



MATH LOVE INSTITUTE

Class 11 Chemistry - Chapter 2



STRUCTURE OF ATOM

Session: 2025-26 | CBSE Curriculum

Complete Study Material with Theory, Numerical, MCQs & Case Studies

Indore, Madhya Pradesh

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CHAPTER INTRODUCTION

Why This Chapter Matters:

The **Structure of Atom** is one of the most important chapters in Class 11 Chemistry! This chapter explains:

- ✓ How atoms are built (electrons, protons, neutrons)
- ✓ How electrons are arranged in atoms (electronic configuration)
- ✓ Why elements have similar properties (periodic table connection)
- ✓ How light and matter interact (quantum mechanics basics)
- ✓ Foundation for Chemical Bonding, Periodic Table, and all organic chemistry!

Board Exam Weightage: 6-8 marks | JEE: 3-4 questions | NEET: 2-3 questions

Learning Strategy for This Chapter:

1. **History First:** Understand how atomic models evolved (Thomson → Rutherford → Bohr → Quantum)
2. **Master Bohr's Model:** This is the most important - practice radius, energy, wavelength numericals
3. **Memorize Electronic Configurations:** First 36 elements by heart!
4. **Practice Quantum Numbers:** n , l , m , s - these are asked every year
5. **Solve Previous Year Questions:** Pattern repeats in boards!



DISCOVERY OF SUBATOMIC PARTICLES

Historical Timeline:

Year	Scientist	Discovery
1897	J.J. Thomson	Electron (cathode ray experiment)
1886	E. Goldstein	Proton (canal rays/anode rays)
1932	James Chadwick	Neutron

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1 Discovery of Electron (J.J. Thomson - 1897)

Cathode Ray Experiment:

Setup: Discharge tube with very low pressure ($\sim 10^{-2}$ mm Hg) and high voltage ($\sim 10,000$ V)

Observation: Rays travelled from cathode to anode, causing fluorescence on the tube wall

Properties of Cathode Rays (Electrons):

- ✓ Travel in straight lines
- ✓ Deflected by electric and magnetic fields (negatively charged)
- ✓ Produce mechanical effects (can rotate a paddle wheel)
- ✓ Produce fluorescence on zinc sulfide screen
- ✓ Same properties regardless of gas or cathode material
- ✓ Negatively charged particles called **electrons**

e/m Ratio of Electron:

$$e/m = 1.758820 \times 10^{11} \text{ C/kg}$$

Determined by Thomson using deflection in electric and magnetic fields

Millikan's Oil Drop Experiment (1906-1914):

What he did: Measured the charge on electron by observing oil drops between charged plates

Result: Charge on electron = $-1.6 \times 10^{-19} \text{ C}$

Using e/m ratio:

Mass of electron = $e/(e/m) = (1.6 \times 10^{-19})/(1.758820 \times 10^{11})$

Mass of electron = $9.1094 \times 10^{-31} \text{ kg}$

2 **Discovery of Proton (E. Goldstein - 1886)**

Canal Ray Experiment (Anode Rays):

Setup: Discharge tube with perforated cathode

Observation: Rays travelling in opposite direction to cathode rays, passing through holes in cathode

Properties of Canal Rays (Protons):

- ✓ Positively charged particles
- ✓ Deflected by electric and magnetic fields (opposite to electrons)
- ✓ e/m ratio depends on gas taken in tube
- ✓ Charge and mass depend on gas
- ✓ **Lightest and smallest: H^+ ion (proton)**

Mass of proton = $1.672622 \times 10^{-27} \text{ kg}$ = 1836 times mass of electron

Charge on proton = $+1.6 \times 10^{-19} \text{ C}$

3 Discovery of Neutron (James Chadwick - 1932)

Neutron Discovery:

What happened: Bombarded beryllium with α -particles, detected electrically neutral particles

Mass of neutron = 1.674927×10^{-27} kg (slightly heavier than proton)

Charge = 0 (neutral)

Present in nucleus with protons

Quick Comparison Table:

Particle	Symbol	Charge	Mass (kg)	Mass (u)	Location
Electron	e^-	-1.6×10^{-19} C	9.1094×10^{-31}	0.000549	Outside nucleus
Proton	p^+	$+1.6 \times 10^{-19}$ C	1.6726×10^{-27}	1.00727	Inside nucleus
Neutron	n	0	1.6749×10^{-27}	1.00866	Inside nucleus

 **Common Mistakes to Avoid:**

- **✗** Saying cathode rays are positive (they are NEGATIVE - electrons!)
- **✗** Confusing canal rays with cathode rays
- **✗** Writing neutron mass = proton mass (neutron is slightly heavier!)
- **✗** Forgetting that e/m ratio of electron is CONSTANT, but for proton it varies with gas

ATOMIC MODELS

1 Thomson's Model (1898) - "Plum Pudding Model"

Description:

- Atom is a sphere of positive charge ($\sim 10^{-10}$ m radius)
- Electrons embedded like plums in a pudding
- Positive and negative charges equal \rightarrow atom is neutral
- Also called "**Raisin Pudding Model**" or "**Watermelon Model**"

Limitations:

- Could not explain α -particle scattering experiment results
- Could not explain atomic spectra
- Could not explain stability of atom

2 Rutherford's Model (1911) - "Nuclear Model"

α -Particle Scattering Experiment (Gold Foil Experiment):

Setup: Bombarded thin gold foil (~ 0.0004 cm thick) with α -particles (He^{2+} ions)

Observations:

1. **Most α -particles passed straight through** (99.9%)
2. **Some deflected at small angles** ($\sim 0.1\%$)
3. **Very few deflected at large angles** (1 in 20,000)
4. **Some bounced back** (1 in 8000)

💡 Conclusions from Experiment:

1. **Most of atom is empty space** (most α -particles passed through)
2. **Positive charge concentrated in center** called **NUCLEUS**
3. **Nucleus is very small** compared to atom size
4. **Nucleus contains most of atom's mass**
5. **Electrons revolve around nucleus** in circular orbits like planets

Key Facts:

- ✓ Nucleus radius $\sim 10^{-15}$ m
- ✓ Atom radius $\sim 10^{-10}$ m
- ✓ **Ratio = 10^5** (nucleus is 100,000 times smaller!)
- ✓ If atom = football stadium, nucleus = small marble!

✗ Limitations of Rutherford's Model:

1. **Stability Problem:** According to classical physics, revolving electron should continuously emit energy and spiral into nucleus. Atom should collapse in 10^{-8} seconds!
2. **Spectral Lines:** Could not explain why atoms emit light of specific wavelengths only (line spectra)

ELECTROMAGNETIC RADIATION & ATOMIC SPECTRA

What is Electromagnetic Radiation?

Properties of EM Radiation:

- ✓ Form of energy propagating through space
- ✓ No medium required (travels through vacuum)
- ✓ Has wave and particle nature (dual nature!)
- ✓ Travels at speed of light: $c = 3 \times 10^8 \text{ m/s}$

Important Relationships:

$$c = v\lambda$$

where c = speed of light, v (nu) = frequency, λ (lambda) = wavelength

$$\text{Wave number } (\bar{\nu}) = 1/\lambda$$

Unit: cm^{-1} or m^{-1}

$$\text{Energy: } E = hv = hc/\lambda$$

where h = Planck's constant = $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$

Electromagnetic Spectrum (Remember Order!):

Type	Wavelength Range	Frequency Range
Radio waves	> 0.1 m	$< 3 \times 10^9$ Hz
Microwaves	0.1 m to 1 mm	3×10^9 to 3×10^{11} Hz
Infrared	1 mm to 700 nm	3×10^{11} to 4.3×10^{14} Hz
Visible light	700 nm to 400 nm	4.3×10^{14} to 7.5×10^{14} Hz
Ultraviolet	400 nm to 1 nm	7.5×10^{14} to 3×10^{17} Hz
X-rays	1 nm to 10^{-3} nm	3×10^{17} to 3×10^{20} Hz
Gamma rays	$< 10^{-3}$ nm	$> 3 \times 10^{20}$ Hz

Memory Trick: "Rude Men In Very Ugly eXtremely Gross" (order of increasing frequency/energy)

Atomic Spectra

Types of Spectra:

1. Continuous Spectrum:

- All wavelengths present (no gaps)
- Example: White light, rainbow
- Produced by hot solids/liquids

2. Line Spectrum (Atomic Spectrum):

- Only specific wavelengths present
- Appears as bright lines on dark background
- Each element has unique line spectrum (fingerprint!)
- Produced by excited atoms/molecules in gas phase

Types of Line Spectra:

- **Emission Spectrum:** Bright lines on dark background (energy released)
- **Absorption Spectrum:** Dark lines on bright background (energy absorbed)

Hydrogen Spectrum - THE MOST IMPORTANT!

Why Hydrogen?

- ✓ Simplest atom (1 proton + 1 electron)
- ✓ Line spectrum can be mathematically predicted
- ✓ Led to development of Bohr's model

Rydberg Formula for Hydrogen Spectrum:

$$1/\lambda = R_H (1/n_1^2 - 1/n_2^2)$$

where:

- R_H = Rydberg constant = $1.097 \times 10^7 \text{ m}^{-1}$ or $109,677 \text{ cm}^{-1}$
- n_1 = lower energy level
- n_2 = higher energy level ($n_2 > n_1$)
- λ = wavelength of emitted/absorbed light

Spectral Series of Hydrogen (Must Memorize!):

Series Name	n_1	n_2	Region	First Line (nm)
Lyman	1	2,3,4...	UV	121.5
Balmer	2	3,4,5...	Visible	656.3 (red)
Paschen	3	4,5,6...	Near IR	1875
Brackett	4	5,6,7...	IR	4050
Pfund	5	6,7,8...	Far IR	7400

Memory Trick: "Little Boys Play Baseball Poorly" (Lyman, Balmer, Paschen, Brackett, Pfund)

 **Example: Balmer Series**

Question: Calculate wavelength of H α line (first line) in Balmer series.

Solution:

For Balmer series: $n_1 = 2$, $n_2 = 3$ (first line)

$$1/\lambda = R_H (1/2^2 - 1/3^2)$$

$$1/\lambda = 1.097 \times 10^7 (1/4 - 1/9)$$

$$1/\lambda = 1.097 \times 10^7 (9-4)/36 = 1.097 \times 10^7 \times 5/36$$

$$1/\lambda = 1.524 \times 10^6 \text{ m}^{-1}$$

$$\lambda = 656.3 \text{ nm (RED line - visible!)}$$



BOHR'S MODEL OF ATOM (1913) - MOST IMPORTANT!



Niels Bohr (1913):

Combined quantum theory with Rutherford's model to explain:

- ✓ Stability of atom
- ✓ Line spectrum of hydrogen



Bohr's Postulates (Learn Word-to-Word!):

Postulate 1 (Stationary Orbits):

Electrons revolve around nucleus in certain fixed circular paths called **orbits** or **stationary states**. In these orbits, electrons do NOT radiate energy.

Postulate 2 (Quantization of Angular Momentum):

Angular momentum of electron is quantized:

$$mvr = nh/2\pi$$

where $n = 1, 2, 3, \dots$ (principal quantum number)

Postulate 3 (Energy Absorption/Emission):

When electron jumps from higher energy orbit (E_2) to lower energy orbit (E_1), energy is emitted as photon:

$$\Delta E = E_2 - E_1 = hv$$

When electron jumps from lower to higher orbit, energy is absorbed.

▢ Important Formulas from Bohr's Model:

1. Radius of nth orbit:

$$r_n = 0.529 \times n^2/Z \text{ \AA}$$

For hydrogen (Z=1): $r_1 = 0.529 \text{ \AA} = 52.9 \text{ pm}$ (Bohr radius)

$$r_2 = 2.12 \text{ \AA}, r_3 = 4.77 \text{ \AA}, r_4 = 8.48 \text{ \AA}$$

2. Velocity of electron in nth orbit:

$$v_n = 2.18 \times 10^6 \times Z/n \text{ m/s}$$

$$\text{For H: } v_1 = 2.18 \times 10^6 \text{ m/s}$$

3. Energy of electron in nth orbit:

$$E_n = -13.6 \times Z^2/n^2 \text{ eV}$$

For H: $E_1 = -13.6 \text{ eV}$, $E_2 = -3.4 \text{ eV}$, $E_3 = -1.51 \text{ eV}$, $E_\infty = 0$

4. Energy of transition ($n_2 \rightarrow n_1$):

$$\Delta E = 13.6 \times Z^2 (1/n_1^2 - 1/n_2^2) \text{ eV}$$

5. Frequency of radiation:

$$\nu = \Delta E/h = (13.6 \times 1.6 \times 10^{-19} \times Z^2/h) (1/n_1^2 - 1/n_2^2)$$

6. Wavelength of radiation:

$$1/\lambda = R_H \times Z^2 (1/n_1^2 - 1/n_2^2)$$

Quick Facts to Remember:

- ✓ Bohr radius (a_0) = 0.529 Å = 52.9 pm
- ✓ Ground state energy of H = -13.6 eV
- ✓ Ionization energy of H = +13.6 eV
- ✓ Energy is NEGATIVE (electron is bound to nucleus)
- ✓ Energy increases as n increases (becomes less negative)
- ✓ At $n = \infty$, $E = 0$ (electron is free/ionized)
- ✓ Radius $\propto n^2$, Velocity $\propto 1/n$, Energy $\propto 1/n^2$

Example 1: Calculate radius of 3rd orbit of H

Solution:

$$r_n = 0.529 \times n^2/Z \text{ \AA}$$

$$r_3 = 0.529 \times 9/1 = 4.761 \text{ \AA}$$

Answer: 4.76 Å

 **Example 2: Calculate energy released when electron jumps from $n=3$ to $n=2$ in H**

Solution:

$$E_3 = -13.6/9 = -1.51 \text{ eV}$$






$$E_2 = -13.6/4 = -3.4 \text{ eV}$$

$$\Delta E = E_3 - E_2 = -1.51 - (-3.4) = 1.89 \text{ eV}$$

Since electron goes down, energy is **released**

Answer: 1.89 eV emitted

 **Common Mistakes in Bohr's Model:**

-  Forgetting negative sign in energy formula
-  Using wrong n values in transition (higher to lower for emission!)
-  Confusing radius formula (n^2 not n !)
-  Not converting units (eV to J, nm to m)
-  Writing $E_1 = +13.6 \text{ eV}$ (it's NEGATIVE!)

✗ Limitations of Bohr's Model:

1. Works only for hydrogen and hydrogen-like species (He^+ , Li^{2+} , etc.)
2. Cannot explain spectra of multi-electron atoms
3. Cannot explain fine structure of spectral lines
4. Cannot explain splitting of spectral lines in magnetic field (Zeeman effect) or electric field (Stark effect)
5. Does not explain chemical bonding
6. Treats electrons as particles only (ignores wave nature)

DUAL NATURE OF MATTER - DE BROGLIE'S THEORY

 **Louis de Broglie (1924):**

If light has both wave and particle nature, then matter should also have wave nature!

 **de Broglie Wavelength:**

$$\lambda = h/mv = h/p$$

where:

- λ = wavelength of particle
- h = Planck's constant = 6.626×10^{-34} J·s
- m = mass of particle (kg)
- v = velocity of particle (m/s)
- p = momentum = mv

Key Points:

- ✓ ALL matter has wave nature (even you!)
- ✓ Wave nature significant only for very small particles (electrons, protons)
- ✓ For macroscopic objects, wavelength is too small to observe
- ✓ Confirmed by Davisson-Germer experiment (electron diffraction)

Example: Calculate de Broglie wavelength of electron moving at 10^6 m/s

Given: $v = 10^6$ m/s, $m = 9.1 \times 10^{-31}$ kg, $h = 6.626 \times 10^{-34}$ J·s

Solution:

$$\lambda = h/mv = (6.626 \times 10^{-34}) / (9.1 \times 10^{-31} \times 10^6)$$

$$\lambda = 7.28 \times 10^{-10} \text{ m} = 0.728 \text{ nm}$$

Answer: 0.728 nm

? HEISENBERG'S UNCERTAINTY PRINCIPLE

The Principle:

$$\Delta x \cdot \Delta p \geq h/4\pi$$

or

$$\Delta x \cdot \Delta(mv) \geq h/4\pi$$

where:

- Δx = uncertainty in position
- Δp = uncertainty in momentum
- h = Planck's constant

Meaning:

It is **IMPOSSIBLE** to simultaneously determine exact position and exact momentum of a particle (like electron).

If we know position precisely (small Δx), momentum becomes uncertain (large Δp), and vice versa.

Consequences:

- ✓ Concept of fixed orbits in Bohr's model is wrong!
- ✓ We can only talk about probability of finding electron
- ✓ Led to development of quantum mechanics

🌀 QUANTUM MECHANICAL MODEL OF ATOM

🎯 Key Concepts:

- ✓ Developed by Schrödinger (1926)
- ✓ Treats electron as a wave (wave function ψ)
- ✓ ψ^2 gives probability of finding electron at a point
- ✓ Electrons exist in **orbitals** not orbits!
- ✓ **Orbital**: 3D region where probability of finding electron is maximum (90-95%)

1 2 3 4 QUANTUM NUMBERS

Four Quantum Numbers (Must Know!)

1 Principal Quantum Number (n):

- ✓ Denotes main energy level/shell
- ✓ Values: $n = 1, 2, 3, 4...$ (positive integers)
- ✓ Shells: K, L, M, N... (for $n = 1, 2, 3, 4...$)
- ✓ Max electrons in shell = $2n^2$
- ✓ Determines size of orbital (larger $n =$ larger orbital)

2 Azimuthal/Angular Momentum Quantum Number (l):

- ✓ Denotes subshell/sublevel and shape of orbital
- ✓ Values: $l = 0$ to $(n-1)$
- ✓ For $n=1$: $l=0$ (only s subshell)
- ✓ For $n=2$: $l=0, 1$ (s, p subshells)
- ✓ For $n=3$: $l=0, 1, 2$ (s, p, d subshells)
- ✓ For $n=4$: $l=0, 1, 2, 3$ (s, p, d, f subshells)

l value	0	1	2	3
Subshell	s	p	d	f
Shape	Spherical	Dumbbell	Double dumbbell	Complex
Max electrons	2	6	10	14

3 Magnetic Quantum Number (m or m_l):

- ✓ Denotes orientation of orbital in space
- ✓ Values: $m = -l$ to $+l$ (including 0)
- ✓ Total values = $2l + 1$

- ✓ For s ($l=0$): $m=0$ (1 orbital)
- ✓ For p ($l=1$): $m=-1,0,+1$ (3 orbitals: p_x, p_y, p_z)
- ✓ For d ($l=2$): $m=-2,-1,0,+1,+2$ (5 orbitals)
- ✓ For f ($l=3$): $m=-3,-2,-1,0,+1,+2,+3$ (7 orbitals)

4 Spin Quantum Number (s or m_s):

- ✓ Denotes spin orientation of electron
- ✓ Values: **+1/2** (clockwise \uparrow) or **-1/2** (anticlockwise \downarrow)
- ✓ Each orbital can hold MAX 2 electrons with opposite spins

Quick Reference Table:

Shell (n)	Subshells	Orbitals	Max Electrons
1 (K)	1s	1	2
2 (L)	2s, 2p	$1+3 = 4$	8
3 (M)	3s, 3p, 3d	$1+3+5 = 9$	18
4 (N)	4s, 4p, 4d, 4f	$1+3+5+7 = 16$	32

 **Example: Write all quantum numbers for 2p electron**

Solution:






For 2p: $n = 2$, $l = 1$ (since p subshell)

$m = -1, 0, +1$ (3 possible values)

$s = +1/2$ or $-1/2$

Total combinations: 3 orbitals \times 2 spins = 6 electrons possible

 **Common Mistakes with Quantum Numbers:**

-  Writing $n = 0$ (minimum is $n = 1$!)
-  Writing $l = n$ (should be $l = n-1$ maximum!)
-  Confusing l values: $s=0$, $p=1$, $d=2$, $f=3$
-  Writing wrong number of orbitals (remember: $s=1$, $p=3$, $d=5$, $f=7$)
-  Forgetting that each orbital holds 2 electrons max

SHAPES OF ORBITALS

s-Orbital ($l = 0$):

- ✓ **Spherical shape**
- ✓ Non-directional (same in all directions)
- ✓ Size increases: $1s < 2s < 3s < 4s$
- ✓ 1s has NO nodes, 2s has 1 node, 3s has 2 nodes
- ✓ Node = region where probability of finding electron is ZERO

p-Orbital ($l = 1$):

- ✓ **Dumbbell shape** (two lobes)
- ✓ Directional (oriented along axes)
- ✓ Three types: p_x (along x-axis), p_y (along y-axis), p_z (along z-axis)
- ✓ Node at nucleus (nodal plane)
- ✓ Starts from $n = 2$ onwards

d-Orbital ($l = 2$):

- ✓ **Double dumbbell shape** (4 lobes)
- ✓ Five types: d_{xy} , d_{yz} , d_{xz} , $d_{x^2-y^2}$, d_{z^2}
- ✓ Four are cloverleaf shaped, one (d_{z^2}) is different
- ✓ Starts from $n = 3$ onwards



ELECTRONIC CONFIGURATION

🎯 Rules for Filling Electrons:

1 Aufbau Principle:

Electrons fill orbitals starting from **lowest energy** to highest energy

Order: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d$

2 Pauli Exclusion Principle:

No two electrons in an atom can have the same set of all four quantum numbers

Result: Maximum 2 electrons per orbital with opposite spins

3 Hund's Rule of Maximum Multiplicity:

Electrons first singly occupy all orbitals of same energy (degenerate orbitals) with parallel spins before pairing starts

Example: In p^3 : $\uparrow \uparrow \uparrow$ (correct), $\uparrow \downarrow \uparrow _$ (wrong!)

$\begin{matrix} 12 \\ 34 \end{matrix}$ (n+l) Rule for Orbital Energy:

Lower (n+l) value = Lower energy = Fills first

If (n+l) same, lower n fills first

Example: 3d vs 4s

3d: $n=3, l=2 \rightarrow n+l=5$

4s: $n=4, l=0 \rightarrow n+l=4$

4s fills before 3d!

💡 Memory Trick for Filling Order:

Draw diagonal arrows:

1s

2s 2p

3s 3p 3d

4s 4p 4d 4f

5s 5p 5d 5f

6s 6p 6d

7s 7p

Follow arrows diagonally from top-right to bottom-left!

 **Electronic Configurations (First 36 Elements -**

MEMORIZE!):

Element	Z	Configuration
H	1	$1s^1$
He	2	$1s^2$
Li	3	$1s^2 2s^1$ or $[\text{He}] 2s^1$
Be	4	$[\text{He}] 2s^2$
B	5	$[\text{He}] 2s^2 2p^1$
C	6	$[\text{He}] 2s^2 2p^2$
N	7	$[\text{He}] 2s^2 2p^3$
O	8	$[\text{He}] 2s^2 2p^4$
F	9	$[\text{He}] 2s^2 2p^5$
Ne	10	$[\text{He}] 2s^2 2p^6$ or $1s^2 2s^2 2p^6$
Na	11	$[\text{Ne}] 3s^1$
Mg	12	$[\text{Ne}] 3s^2$
Al	13	$[\text{Ne}] 3s^2 3p^1$
Si	14	$[\text{Ne}] 3s^2 3p^2$
P	15	$[\text{Ne}] 3s^2 3p^3$
S	16	$[\text{Ne}] 3s^2 3p^4$
Cl	17	$[\text{Ne}] 3s^2 3p^5$

Ar	18	[Ne] 3s ² 3p ⁶
K	19	[Ar] 4s ¹
Ca	20	[Ar] 4s ²
Sc	21	[Ar] 3d ¹ 4s ²
Ti	22	[Ar] 3d ² 4s ²
V	23	[Ar] 3d ³ 4s ²
Cr ⚠	24	[Ar] 3d ⁵ 4s ¹ (Exception!)
Mn	25	[Ar] 3d ⁵ 4s ²
Fe	26	[Ar] 3d ⁶ 4s ²
Co	27	[Ar] 3d ⁷ 4s ²
Ni	28	[Ar] 3d ⁸ 4s ²
Cu ⚠	29	[Ar] 3d ¹⁰ 4s ¹ (Exception!)
Zn	30	[Ar] 3d ¹⁰ 4s ²
Ga	31	[Ar] 3d ¹⁰ 4s ² 4p ¹
Ge	32	[Ar] 3d ¹⁰ 4s ² 4p ²
As	33	[Ar] 3d ¹⁰ 4s ² 4p ³
Se	34	[Ar] 3d ¹⁰ 4s ² 4p ⁴
Br	35	[Ar] 3d ¹⁰ 4s ² 4p ⁵
Kr	36	[Ar] 3d ¹⁰ 4s ² 4p ⁶

⚠ **IMPORTANT EXCEPTIONS (Learn These!):**

- **Cr (24):** $3d^5 4s^1$ NOT $3d^4 4s^2$ (half-filled d is stable!)
- **Cu (29):** $3d^{10} 4s^1$ NOT $3d^9 4s^2$ (fully-filled d is stable!)
- **Mo (42):** $4d^5 5s^1$
- **Ag (47):** $4d^{10} 5s^1$
- **Au (79):** $5d^{10} 6s^1$

Reason: Half-filled (d^5) and fully-filled (d^{10}) subshells are more stable due to exchange energy!

 **Example: Write electronic configuration of Fe (Z=26) and Fe²⁺**

Solution:

Fe (26): $[\text{Ar}] 3d^6 4s^2$

For Fe²⁺: Remove 2 electrons from **4s first** (not 3d!)

Fe²⁺ (24): $[\text{Ar}] 3d^6 4s^0$ or simply $[\text{Ar}] 3d^6$

Remember: When forming ions, remove 4s electrons before 3d!



PRACTICE PROBLEMS - NUMERICAL

Problem 1: Calculate energy and wavelength emitted when electron in H jumps from $n=4$ to $n=2$

Solution:

$$E_4 = -13.6/16 = -0.85 \text{ eV}$$

$$E_2 = -13.6/4 = -3.4 \text{ eV}$$

$$\Delta E = E_4 - E_2 = -0.85 - (-3.4) = 2.55 \text{ eV}$$

$$\text{Converting to Joules: } 2.55 \times 1.6 \times 10^{-19} = 4.08 \times 10^{-19} \text{ J}$$

$$\lambda = hc/E = (6.626 \times 10^{-34} \times 3 \times 10^8) / (4.08 \times 10^{-19})$$

$$\lambda = 487 \text{ nm (Blue-green line in Balmer series!)}$$

Problem 2: Calculate velocity of electron in 1st Bohr orbit of H

Solution:

$$v_1 = 2.18 \times 10^6 \times Z/n \text{ m/s}$$

For H: $Z=1$, $n=1$

$$v_1 = 2.18 \times 10^6 \text{ m/s}$$

$$\text{Answer: } 2.18 \times 10^6 \text{ m/s}$$

Problem 3: Calculate de Broglie wavelength of ball (m=100g) moving at 35 m/s

Solution:

$$m = 100 \text{ g} = 0.1 \text{ kg}, v = 35 \text{ m/s}$$

$$\lambda = h/mv = (6.626 \times 10^{-34}) / (0.1 \times 35)$$

$$\lambda = 1.89 \times 10^{-34} \text{ m}$$

This is SO SMALL - wave nature not observable for macroscopic objects!

Problem 4: How many electrons in an atom can have n=4, l=1?

Solution:

n=4, l=1 means 4p subshell

p subshell has 3 orbitals

Each orbital holds 2 electrons

Answer: $3 \times 2 = 6$ electrons

Problem 5: Calculate ionization energy of Li^{2+} ion

Solution:

Li^{2+} is hydrogen-like with $Z=3$, $n=1$

$$E_1 = -13.6 \times Z^2/n^2 = -13.6 \times 9/1 = -122.4 \text{ eV}$$

Ionization energy = Energy needed to remove electron from ground state

Answer: +122.4 eV

✓ MCQ PRACTICE (CBSE Pattern)

Q1. Which of the following has maximum wavelength?

- (a) Gamma rays
- (b) X-rays
- (c) Infrared
- (d) Radio waves

Q2. The energy of electron in n th orbit of hydrogen is given by:

- (a) $-13.6/n$ eV
- (b) $-13.6 \times n^2$ eV
- (c) $-13.6/n^2$ eV
- (d) $-13.6 \times n$ eV

Q3. Chromium ($Z=24$) has electronic configuration:

- (a) $[\text{Ar}] 3d^4 4s^2$
- (b) $[\text{Ar}] 3d^5 4s^1$
- (c) $[\text{Ar}] 3d^6 4s^0$
- (d) $[\text{Ar}] 3d^3 4s^3$

Q4. Maximum number of electrons in a subshell with $l=3$ is:

- (a) 2
- (b) 6
- (c) 10
- (d) 14

Q5. Which series of hydrogen spectrum lies in visible region?

- (a) Lyman
- (b) Balmer
- (c) Paschen
- (d) Brackett

Q6. de Broglie wavelength is associated with:

- (a) Only charged particles
- (b) Only uncharged particles
- (c) All material particles
- (d) Only electrons

Q7. Number of orbitals in $n=4$ shell is:

- (a) 4
- (b) 8
- (c) 16
- (d) 32

Q8. Heisenberg's uncertainty principle is significant for:

- (a) Macroscopic particles only
- (b) Microscopic particles only
- (c) All particles
- (d) None of these

Q9. Which quantum number determines the shape of orbital?

- (a) Principal (n)
- (b) Azimuthal (l)
- (c) Magnetic (m)
- (d) Spin (s)

Q10. Cathode rays are:

- (a) Electromagnetic waves
- (b) Stream of electrons
- (c) Stream of protons
- (d) Stream of α -particles

Q11. Radius of first Bohr orbit for H atom is:

- (a) 0.529 pm
- (b) 5.29 pm
- (c) 52.9 pm
- (d) 529 pm

Q12. Maximum number of electrons in M shell is:

- (a) 2
- (b) 8
- (c) 18
- (d) 32

Q13. Which orbital is dumbbell shaped?

- (a) s
- (b) p
- (c) d
- (d) f

Q14. Planck's constant has units of:

- (a) Energy
- (b) Linear momentum
- (c) Angular momentum
- (d) Power

Q15. Electronic configuration of Cu^+ ($Z=29$) is:

- (a) $[\text{Ar}] 3d^9 4s^1$
- (b) $[\text{Ar}] 3d^{10} 4s^0$
- (c) $[\text{Ar}] 3d^8 4s^2$
- (d) $[\text{Ar}] 3d^9 4s^0$

Q16. Aufbau principle is:

- (a) Electrons fill lowest energy orbitals first
- (b) No two electrons have same quantum numbers
- (c) Electrons singly occupy orbitals first
- (d) Electrons have spin

Q17. Who discovered neutron?

- (a) Thomson
- (b) Rutherford
- (c) Chadwick
- (d) Bohr

Q18. Energy difference between successive energy levels:

- (a) Increases with increasing n
- (b) Decreases with increasing n
- (c) Remains constant
- (d) None of these

Q19. Which of following violates Hund's rule?

- (a) $\uparrow \uparrow \uparrow (p^3)$
- (b) $\uparrow \downarrow \uparrow _ (p^3)$
- (c) $\uparrow \uparrow _ (p^2)$
- (d) $\uparrow \downarrow _ _ (p^2)$

Q20. Photoelectric effect supports:

- (a) Wave nature of light
- (b) Particle nature of light
- (c) Dual nature of matter
- (d) None of these

 **ANSWER KEY:**

1.(d) 2.(c) 3.(b) 4.(d) 5.(b) 6.(c) 7.(c) 8.(b) 9.(b) 10.(b)

11.(c) 12.(c) 13.(b) 14.(c) 15.(b) 16.(a) 17.(c) 18.(b) 19.(b) 20.(b)

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 **CASE-BASED QUESTIONS (As per New CBSE**

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Pattern)

Case Study 1: Atomic Spectra

Passage:

When an electron in hydrogen atom jumps from higher energy level to lower energy level, energy is emitted in the form of electromagnetic radiation. The wavelength of emitted radiation can be calculated using Rydberg formula. Different series of spectral lines are obtained depending on the lower energy level to which transition occurs. Balmer series (visible region) corresponds to transitions ending at $n=2$.

Q1. Lyman series lies in which region?

- (a) Visible
- (b) UV
- (c) IR
- (d) Radio

Q2. First line of Balmer series is due to transition:

- (a) $n=2$ to $n=1$
- (b) $n=3$ to $n=2$
- (c) $n=4$ to $n=2$
- (d) $n=3$ to $n=1$

Q3. Energy is emitted when electron:

- (a) Jumps to higher orbit
- (b) Jumps to lower orbit
- (c) Remains in same orbit
- (d) Is removed from atom

Q4. Rydberg constant for hydrogen is:

- (a) $1.097 \times 10^5 \text{ m}^{-1}$
- (b) $1.097 \times 10^6 \text{ m}^{-1}$
- (c) $1.097 \times 10^7 \text{ m}^{-1}$
- (d) $1.097 \times 10^8 \text{ m}^{-1}$

Answers: 1.(b) 2.(b) 3.(b) 4.(c)

Case Study 2: Quantum Numbers

Passage:

Four quantum numbers are required to completely describe an electron in an atom. Principal quantum number (n) describes the shell, azimuthal quantum number (l) describes the subshell and shape, magnetic quantum number (m) describes orientation, and spin quantum number (s) describes the spin direction. Each electron in an atom has a unique set of these four quantum numbers according to Pauli exclusion principle.

Q1. For d-subshell, value of l is:

- (a) 0
- (b) 1
- (c) 2
- (d) 3

Q2. Maximum number of orbitals in a shell is:

- (a) $2n$
- (b) n^2
- (c) $2n^2$
- (d) n

Q3. Number of orbitals in p-subshell is:

- (a) 1
- (b) 3
- (c) 5
- (d) 7

Q4. Possible values of spin quantum number are:

- (a) 0, 1, 2
- (b) $+1/2, -1/2$
- (c) -1, 0, +1
- (d) 1, 2, 3

Answers: 1.(c) 2.(b) 3.(b) 4.(b)

Case Study 3: Electronic Configuration

Passage:

Electronic configuration describes the distribution of electrons in different orbitals of an atom. It follows three main principles: Aufbau principle (lowest energy orbitals fill first), Pauli exclusion principle (maximum 2 electrons per orbital with opposite spins), and Hund's rule (electrons singly occupy orbitals before pairing). Some elements like Cr and Cu show exceptional configurations due to extra stability of half-filled and fully-filled d-subshells.

Q1. Electronic configuration of nitrogen (Z=7) is:

- (a) $1s^2 2s^2 2p^3$
- (b) $1s^2 2s^1 2p^4$
- (c) $1s^2 2s^3 2p^2$
- (d) $1s^1 2s^2 2p^4$

Q2. Which shows exceptional configuration?

- (a) Fe (26)
- (b) Cr (24)
- (c) Mn (25)
- (d) Co (27)

Q3. When Fe^{2+} is formed from Fe, electrons are lost from:

- (a) 3d orbital
- (b) 4s orbital
- (c) 3p orbital
- (d) 4p orbital

Q4. Half-filled configuration is more stable due to:

- (a) Symmetry
- (b) Exchange energy
- (c) Both (a) and (b)
- (d) None of these

Answers: 1.(a) 2.(b) 3.(b) 4.(c)

Topic-wise Preparation Priority:

1. HIGH PRIORITY (Must Score 100%):

- Bohr's model numericals (radius, energy, wavelength)
- Electronic configuration (especially exceptions!)
- Quantum numbers (very easy marks!)
- Spectral series of hydrogen

2. MEDIUM PRIORITY (Important):

- Discovery of subatomic particles
- Atomic models and their limitations
- de Broglie wavelength problems
- Shapes of orbitals

3. LOW PRIORITY (For extra marks):

- Historical dates and names
- Thomson's experiment details
- Mathematical derivations

⚠️ Last Minute Checklist (1 Day Before Exam):

- Memorize all formulas (write them 3 times!)
- Electronic configurations of first 36 elements
- Exceptions: Cr, Cu configurations
- Spectral series: Lyman, Balmer, Paschen, Brackett, Pfund
- Values: Bohr radius, ground state energy, Rydberg constant
- Practice 10 Bohr model numericals
- Quantum number rules
- Three principles: Aufbau, Pauli, Hund
- Important constants: h , c , e , m_e

🎯 Exam Writing Tips:

1. **Numericals:** Always show all steps, write formula first, then substitute
2. **Units:** Convert to SI units if needed, write units in final answer
3. **Definitions:** Write word-to-word from NCERT
4. **Diagrams:** If asked for atomic models, draw clean labeled diagrams
5. **Electronic Config:** Use noble gas notation [Ne], [Ar] to save time
6. **Negative Signs:** Never forget negative sign in energy!
7. **Check:** Recheck quantum number problems - very easy to make silly mistakes

 **Formula Sheet - Write This on Last Page!**

Concept	Formula
$c = v\lambda$	$c = 3 \times 10^8 \text{ m/s}$
$E = hv$	$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$
Rydberg Formula	$1/\lambda = R_H (1/n_1^2 - 1/n_2^2)$, $R_H = 1.097 \times 10^7 \text{ m}^{-1}$
Bohr Radius	$r_n = 0.529 \times n^2/Z \text{ \AA}$
Bohr Velocity	$v_n = 2.18 \times 10^6 \times Z/n \text{ m/s}$
Bohr Energy	$E_n = -13.6 \times Z^2/n^2 \text{ eV}$
de Broglie	$\lambda = h/mv$
Uncertainty	$\Delta x \cdot \Delta p \geq h/4\pi$
Angular Momentum	$mvr = nh/2\pi$
Max electrons	Shell: $2n^2$, Subshell: $2(2l+1)$

QUICK REVISION NOTES

Key Definitions (Must Know Word-to-Word)

- 1. Electron:** Negatively charged subatomic particle discovered by J.J. Thomson in cathode ray experiment.
- 2. Proton:** Positively charged subatomic particle present in nucleus.
- 3. Neutron:** Electrically neutral subatomic particle present in nucleus, discovered by Chadwick.
- 4. Nucleus:** Central core of atom containing protons and neutrons, carries positive charge.
- 5. Orbital:** Three-dimensional region around nucleus where probability of finding electron is maximum.
- 6. Quantum Numbers:** Set of four numbers (n, l, m, s) that completely describe an electron in atom.
- 7. Aufbau Principle:** Electrons fill orbitals in order of increasing energy.
- 8. Pauli Exclusion Principle:** No two electrons in atom can have same set of all four quantum numbers.
- 9. Hund's Rule:** Electrons singly occupy all degenerate orbitals before pairing begins.
- 10. Emission Spectrum:** Spectrum of radiation emitted by excited atoms as bright lines on dark background.

⚡ One-Line Facts (Speed Revision)

- e/m of electron = 1.758820×10^{11} C/kg
- Charge on electron = -1.6×10^{-19} C
- Mass of electron = 9.1094×10^{-31} kg
- Mass of proton = $1836 \times$ mass of electron
- Nucleus is 10^5 times smaller than atom
- Speed of light = 3×10^8 m/s
- Planck's constant = 6.626×10^{-34} J·s
- Bohr radius = $0.529 \text{ \AA} = 52.9 \text{ pm}$
- Ground state energy of H = -13.6 eV
- Ionization energy of H = $+13.6 \text{ eV}$
- Rydberg constant = $1.097 \times 10^7 \text{ m}^{-1}$
- Max electrons in shell = $2n^2$
- s orbital: spherical, p orbital: dumbbell
- Cr: $3d^5 4s^1$, Cu: $3d^{10} 4s^1$ (exceptions!)
- Balmer series: visible region


















Top 10 Question Types (Practice These!)

1. Calculate energy/wavelength for electron transition in H atom
2. Calculate radius of nth Bohr orbit
3. Write electronic configuration (watch for exceptions!)
4. Determine quantum numbers for given electron
5. Calculate de Broglie wavelength
6. Identify spectral series from transition
7. Calculate ionization energy
8. Apply Hund's rule to fill orbitals
9. Calculate maximum electrons in shell/subshell
10. Explain limitations of atomic models

Study Material Information

This comprehensive study material on **Structure of Atom** has been meticulously prepared following the latest CBSE curriculum and examination pattern for Class 11 Chemistry (2025-26 session). The content includes detailed explanations of all topics, discovery of subatomic particles (electron, proton, neutron), atomic models (Thomson, Rutherford, Bohr), electromagnetic radiation and atomic spectra, hydrogen spectrum and spectral series, Bohr's model with complete formulas and numericals, de Broglie's wave theory, Heisenberg's uncertainty principle, quantum mechanical model, four quantum numbers (n, l, m, s), shapes of orbitals (s, p, d, f), electronic configuration rules (Aufbau, Pauli, Hund), complete configurations of first 36 elements, 20 CBSE-pattern MCQs, 3 case-based questions, numerical problem solving, and exam strategy with tips and tricks.

Key Features of This Material:

-  Complete Chapter 2 coverage with crystal-clear concepts
-  Historical development: Thomson → Rutherford → Bohr → Quantum model
-  All subatomic particles with properties and discovery
-  Bohr's model - most important for numericals!
-  All formulas: radius, velocity, energy, wavelength
-  Hydrogen spectrum: Lyman, Balmer, Paschen, Brackett, Pfund series
-  Complete quantum numbers explanation with examples
-  Electronic configurations of first 36 elements (memorize!)
-  Important exceptions: Cr and Cu configurations explained
-  Orbital shapes (s, p, d, f) with clear descriptions
-  Three golden rules: Aufbau, Pauli, Hund
-  20 MCQs + 3 case studies as per latest CBSE pattern
-  Step-by-step solved numericals with all formulas
-  Memory tricks and shortcuts for fast learning
-  Common mistakes to avoid (very important!)
-  Exam strategy and last-minute checklist
-  Quick revision notes and complete formula sheet

Why This Chapter is SUPER Important:

Structure of Atom is the **foundation of ALL chemistry!**

Understanding this chapter is essential for:

- ✓ Chemical Bonding (Class 11 - next chapter!)
- ✓ Periodic Classification (Class 11)
- ✓ Coordination Compounds (Class 12)
- ✓ ALL of organic chemistry (hybridization!)
- ✓ JEE & NEET (direct 4-5 questions every year!)
- ✓ Understanding why elements behave the way they do

Master electronic configuration = 30% of chemistry becomes easy!

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This study material is provided for educational purposes only.

Content is based on NCERT syllabus and CBSE guidelines for 2025-26.

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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 2: Structure of Atom from Class 11 Chemistry NCERT textbook (Reprint 2025-26). All numerical values, formulas, and definitions are as per latest NCERT edition. Special focus on Bohr's model as it's the most important for numerical problems in boards and competitive exams.

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