ K}$, $T_2 = 400\text{ K}$. By Charles's law, $V_2 = V_1T_2/T_1 = 300 \times 400/300 = \mathbf{400\text{ mL}}$.

Example 3

Calculate the pressure exerted by 2 mol of an ideal gas in a 10 L container at 300 K. ($R = 0.0821\text{ L atm K}^{-1}\text{ mol}^{-1}$)

Answer: $p = nRT/V = (2 \times 0.0821 \times 300)/10 = 49.26/10 = \mathbf{4.93\text{ atm}}$.

Example 4

A mixture contains 4 g of O_2 and 2 g of H_2 at a total pressure of 3 atm. Find the partial pressure of H_2 .

Answer: Moles: $\text{O}_2 = 4/32 = 0.125$; $\text{H}_2 = 2/2 = 1$. Total = 1.125. $x(\text{H}_2) = 1/1.125 = 0.889$. $p(\text{H}_2) = 0.889 \times 3 = \mathbf{2.67\text{ atm}}$.

Example 5

Calculate the root mean square speed of oxygen molecules at 300 K. ($R = 8.314\text{ J K}^{-1}\text{ mol}^{-1}$, $M = 32 \times 10^{-3}\text{ kg mol}^{-1}$)

Answer: $u_{\text{rms}} = \sqrt{(3RT/M)} = \sqrt{(3 \times 8.314 \times 300 / 0.032)} = \sqrt{(233\ 831)} \approx \mathbf{483.6\text{ m/s}}$.

Example 6

The compressibility factor of a real gas at a certain p and T is 0.8. Is the gas more or less compressible than an ideal gas, and which effect dominates?

Answer: Since $Z = 0.8 < 1$, the gas is **more compressible than ideal**; **attractive intermolecular forces dominate** at this pressure.

Important Questions for Board Exams

1-Mark Questions (VSA)

1. State Boyle's law.

2. What is the value of the compressibility factor Z for an ideal gas?
3. Why are liquid drops spherical in shape?
4. Define critical temperature.
5. Why does the viscosity of a liquid decrease with rise in temperature?

2–3-Mark Questions (SA)

1. Derive the combined gas law from Boyle's and Charles's laws.
2. State Dalton's law of partial pressures and express the partial pressure in terms of mole fraction.
3. Distinguish between average speed, most probable speed and root mean square speed, and give their ratio.
4. Explain why real gases deviate from ideal behaviour at high pressure and low temperature.

5-Mark Questions (LA)

1. State the postulates of the kinetic molecular theory of gases and explain how it accounts for gas pressure.
 2. Write the van der Waals equation, explain the significance of the constants a and b , and how it corrects the ideal gas equation.
 3. Explain vapour pressure, surface tension and viscosity, and describe how each varies with temperature.
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Quick Revision Points

- Intermolecular forces (van der Waals): dispersion < dipole–dipole < hydrogen bonding
 - State of matter = balance of intermolecular forces vs thermal energy
 - Boyle: $pV = \text{constant}$; Charles: $V/T = \text{constant}$; Gay-Lussac: $p/T = \text{constant}$; Avogadro: $V \propto n$
 - Ideal gas equation: $pV = nRT$; $R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1} = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$
 - Dalton's law: $p_{\text{total}} = p_1 + p_2 + \dots$; partial pressure = mole fraction $\times$ total pressure
 - KMT: elastic collisions, no attraction, $KE \propto T$; average KE per mole = $(3/2)RT$
 - Speeds: $u_{\text{mp}} : u_{\text{av}} : u_{\text{rms}} = 1 : 1.128 : 1.224$; $u_{\text{rms}} = \sqrt{(3RT/M)}$
 - Real gases ideal at low p , high T ; corrected by van der Waals: $(p + an^2/V^2)(V - nb) = nRT$
 - Compressibility factor $Z = pV/nRT$; $Z = 1$ (ideal), $Z < 1$ (attraction), $Z > 1$ (volume)
 - Critical temperature T_c : above it a gas cannot be liquefied by pressure alone
 - Surface tension and vapour pressure of a liquid; viscosity decreases as temperature rises
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Next Chapter: Chapter 6 — Thermodynamics



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