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Class 12 Physics - Chapter 12: ATOMS



ATOMS

Chapter Overview

This chapter explores the structure of the atom and the revolutionary models that helped us understand matter at the atomic level!

- Thomson's Plum Pudding Model (1898)
- Rutherford's Alpha Scattering Experiment (1911)
- Rutherford's Nuclear Model of Atom
- Bohr's Model of Hydrogen Atom (1913)
- Energy Levels and Quantization
- Hydrogen Spectrum (Lyman, Balmer, Paschen Series)
- De Broglie's Wave-Particle Duality

1. INTRODUCTION

Historical Background

By the 19th century:

- Atomic hypothesis of matter was well established
- **1897:** J.J. Thomson discovered electrons - negatively charged constituents
- **Key Question:** How are positive and negative charges arranged inside the atom?

Thomson's Plum Pudding Model (1898)

Key Features

Thomson's Model:

- Positive charge uniformly distributed throughout atom volume
- Electrons embedded in it like seeds in watermelon
- Also called "Plum Pudding Model"
- **Limitation:** Could not explain experimental observations

Atomic Spectra

Spectral Observations

Continuous Spectrum:

- Emitted by solids, liquids, and dense gases
- Contains all wavelengths in a range

Line Spectrum:

- Emitted by rarefied gases (low pressure)
- Contains only specific discrete wavelengths
- Each element has characteristic spectrum - like "fingerprint"
- **Example:** Hydrogen spectrum shows fixed lines

Significance: This suggested intimate relationship between atomic structure and emitted radiation!

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★ 2. ALPHA-PARTICLE SCATTERING AND RUTHERFORD'S NUCLEAR MODEL

🎯 The Geiger-Marsden Experiment (1911)

Setup:

- Radioactive source: $^{214}_{83}\text{Bi}$ emitting 5.5 MeV α -particles
- Target: Thin gold foil (thickness 2.1×10^{-7} m)
- Detector: Zinc sulphide screen + microscope (rotatable)
- α -particles cause scintillations on screen

Observations:

- ✓ Most α -particles passed straight through foil
- ✓ Only 0.14% scattered by more than 1°
- ✓ About 1 in 8000 deflected by more than 90°
- ✓ Some particles bounced back (nearly 180°)

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🧠 Rutherford's Reasoning

To deflect α -particle backwards:

- Must experience very large repulsive force
- This requires concentration of positive charge in tiny space
- Most of atom must be empty space (explains easy passage)

🎯 **Conclusion:** Most mass and all positive charge concentrated in small central nucleus!

🌟 Rutherford's Nuclear Model

★ KEY FEATURES ★

- **Nucleus:** Contains entire positive charge and most mass
- **Size:** Nuclear radius $\sim 10^{-15}$ m to 10^{-14} m
- **Atomic radius:** $\sim 10^{-10}$ m
- **Ratio:** Atom is 10,000 to 100,000 times larger than nucleus
- **Electrons:** Revolve around nucleus like planets around sun
- **Most of atom:** Empty space!

Force on α -Particle

Coulomb Force:

$$F = (1/4\pi\epsilon_0) \times (2e)(Ze)/r^2$$

Where:

- $2e$ = charge on α -particle (helium nucleus)
- Ze = charge on target nucleus (Z = atomic number)
- r = distance between α -particle and nucleus

Impact Parameter and Scattering

Impact Parameter (b)

Definition: Perpendicular distance of initial velocity vector from center of nucleus

Impact Parameter	Scattering Angle (θ)	Result
$b = 0$ (Head-on)	$\theta \approx \pi$ (180°)	α -particle rebounds back
Small b	Large θ	Large scattering
Large b	Small $\theta \approx 0$	Nearly undeviated

NCERT Example 12.1 - Atomic Empty Space

Problem: Compare atom proportions with solar system. If electron orbit/nucleus = earth orbit/sun, where would earth be?

Given:

- Nuclear radius $\sim 10^{-15}$ m
- Atomic radius $\sim 10^{-10}$ m
- Earth's orbit radius = 1.5×10^{11} m
- Sun's radius = 7×10^8 m

Solution:

Step 1: Ratio in atom

$$\text{Electron orbit/nucleus} = (10^{-10} \text{ m}) / (10^{-15} \text{ m}) = 10^5$$

Step 2: If same ratio for solar system

$$\text{Earth orbit} = 10^5 \times \text{Sun radius}$$

$$= 10^5 \times 7 \times 10^8 \text{ m} = 7 \times 10^{13} \text{ m}$$

Step 3: Compare with actual

This is more than **100 times** the actual orbital radius!

✨ **Conclusion:** Atom contains much greater fraction of empty space than solar system!

NCERT Example 12.2 - Distance of Closest Approach

Problem: In Geiger-Marsden experiment, what is distance of closest approach for 7.7 MeV α -particle to gold nucleus?

Solution:

Key Concept: Energy conservation

Initial KE = Final PE (at stopping point)

Step 1: Conservation equation

$$K = (1/4\pi\epsilon_0) \times (2e)(Ze)/d$$

Step 2: Distance of closest approach

$$d = (1/4\pi\epsilon_0) \times (2Ze^2/K)$$

Step 3: Substitute values

$$K = 7.7 \text{ MeV} = 1.2 \times 10^{-12} \text{ J}$$

$$1/4\pi\epsilon_0 = 9.0 \times 10^9 \text{ N}\cdot\text{m}^2/\text{C}^2$$

$$e = 1.6 \times 10^{-19} \text{ C}$$

$$Z = 79 \text{ (for gold)}$$

$$d = (2)(9.0 \times 10^9)(1.6 \times 10^{-19})^2(79)/(1.2 \times 10^{-12})$$

$$\mathbf{d = 3.0 \times 10^{-14} \text{ m} = 30 \text{ fm}}$$

 This gives upper limit for nuclear radius!

⚡ Electron Orbits in Rutherford Model

▢ Stable Orbit Condition

For dynamically stable orbit:

Electrostatic force = Centripetal force

$$(1/4\pi\epsilon_0) \times (e^2/r^2) = mv^2/r$$

This gives:

$$r = e^2/(4\pi\epsilon_0 mv^2)$$

Energy Calculations:

- **Kinetic Energy:** $K = \frac{1}{2}mv^2 = e^2/(8\pi\epsilon_0 r)$
- **Potential Energy:** $U = -e^2/(4\pi\epsilon_0 r)$
- **Total Energy:** $E = K + U = -e^2/(8\pi\epsilon_0 r)$

⚠ **Negative total energy** means electron is bound to nucleus!

✗ Problems with Rutherford Model

⚠ FATAL FLAWS

Problem 1: Instability

- Revolving electron undergoes centripetal acceleration
- According to classical EM theory: accelerated charge radiates energy
- Electron should lose energy and spiral into nucleus
- **Contradiction:** Atoms ARE stable!

Problem 2: Continuous Spectrum

- As electron spirals in, frequency changes continuously
- Should emit continuous spectrum
- **Contradiction:** Atoms emit line spectrum!

💡 **Solution needed:** Classical physics fails at atomic level. New quantum theory required!

☀️ 3. BOHR MODEL OF HYDROGEN ATOM (1913)

👨‍🔬 Niels Bohr (1885-1962)

- Danish physicist who revolutionized atomic theory
- Studied in Rutherford's laboratory (1912)
- Combined classical and quantum concepts
- **1922:** Awarded Nobel Prize in Physics

★ Bohr's Three Postulates

🎯 BOHR'S POSTULATES

📌 POSTULATE 1: Stationary States

Statement: An electron can revolve in certain stable orbits without emitting radiation.

- Each atom has definite stable states (stationary states)
- Each state has definite total energy
- No energy is radiated in these orbits
- **Revolutionary:** Contradicts classical EM theory!

✦ POSTULATE 2: Quantization of Angular Momentum

$$L = nh/2\pi$$

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

$h = 6.63 \times 10^{-34}$ J·s (Planck's constant)

Meaning:

- Angular momentum can only have discrete values
- $n = 1$: ground state (lowest energy)
- $n = 2, 3, \dots$: excited states
- Only certain orbits are allowed!

✦ POSTULATE 3: Energy Transitions

Statement: When electron transitions between orbits, photon is emitted/absorbed.

$$h\nu = E_i - E_f$$

E_i = Initial energy

E_f = Final energy

ν = Frequency of photon

Two Processes:

- **Emission:** $E_i > E_f \rightarrow$ Photon emitted
- **Absorption:** $E_i + h\nu = E_f \rightarrow$ Photon absorbed

Radius of Bohr Orbits

Derivation of Orbital Radius

From stability condition:

$$(1/4\pi\epsilon_0) \times (e^2/r^2) = mv^2/r$$

$$\text{Therefore: } r = e^2/(4\pi\epsilon_0mv^2) \dots (1)$$

From quantization:

$$mvr = nh/2\pi$$

$$\text{Therefore: } v = nh/(2\pi mr) \dots (2)$$

Substitute (2) in (1):

$$r = e^2/(4\pi\epsilon_0m) \times [2\pi mr/(nh)]^2$$

$$r = (4\pi\epsilon_0/me^2) \times (n^2h^2/4\pi^2m)$$

$$r = (\epsilon_0h^2/\pi me^2) \times n^2$$

$$\mathbf{r_n = (n^2/m) \times (h/2\pi)^2 \times (4\pi\epsilon_0/e^2)}$$

$$\mathbf{r_n = n^2 \times r_1}$$

Bohr Radius (n=1):

$$r_1 = 0.529 \times 10^{-10} \text{ m} = 0.529 \text{ \AA} = 53 \text{ pm}$$

⚡ Energy of Bohr Orbits

▢ Derivation of Energy Levels

Total Energy:

$$E = K + U = \frac{e^2}{(8\pi\epsilon_0 r)} - \frac{e^2}{(4\pi\epsilon_0 r)}$$

$$E = -\frac{e^2}{(8\pi\epsilon_0 r)}$$

Substitute r_n :

$$E_n = -\frac{e^2}{(8\pi\epsilon_0)} \times \frac{1}{[n^2(\epsilon_0 h^2 / \pi m e^2)]}$$

$$E_n = -\frac{(m e^4)}{(8\epsilon_0^2 h^2 n^2)}$$

★ ENERGY FORMULA ★

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

$$\text{Or: } E_n = -2.18 \times 10^{-18}/n^2 \text{ J}$$

💡 Understanding Energy Values

n (orbit)	Energy E_n (eV)	State
1	-13.6 eV	Ground state
2	-3.4 eV	First excited
3	-1.51 eV	Second excited
4	-0.85 eV	Third excited
∞	0 eV	Ionized (free)

Key Points:

- Negative energy means electron is bound
- Energy increases (becomes less negative) with n
- Energy levels get closer as n increases
- $E = 0$ means electron is free (ionized)

NCERT Example 12.3 - Hydrogen Atom Ground State

Problem: 13.6 eV energy is required to ionize hydrogen. Find orbital radius and electron velocity.

Solution:

Step 1: Energy in ground state

$$E = -13.6 \text{ eV} = -13.6 \times 1.6 \times 10^{-19} \text{ J} = -2.18 \times 10^{-18} \text{ J}$$

Step 2: Find orbital radius

$$E = -e^2/(8\pi\epsilon_0 r)$$

$$r = -e^2/(8\pi\epsilon_0 E)$$

$$r = -(9 \times 10^9)(1.6 \times 10^{-19})^2/[2(-2.18 \times 10^{-18})]$$

$$\mathbf{r = 5.3 \times 10^{-11} \text{ m} = 0.53 \text{ \AA}}$$

Step 3: Find velocity

$$\text{From } r = e^2/(4\pi\epsilon_0 mv^2)$$

$$v = e/\sqrt{(4\pi\epsilon_0 mr)}$$

$$v = (1.6 \times 10^{-19})/\sqrt{[(4\pi)(8.85 \times 10^{-12})(9.1 \times 10^{-31})(5.3 \times 10^{-11})]}$$

$$\mathbf{v = 2.2 \times 10^6 \text{ m/s}}$$

This is about 0.7% of speed of light!

Ionization Energy

Key Concept

Ionization Energy: Minimum energy required to remove electron from ground state to infinity.

$$E_{\text{ionization}} = E_{\infty} - E_1 = 0 - (-13.6) = \mathbf{13.6 \text{ eV}}$$

Excitation Energies:

- **n=1 to n=2:** $\Delta E = -3.4 - (-13.6) = 10.2 \text{ eV}$
- **n=1 to n=3:** $\Delta E = -1.51 - (-13.6) = 12.09 \text{ eV}$
- **n=1 to n=4:** $\Delta E = -0.85 - (-13.6) = 12.75 \text{ eV}$

4. LINE SPECTRA OF HYDROGEN ATOM

Spectral Series

When electron transitions occur:

- Photon frequency: $\nu = (E_i - E_f)/h$
- Wavelength: $\lambda = c/\nu = hc/(E_i - E_f)$

General Formula:

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

Rydberg Constant:

$$R_H = 1.097 \times 10^7 \text{ m}^{-1}$$

Hydrogen Spectral Series

Series Name	Final State (n_f)	Initial State (n_i)	Region	Wavelength Range
Lyman Series	$n_f = 1$	$n_i = 2, 3, 4, \dots$	Ultraviolet	91 nm - 122 nm
Balmer Series	$n_f = 2$	$n_i = 3, 4, 5, \dots$	Visible	365 nm - 656 nm
Paschen Series	$n_f = 3$	$n_i = 4, 5, 6, \dots$	Infrared	820 nm - 1875 nm
Brackett Series	$n_f = 4$	$n_i = 5, 6, 7, \dots$	Infrared	1.46 μm - 4.05 μm
Pfund Series	$n_f = 5$	$n_i = 6, 7, 8, \dots$	Infrared	2.28 μm - 7.46 μm

Balmer Series - Visible Light

Famous lines in visible spectrum:

- **H- α ($n=3 \rightarrow 2$):** 656 nm (Red)
- **H- β ($n=4 \rightarrow 2$):** 486 nm (Blue-green)
- **H- γ ($n=5 \rightarrow 2$):** 434 nm (Violet)
- **H- δ ($n=6 \rightarrow 2$):** 410 nm (Violet)

Balmer Formula (1885):

$$\lambda = 364.6 \frac{n^2}{(n^2 - 4)} \text{ nm, where } n = 3, 4, 5, \dots$$

5. DE BROGLIE'S EXPLANATION

Louis de Broglie (1923)

Revolutionary Idea: Electrons have wave nature!

- De Broglie wavelength: $\lambda = h/p = h/(mv)$
- Verified by Davisson-Germer experiment (1927)
- Explained Bohr's mysterious quantization condition

🎵 Standing Wave Condition

📐 Explanation of Quantization

For standing wave in circular orbit:

Circumference must equal whole number of wavelengths

$$2\pi r_n = n\lambda, \text{ where } n = 1, 2, 3, \dots$$

Substitute de Broglie wavelength:

$$\lambda = h/(mv_n)$$

$$2\pi r_n = n \times h/(mv_n)$$

$$mv_n r_n = nh/(2\pi)$$

$$\star \mathbf{L = nh/2\pi} \star$$

This is EXACTLY Bohr's second postulate!

✨ **De Broglie explained quantization!** Only standing waves persist in atom. Orbits with non-integer wavelengths destructively interfere.



6. KEY FORMULAE AND CONSTANTS

⚡ FUNDAMENTAL CONSTANTS

Constant	Symbol	Value
Planck's constant	h	$6.63 \times 10^{-34} \text{ J}\cdot\text{s}$
Electron mass	m_e	$9.1 \times 10^{-31} \text{ kg}$
Electron charge	e	$1.6 \times 10^{-19} \text{ C}$
Permittivity	ϵ_0	$8.85 \times 10^{-12} \text{ F/m}$
Speed of light	c	$3.0 \times 10^8 \text{ m/s}$
Rydberg constant	R_H	$1.097 \times 10^7 \text{ m}^{-1}$

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▶ IMPORTANT FORMULAE

Formula	Description
$L = nh/2\pi$	Angular momentum quantization
$r_n = n^2 a_0$	Bohr orbit radius ($a_0 = 0.529 \text{ \AA}$)
$E_n = -13.6/n^2 \text{ eV}$	Energy levels
$v_n = (c/137) \times (1/n)$	Electron velocity ($c/137 = 2.2 \times 10^6 \text{ m/s}$)
$h\nu = E_i - E_f$	Photon energy in transition
$1/\lambda = R_H(1/n_f^2 - 1/n_i^2)$	Rydberg formula for wavelength
$\lambda = h/(mv)$	De Broglie wavelength
$2\pi r_n = n\lambda$	Standing wave condition

⚠ 7. COMMON MISTAKES

✘ MISTAKE 1: Confusing Energy Signs

Wrong: "Energy of ground state is +13.6 eV"

Correct: $E_1 = -13.6 \text{ eV}$ (negative means bound state)

Remember: Ionization energy = 13.6 eV (energy needed to ionize)

✘ MISTAKE 2: Quantum Number Starting Point

Wrong: "Ground state has $n = 0$ "

Correct: Ground state has $n = 1$

Remember: $n = 1, 2, 3, \dots$ (positive integers only)

✘ MISTAKE 3: Emission vs Absorption

Wrong: "Photon emitted when electron goes from $n=2$ to $n=3$ "

Correct: Photon is **absorbed** (going to higher energy)

Emission: $n_i > n_f$ (higher \rightarrow lower)

Absorption: $n_i < n_f$ (lower \rightarrow higher)

✗ MISTAKE 4: Series Identification

Wrong: "All transitions ending at $n=1$ are in visible region"

Correct: Lyman series ($n \rightarrow 1$) is in **UV region**

Only Balmer series ($n \rightarrow 2$) is in visible region!

✗ MISTAKE 5: De Broglie Wavelength

Wrong: " $\lambda = h/m$ (missing velocity)"

Correct: $\lambda = h/(mv)$ or $\lambda = h/p$

Remember: Wavelength depends on momentum, not just mass!

8. PRACTICE QUESTIONS

MCQ Section

Q1. The angular momentum of electron in the ground state of hydrogen atom is:

- (a) h
- (b) $h/2\pi$ ✓
- (c) h/π
- (d) $2h/\pi$

Answer: (b) For $n=1$, $L = 1 \times h/2\pi$

Q2. In Bohr model, the ratio of radii of $n=2$ to $n=1$ orbits is:

- (a) 2:1
- (b) 4:1 ✓
- (c) 8:1
- (d) 1:4

Answer: (b) $r_n \propto n^2$, so $r_2/r_1 = 4/1$

Q3. Balmer series lies in which region?

- (a) UV
- (b) Visible ✓
- (c) IR
- (d) X-ray

Answer: (b) Transitions to $n=2$ are visible

Q4. Energy of first excited state of hydrogen is:

- (a) -13.6 eV
- (b) -3.4 eV ✓
- (c) -1.51 eV
- (d) -6.8 eV

Answer: (b) $E_2 = -13.6/4 = -3.4$ eV

Q5. De Broglie wavelength of electron is inversely proportional to:

- (a) Mass
- (b) Velocity
- (c) Momentum ✓
- (d) Energy

Answer: (c) $\lambda = h/p$, so $\lambda \propto 1/p$

Short Answer Questions (2-3 marks)

Q6. State two limitations of Rutherford's nuclear model.

Answer:

Limitation 1: Cannot explain stability of atom

- Revolving electron should radiate energy continuously
- Should spiral into nucleus in $\sim 10^{-8}$ seconds
- Contradicts observed atomic stability

Limitation 2: Cannot explain line spectrum

- Classical theory predicts continuous spectrum
- Atoms actually emit discrete line spectrum
- No explanation for characteristic spectra

Q7. Calculate wavelength of photon emitted in transition $n=3$ to $n=2$ in hydrogen.

Answer:

Energy difference:

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

$$\Delta E = 1.89 \times 1.6 \times 10^{-19} \text{ J} = 3.02 \times 10^{-19} \text{ J}$$

Wavelength:

$$hc/\lambda = \Delta E$$

$$\lambda = hc/\Delta E = (6.63 \times 10^{-34})(3 \times 10^8)/(3.02 \times 10^{-19})$$

$\lambda = 656 \text{ nm}$ (H-alpha line, red)

Q8. Why did Bohr quantize angular momentum and not some other quantity?

Answer:

Reasons:

- Planck's constant h has dimensions of angular momentum ($\text{J}\cdot\text{s}$)
- For circular orbits, angular momentum is natural variable
- Quantization gives discrete orbits naturally
- Leads to correct energy levels matching experimental data

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Long Answer Questions (5 marks)

Q9. Derive expression for radius of nth Bohr orbit and calculate Bohr radius.

Answer:

Step 1: Stable orbit condition

Electrostatic force = Centripetal force

$$(1/4\pi\epsilon_0)(e^2/r^2) = mv^2/r$$

$$\text{Therefore: } r = e^2/(4\pi\epsilon_0mv^2) \dots (1)$$

Step 2: Quantization condition

$$L = mvr = nh/2\pi$$

$$v = nh/(2\pi mr) \dots (2)$$

Step 3: Substitute (2) in (1)

$$r = e^2/(4\pi\epsilon_0m) \times (2\pi mr/nh)^2$$

$$r = e^2 \times 4\pi^2m^2r^2/(4\pi\epsilon_0m \times n^2h^2)$$

$$1 = e^2\pi mr/(\epsilon_0n^2h^2)$$

$$r = \epsilon_0n^2h^2/(\pi me^2)$$

Step 4: Final formula

$$r_n = (n^2/m)(h/2\pi)^2(4\pi\epsilon_0/e^2)$$

Step 5: Bohr radius (n=1)

$$r_1 = \epsilon_0h^2/(\pi me^2)$$

$$= (8.85 \times 10^{-12})(6.63 \times 10^{-34})^2/[\pi(9.1 \times 10^{-31})(1.6 \times 10^{-19})^2]$$

$$r_1 = 0.529 \times 10^{-10} \text{ m} = 0.529 \text{ \AA}$$

Q10. Explain de Broglie's explanation of Bohr's quantization condition with diagram.

Answer:

De Broglie's Wave-Particle Duality:

Electrons behave as both particles and waves with wavelength $\lambda = h/(mv)$

Standing Wave Condition:

- Electron wave must form standing wave in circular orbit
- Only certain wavelengths can fit exactly
- Like standing waves on string with fixed ends

Mathematical Derivation:

For standing wave: Circumference = $n \times$ wavelength

$$2\pi r_n = n\lambda$$

$$2\pi r_n = n(h/mv_n)$$

$$mv_n r_n = nh/2\pi$$









$$L = nh/2\pi$$

Significance:

- This is exactly Bohr's second postulate!
- Quantization arises naturally from wave nature
- Orbits with non-integer wavelengths destructively interfere
- Only resonant standing waves persist

9. EXAM PREPARATION

Must Master Concepts:

-  Thomson's vs Rutherford's vs Bohr's models
-  Geiger-Marsden experiment and conclusions
-  Bohr's three postulates
-  Derivation of r_n and E_n formulae
-  Energy level diagram
-  Hydrogen spectral series
-  De Broglie's wave-particle duality
-  Limitations of Bohr model

EXAM TIPS

For Derivations:

- Start with force balance equation clearly
- State quantization condition explicitly
- Show all substitution steps
- Box final formulas
- Don't skip unit conversions

Common Question Types (with marks):

1. State Bohr's postulates (3 marks)
2. Derive radius/energy formula (5 marks)
3. Calculate wavelength in transitions (3 marks)
4. Explain de Broglie hypothesis (3 marks)
5. Compare atomic models (3 marks)
6. Identify spectral series (2 marks)
7. Distance of closest approach (3 marks)

Important Values to Memorize

Quantity	Value
Ground state energy	$E_1 = -13.6 \text{ eV}$
Bohr radius	$a_0 = 0.529 \text{ \AA}$
Ionization energy	13.6 eV
Planck's constant	$h = 6.63 \times 10^{-34} \text{ J}\cdot\text{s}$
Rydberg constant	$R_H = 1.097 \times 10^7 \text{ m}^{-1}$
Electron velocity (n=1)	$v_1 = 2.2 \times 10^6 \text{ m/s}$

Key Formulas to Remember

1. **Angular momentum:** $L = nh/2\pi$
2. **Orbit radius:** $r_n = n^2 a_0 = n^2(0.529 \text{ \AA})$
3. **Energy:** $E_n = -13.6/n^2 \text{ eV}$
4. **Velocity:** $v_n = (2.2 \times 10^6)/n \text{ m/s}$
5. **Photon energy:** $h\nu = E_i - E_f$
6. **Rydberg formula:** $1/\lambda = R_H(1/n_f^2 - 1/n_i^2)$
7. **De Broglie:** $\lambda = h/(mv)$

Spectral Series Quick Reference

- **Lyman (UV):** $n \rightarrow 1$
- **Balmer (Visible):** $n \rightarrow 2$
- **Paschen (IR):** $n \rightarrow 3$
- **Brackett (IR):** $n \rightarrow 4$
- **Pfund (IR):** $n \rightarrow 5$

Study Material Information

This comprehensive study material on **Atoms** (Chapter 12, Class 12 Physics) has been prepared following the latest CBSE curriculum. The content includes detailed explanations, NCERT solved examples, practice questions, and exam-focused tips.

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