

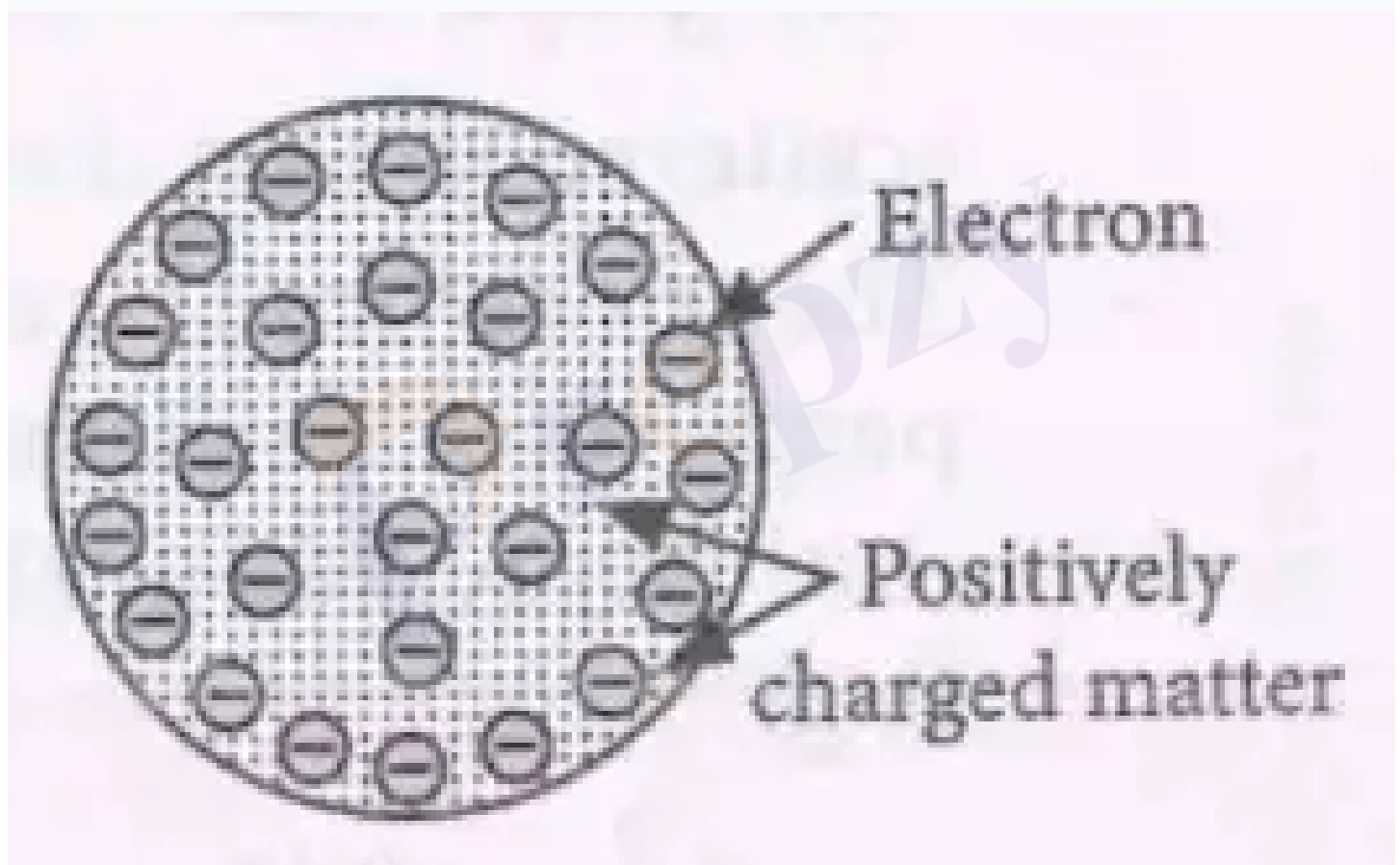
Table of Contents

1. Plum Pudding Model
2. Alpha Particle Scattering
3. Rutherford's Nuclear Model
4. Electron Orbits
5. Atomic Spectra
6. Absorption Spectrum
7. Spectral Series
8. Bohr Model
9. Energy Transitions
10. Orbit Radius & Energy Levels
11. Energy Quantization
12. Line Spectra of Hydrogen
13. Continuous and Characteristic X-rays
14. De-Broglie's Explanation of Bohr's Second Postulate of Quantisation
15. Limitations of Bohr's Theory
16. Moseley's law
17. Bragg's law

Plum Pudding Model

Model Description

- Proposed by J.J. Thomson in 1898.
- Atom consists of a sphere of positive charge with electrons embedded like seeds in a watermelon.
- Positive charge is uniformly distributed throughout the atom's volume.
- Electrons are negatively charged and scattered inside the positive sphere.
- Model was later disproved by experiments showing different charge distribution.



Solved Examples

- **Ex 1:** Identify the charge distribution in the plum pudding model.
Answer: Positive charge is spread uniformly; electrons embedded inside.

Practice Set

- Q1: What does the plum pudding model propose about the atom?
 - A. Nucleus at center
 - B. Electrons orbit nucleus
 - C. Positive charge spread uniformly with embedded electrons
 - D. Electrons outside nucleus
- Q2: According to the plum pudding model, electrons are:
 - A. Orbiting nucleus
 - B. Embedded in positive charge
 - C. Free particles
 - D. Absent
- Q3: Which experiment disproved the plum pudding model?
 - A. Photoelectric effect
 - B. Rutherford's alpha scattering
 - C. Compton scattering
 - D. Davisson-Germer experiment

Answer Key

- Q1: C - Positive charge is uniformly spread with electrons embedded inside.
- Q2: B - Electrons are embedded in the positive charge sphere.
- Q3: B - Rutherford's alpha scattering experiment disproved this model.

Memory Aids

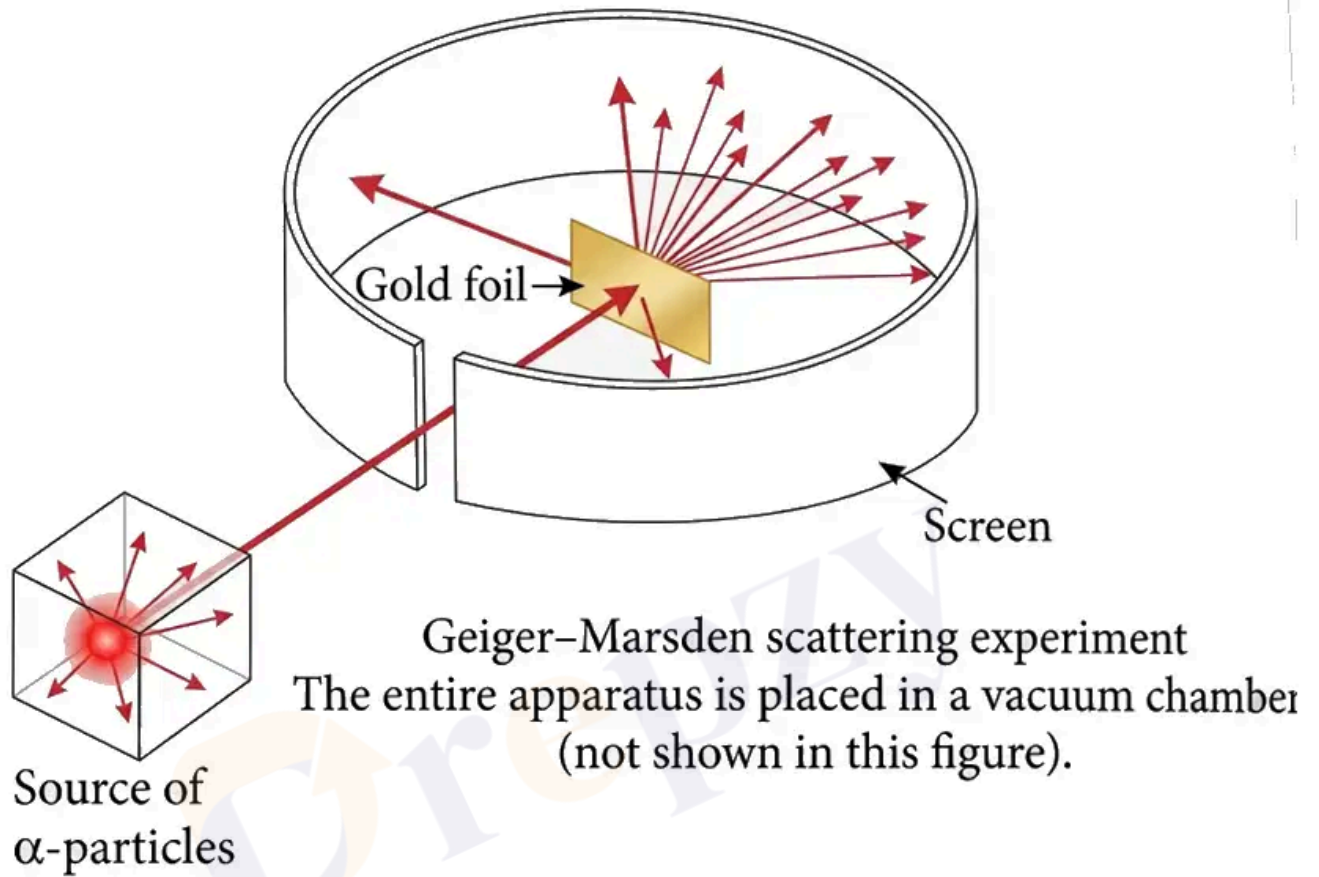
- Think "plums in pudding" for electrons in positive charge.
- Positive charge = pudding; electrons = plums.

Alpha Particle Scattering

Experiment Setup

- Performed by Geiger and Marsden under Rutherford's suggestion in 1911.

- 5.5 MeV alpha particles from $^{214}_{83}\text{Bi}$ source directed at thin gold foil (thickness 2.1×10^{-7} m).
- Alpha particles collimated into narrow beam using lead bricks.
- Scattered particles detected by rotatable zinc sulphide screen and microscope.
- Scintillations on ZnS screen indicate alpha particle impacts.



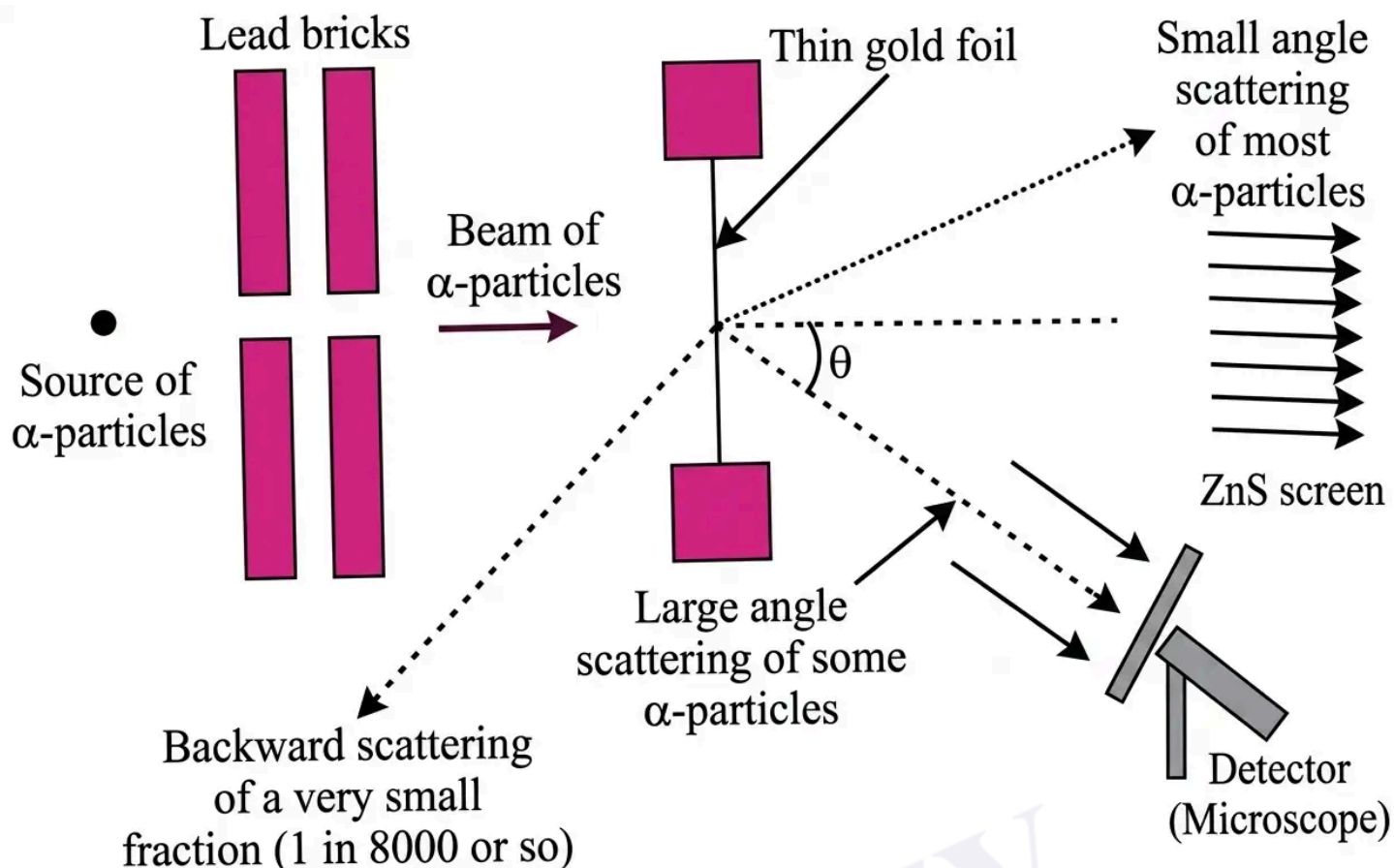


Figure : Schematic arrangement of the Geiger-Marsden experiment

Observations

- Most alpha particles pass through gold foil undeflected $\rightarrow$ atom mostly empty space.
- Some alpha particles deflect at various angles $\rightarrow$ presence of small, dense, positively charged nucleus.
- Very few (1 in 8000) deflect by $180^\circ \rightarrow$ nucleus size is very small ($\sim 1/8000$ of atom size).
- Number of scattered particles $N \propto \frac{1}{\sin^4(\theta/2)}$.

Solved Examples

- **Ex 1:** What does the deflection of alpha particles indicate?
Answer: Presence of a small, dense, positively charged nucleus.
- **Ex 2:** Calculate the ratio of nucleus size to atom size if 1 in 8000 alpha particles deflect by 180° .
Answer: Nucleus size $\approx \frac{1}{8000}$ of atom size.

Practice Set

- Q1: What does most alpha particles passing through gold foil indicate?
 - A. Atom is solid
 - B. Atom is mostly empty
 - C. Atom has no nucleus
 - D. Alpha particles are absorbed
- Q2: The scattering angle of alpha particles depends on:
 - A. Electron charge
 - B. Impact parameter
 - C. Alpha particle mass
 - D. Foil thickness
- Q3: The number of scattered alpha particles N varies with scattering angle θ as:
 - A. $N \propto \sin^2 \theta$
 - B. $N \propto \cos^2 \theta$
 - C. $N \propto \frac{1}{\sin^4(\theta/2)}$
 - D. $N \propto \tan \theta$

Answer Key

- Q1: B - Most alpha particles pass through, showing atom is mostly empty.
- Q2: B - Scattering angle depends on impact parameter (distance from nucleus).
- Q3: C - Number of scattered particles varies as $1/\sin^4(\theta/2)$.

Memory Aids

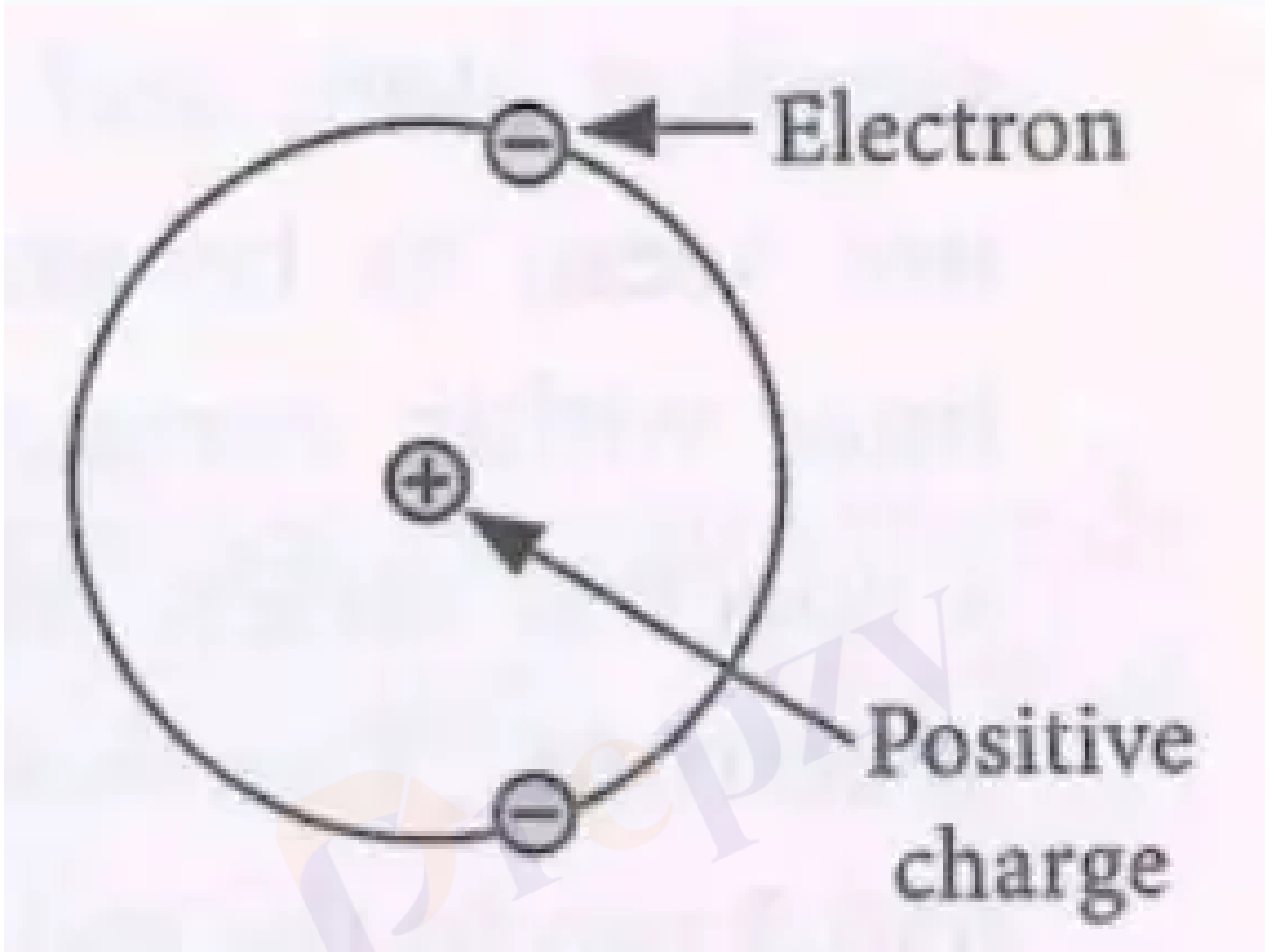
- Alpha scattering → nucleus discovery.
- Deflection angle ↔ impact parameter.
- $N \propto 1/\sin^4(\theta/2)$ curve shape.

Rutherford's Nuclear Model

Postulates

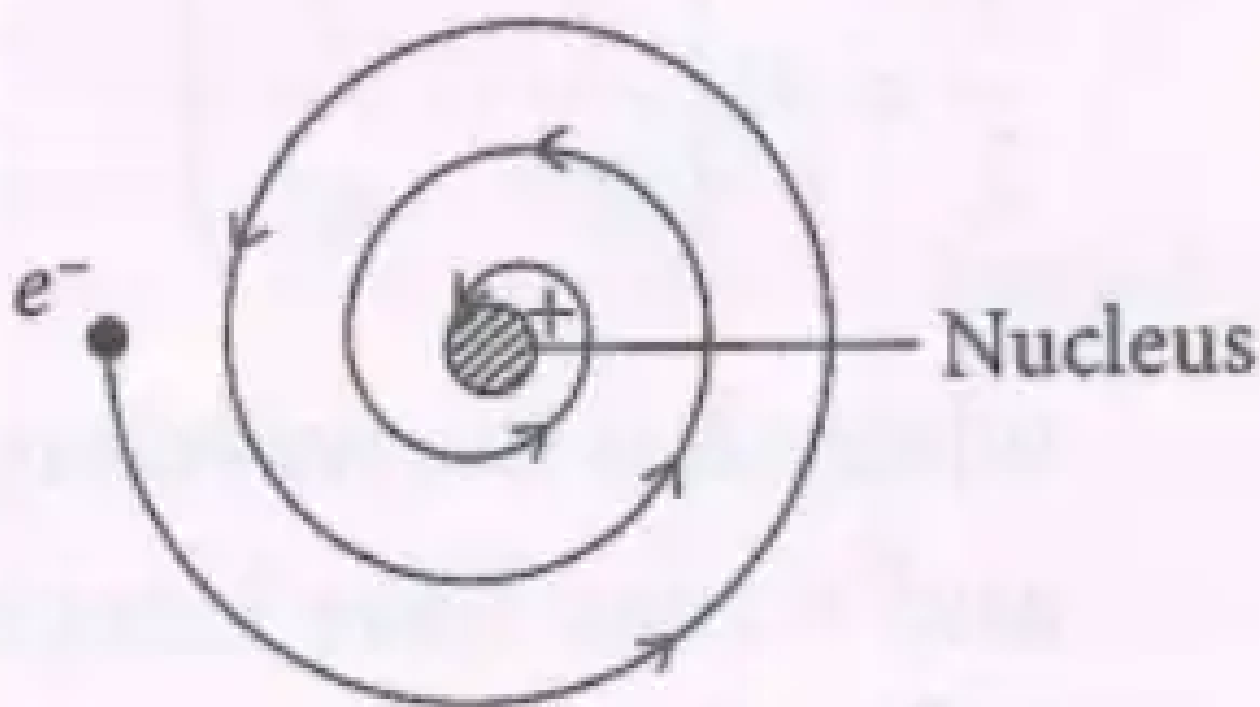
- Positive charge and most mass concentrated in tiny nucleus ($\sim 10^{-14}$ m diameter).

- Electrons revolve around nucleus in empty space, held by electrostatic attraction.
- Atom is electrically neutral: total positive charge = total negative charge.



Drawbacks

- Accelerated electrons should radiate energy and spiral into nucleus, collapsing atom (not observed).
- Predicted continuous energy spectrum, but actual atomic spectra are discrete.



Spiral Orbit of e^-

Solved Examples

- **Ex 1:** Why does Rutherford model predict atom collapse?
Answer: Accelerated electrons radiate energy, lose it, and spiral into nucleus.
- **Ex 2:** What is the size difference between atom and nucleus?
Answer: Atom $\sim 10^{-10}$ m, nucleus $\sim 10^{-14}$ m.

Practice Set

- Q1: What is concentrated in the nucleus according to Rutherford?
A. Electrons
B. Positive charge and mass
C. Neutrons only
D. Empty space
- Q2: Why does Rutherford model fail to explain atomic stability?
A. Electrons do not revolve
B. Electrons radiate energy and collapse

- C. Nucleus is too big
- D. Atoms are not neutral
- Q3: What type of spectrum does Rutherford model fail to explain?
 - A. Continuous
 - B. Discrete
 - C. Emission
 - D. Absorption

Answer Key

- Q1: B - Positive charge and mass are concentrated in nucleus.
- Q2: B - Accelerated electrons radiate energy and collapse atom.
- Q3: A - Model predicts continuous spectrum, but actual is discrete.

Memory Aids

- Nucleus = dense positive core.
- Electron spiral collapse = drawback.
- Atom mostly empty space.

Electron Orbits

Forces and Energy

- Electrostatic force F_e provides centripetal force F_c for electron orbit.
- $F_e = F_c \Rightarrow \frac{1}{4\pi\epsilon_0} \frac{e^2}{r^2} = \frac{mv^2}{r}$
- Orbit radius: $r = \frac{e^2}{4\pi\epsilon_0 mv^2}$
- Kinetic energy: $K = \frac{1}{2}mv^2 = \frac{e^2}{8\pi\epsilon_0 r}$
- Potential energy: $U = -\frac{e^2}{4\pi\epsilon_0 r}$
- Total energy: $E = K + U = -\frac{e^2}{8\pi\epsilon_0 r}$ (negative, electron bound to nucleus)

Solved Examples

- **Ex 1:** Calculate orbit radius if electron velocity v is known.

Use $r = \frac{e^2}{4\pi\epsilon_0 m v^2}$.

- **Ex 2:** Find total energy of electron in orbit radius r .

Use $E = -\frac{e^2}{8\pi\epsilon_0 r}$.

Practice Set

- Q1: What force keeps electron in orbit?
A. Gravitational
B. Electrostatic attraction
C. Magnetic
D. Nuclear
- Q2: Total energy of electron in orbit is:
A. Positive
B. Zero
C. Negative
D. Infinite
- Q3: If electron velocity increases, orbit radius:
A. Increases
B. Decreases
C. Remains same
D. Becomes zero

Answer Key

- Q1: B - Electrostatic force provides centripetal force.
- Q2: C - Total energy is negative, electron is bound.
- Q3: B - Radius inversely proportional to square of velocity.

Memory Aids

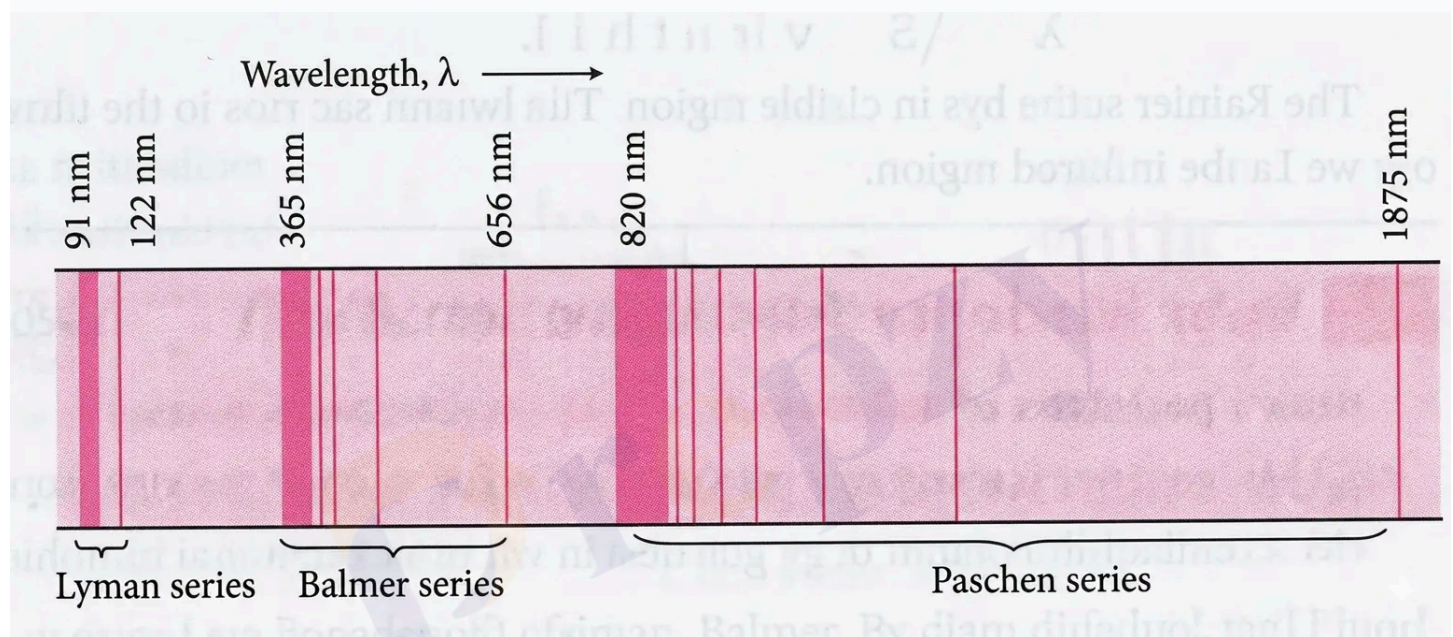
- Orbit radius $r \propto \frac{1}{v^2}$.
- Total energy negative $\rightarrow$ bound electron.

- Electrostatic force = centripetal force.

Atomic Spectra

Emission Line Spectrum

- Excited atomic gas emits radiation at specific wavelengths only.
- Emission line spectrum: bright lines on dark background.
- Example: Hydrogen emission spectrum shows distinct lines.



Explanation of Hydrogen Spectrum

- Electrons change energy levels, emitting light at specific wavelengths.
- Series: Lyman (UV), Balmer (visible), Paschen (IR).
- Lyman: 91-122 nm, Balmer: 365-656 nm, Paschen: 820-1875 nm.
- Each line = electron dropping to lower energy level, releasing photon.

Solved Examples

- **Ex 1:** Identify the series for a line at 486.1 nm.
Solution: 486.1 nm lies in visible range → Balmer series.

Practice Set

- Q1: Which series of hydrogen spectrum lies in the ultraviolet region?
A. Balmer
B. Paschen
C. Lyman
D. Brackett
- Q2: The emission line at 656.3 nm belongs to which series?
A. Lyman
B. Balmer
C. Paschen
D. Pfund
- Q3: What happens when an electron falls from $n=4$ to $n=2$ in hydrogen?
A. Emits photon in Balmer series
B. Emits photon in Lyman series
C. Absorbs photon
D. Emits photon in Paschen series

Answer Key

- Q1: C - Lyman series is ultraviolet.
- Q2: B - 656.3 nm is $H\alpha$ line in Balmer series.
- Q3: A - Transition from $n=4$ to $n=2$ emits Balmer series photon.

Memory Aids

- Lyman: UV, $n=1$ final level.
- Balmer: Visible, $n=2$ final level.
- Paschen: IR, $n=3$ final level.

Absorption Spectrum

Key Points

- White light passing through gas shows dark lines at specific wavelengths.

- Dark lines match emission lines of the gas.
- Called absorption spectrum of the gas.

Solved Examples

- **Ex 1:** Why do dark lines appear in absorption spectrum?
Solution: Gas atoms absorb light at specific wavelengths, causing dark lines.

Practice Set

- Q1: What causes dark lines in absorption spectrum?
A. Emission of light
B. Absorption of specific wavelengths
C. Scattering of light
D. Reflection of light
- Q2: Absorption lines correspond to which of the following?
A. Random wavelengths
B. Emission lines of the gas
C. Continuous spectrum
D. Infrared radiation only
- Q3: Absorption spectrum is observed when white light passes through?
A. Vacuum
B. Gas
C. Solid
D. Liquid

Answer Key

- Q1: B - Gas absorbs specific wavelengths causing dark lines.
- Q2: B - Absorption lines match emission lines of the gas.
- Q3: B - Absorption spectrum occurs when light passes through gas.

Memory Aids

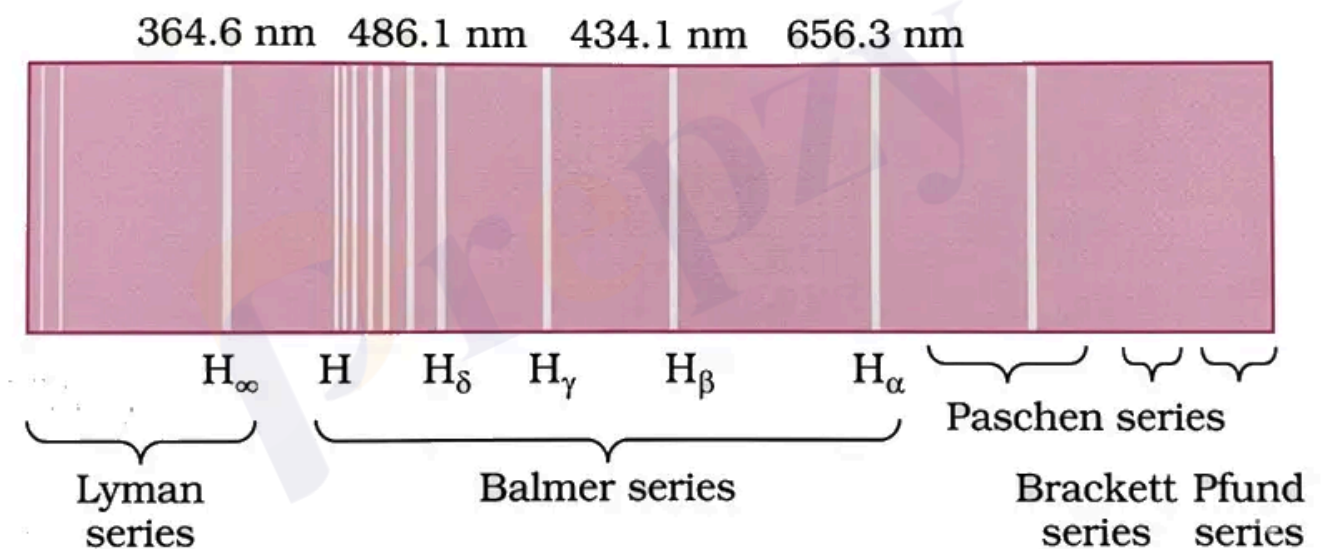
- Absorption = dark lines, emission = bright lines.

- Absorption lines match emission lines of same gas.

Spectral Series

Definition and Pattern

- Light frequencies emitted by elements show regular patterns called spectral series.
- Hydrogen has simplest spectrum with sets of lines called series.
- Balmer series observed in visible region (1885).
- Lines get closer and weaker as wavelength decreases.
- Balmer formula: $\frac{1}{\lambda} = R \left(\frac{1}{2^2} - \frac{1}{n^2} \right), n = 3, 4, 5, \dots$
- R = Rydberg constant = $1.097 \times 10^7 \text{ m}^{-1}$



Spectral Series of Hydrogen

- Lyman: $\frac{1}{\lambda} = R \left(\frac{1}{1^2} - \frac{1}{n^2} \right), n = 2, 3, 4, \dots$ (UV)
- Balmer: $\frac{1}{\lambda} = R \left(\frac{1}{2^2} - \frac{1}{n^2} \right), n = 3, 4, 5, \dots$ (Visible)
- Paschen: $\frac{1}{\lambda} = R \left(\frac{1}{3^2} - \frac{1}{n^2} \right), n = 4, 5, 6, \dots$ (IR)
- Brackett: $\frac{1}{\lambda} = R \left(\frac{1}{4^2} - \frac{1}{n^2} \right), n = 5, 6, 7, \dots$ (IR)
- Pfund: $\frac{1}{\lambda} = R \left(\frac{1}{5^2} - \frac{1}{n^2} \right), n = 6, 7, 8, \dots$ (IR)

Solved Examples

- **Ex 1:** Calculate wavelength for Balmer line with $n = 3$.

Solution: $\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 1.097 \times 10^7 \left(\frac{1}{4} - \frac{1}{9}\right) = 1.52 \times 10^6 \text{ m}^{-1}$

$\lambda = 6.58 \times 10^{-7} \text{ m} = 658 \text{ nm}$

Practice Set

- Q1: Which series corresponds to electron transition to $n = 1$?
A. Balmer
B. Lyman
C. Paschen
D. Brackett
- Q2: The Rydberg constant value is?
A. $1.097 \times 10^7 \text{ m}^{-1}$
B. 9.8 m/s^2
C. $3 \times 10^8 \text{ m/s}$
D. $6.63 \times 10^{-34} \text{ Js}$
- Q3: Wavelength of Paschen series line for transition $n = 5 \rightarrow n = 3$ is?
A. Calculate using $\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{5^2}\right)$
B. 656.3 nm
C. 486.1 nm
D. 91.2 nm

Answer Key

- Q1: B - Lyman series ends at $n = 1$.
- Q2: A - Rydberg constant is $1.097 \times 10^7 \text{ m}^{-1}$.
- Q3: A - Use formula $\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{5^2}\right)$ to calculate.

Memory Aids

- Series final levels: Lyman=1, Balmer=2, Paschen=3.
- Balmer visible, Lyman UV, others IR.
- Formula: $\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$.

Bohr Model

Bohr's Postulates

- Electrons revolve in fixed stable orbits without emitting energy.
- Angular momentum of electron in orbit: $mvr = n\frac{h}{2\pi}$, where $n = 1, 2, 3, \dots$
- Electron transitions between orbits emit or absorb energy as photons: $h\nu = E_i - E_f$.
- Energy quantization: electrons have discrete energy levels.

Postulates

- An electron in an atom could revolve in certain stable orbit without the emission of radiant energy.
- Electron revolves around the nucleus only in those orbits for which the angular momentum is an integral multiple of $h/2\pi$.
 $L = n(h/2\pi)$, $n = 1, 2, 3, \dots$



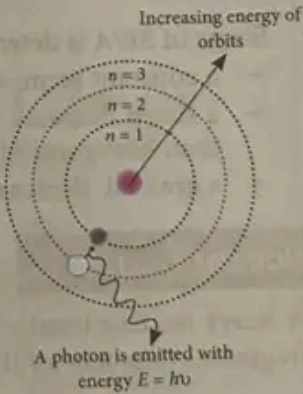
Stability of an Atom

Results :

- Velocity: $v_n = \frac{2\pi^2 ke^2}{nh}$, Radius: $r_n = \frac{n^2 h^2}{4\pi^2 kme^2}$
- Angular momentum: $mvr_n = \frac{nh}{2\pi}$

Energy of an Electron

- Each orbit has different energy, and electrons can move to a higher orbit by absorbing energy and drop to a lower orbit by emitting energy.
- The change in energy of the electron when it moves from one orbit to another is
 $\Delta E = E_i - E_f = h\nu$ where $h\nu$ is the energy of the observed or emitted photon.



Solved Examples

- **Ex 1:** Calculate angular momentum for $n = 2$: $L = 2 \times \frac{h}{2\pi} = \frac{2h}{2\pi}$.
- **Ex 2:** Energy of photon emitted when electron falls from $n = 3$ to $n = 2$: $h\nu = E_3 - E_2$.

Practice Set

- Q1: What is the angular momentum of electron in $n = 1$ orbit?
 - A) $\frac{h}{2\pi}$

- B) h
- C) $2h$
- D) Zero
- Q2: When electron jumps from higher to lower orbit, it:
 - A) Absorbs energy
 - B) Emits energy as photon
 - C) Remains same energy
 - D) Emits sound
- Q3: Electron in $n = 3$ orbit has angular momentum:
 - A) $3 \frac{h}{2\pi}$
 - B) $\frac{h}{2\pi}$
 - C) $2 \frac{h}{2\pi}$
 - D) Zero

Answer Key

- Q1: A - By Bohr's postulate, angular momentum is $n \frac{h}{2\pi}$, for $n = 1$ it is $\frac{h}{2\pi}$.
- Q2: B - Electron emits photon energy when falling to lower orbit.
- Q3: A - Angular momentum is integral multiple of $\frac{h}{2\pi}$, here $n = 3$.

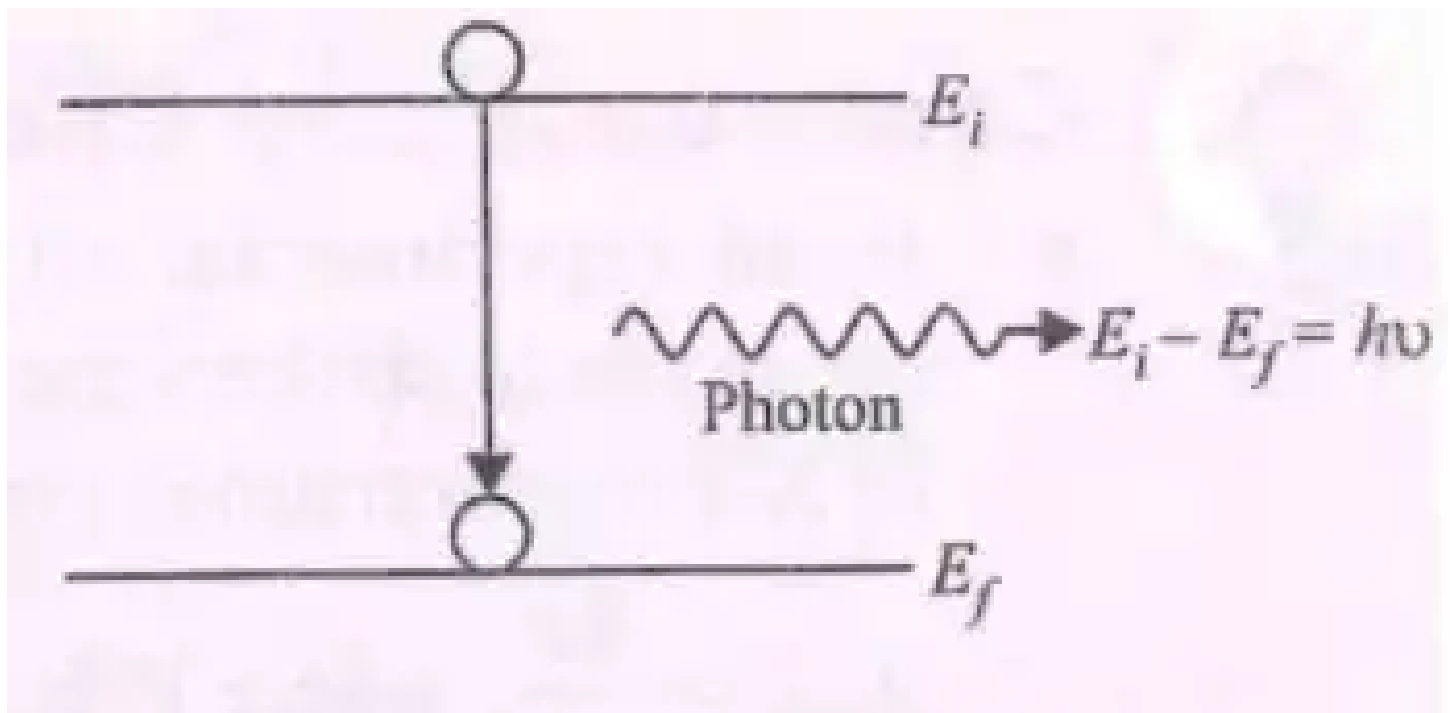
Memory Aids

- Angular momentum $= n \frac{h}{2\pi}$
- Energy emitted/absorbed $= h\nu = E_i - E_f$
- Electron orbits are stable and quantized

Energy Transitions

Electron Orbit Transitions

- Electron moves from higher energy E_i to lower E_f orbit emitting photon energy $h\nu = E_i - E_f$.
- Photon frequency ν relates to energy difference.
- Energy levels are discrete and quantized.



Solved Examples

- **Ex 1:** Calculate photon energy when electron falls from $E_3 = -1.51$ eV to $E_2 = -3.4$ eV:
 $h\nu = |-1.51 - (-3.4)| = 1.89$ eV.

Practice Set

- Q1: What happens when electron jumps from higher to lower orbit?
 - A) Absorbs photon
 - B) Emits photon
 - C) No energy change
 - D) Emits electron
- Q2: Energy of photon emitted equals:
 - A) $E_f - E_i$
 - B) $E_i - E_f$
 - C) $E_i + E_f$
 - D) Zero
- Q3: Electron falls from $n = 4$ to $n = 2$, energy emitted is:
 - A) $0.85 - 3.4$ eV
 - B) $3.4 - 0.85$ eV
 - C) $|-0.85 - (-3.4)|$ eV
 - D) 0 eV

Answer Key

- Q1: B - Electron emits photon when moving to lower orbit.
- Q2: B - Photon energy equals difference $E_i - E_f$.
- Q3: C - Energy emitted is difference of energy levels.

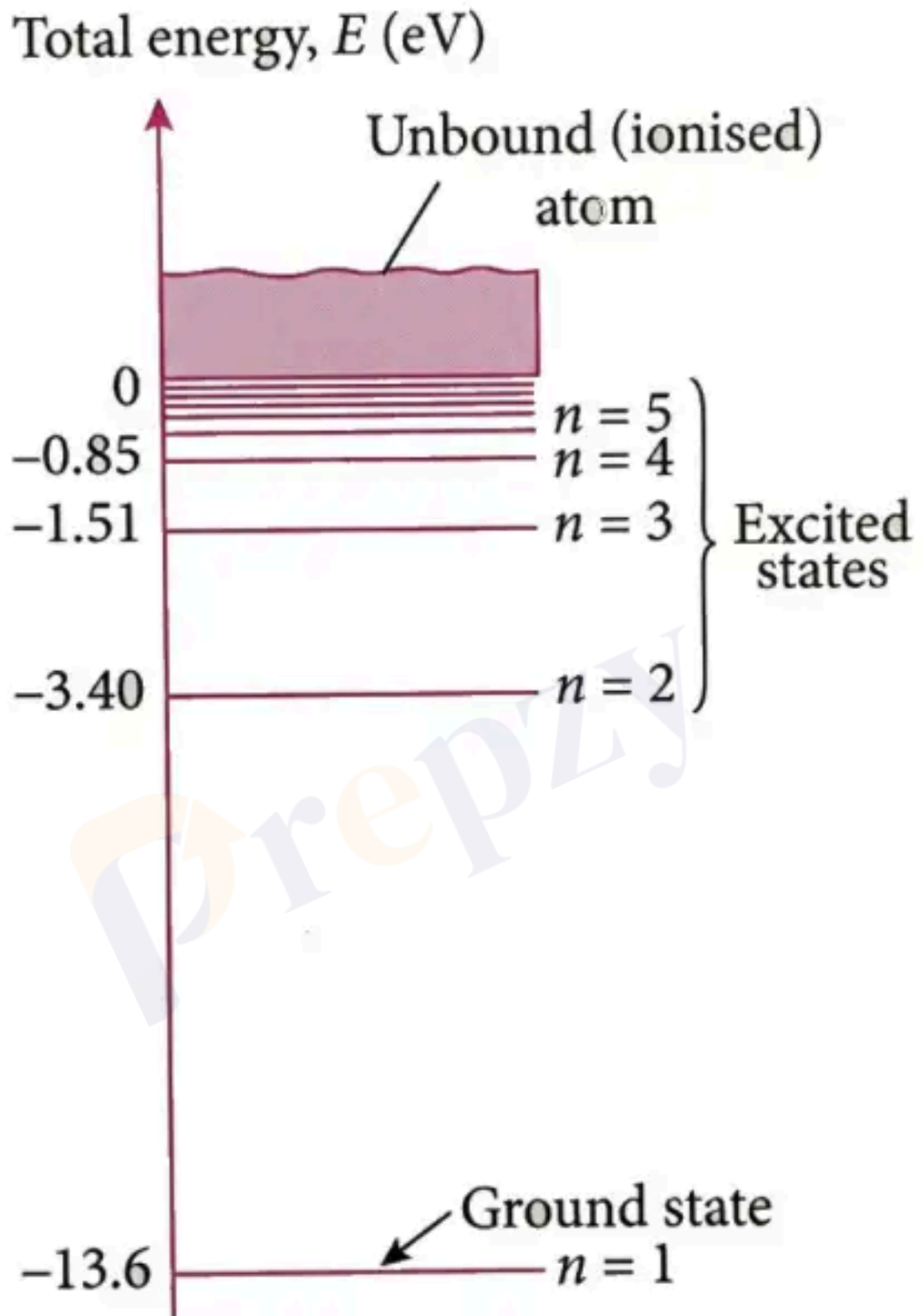
Memory Aids

- Energy emitted = $h\nu = E_i - E_f$
- Photon frequency ν proportional to energy difference
- Electron jumps cause emission or absorption

Orbit Radius & Energy Levels

Formulas and Energy Levels

- Centripetal force: $\frac{mv_n^2}{r_n} = \frac{1}{4\pi\epsilon_0} \frac{eZe}{r_n^2}$
- Angular momentum: $L_n = mv_n r_n = n \frac{h}{2\pi}$
- Radius of nth orbit: $r_n = \frac{n^2 h^2 \epsilon_0}{\pi m e^2 Z} = 0.53 \frac{n^2}{Z} \text{ \AA}$
- Energy levels: $E_n = -\frac{13.6}{n^2} \text{ eV}$ for hydrogen atom ($Z=1$)
- Ground state ($n=1$) energy: -13.6 eV , higher states have less negative energy.



Solved Examples

- **Ex 1:** Calculate radius for $n = 2$ in hydrogen: $r_2 = 0.53 \times \frac{2^2}{1} = 2.12 \text{ \AA}$.
- **Ex 2:** Energy at $n = 3$: $E_3 = -\frac{13.6}{3^2} = -1.51 \text{ eV}$.

Practice Set

- Q1: Radius of $n = 1$ orbit in hydrogen is:
 - A) $0.53 \text{ \AA}$
 - B) $1.06 \text{ \AA}$
 - C) $2.12 \text{ \AA}$
 - D) $0.26 \text{ \AA}$
- Q2: Energy of electron in $n = 4$ orbit is:
 - A) -0.85 eV
 - B) -3.4 eV
 - C) -13.6 eV
 - D) -1.51 eV
- Q3: Radius r_n is proportional to:
 - A) n
 - B) n^2/Z
 - C) Z/n^2
 - D) $1/n$

Answer Key

- Q1: A - Bohr radius for hydrogen ground state is $0.53 \text{ \AA}$.
- Q2: A - Energy at $n = 4$ is $-13.6/4^2 = -0.85 \text{ eV}$.
- Q3: B - Radius varies as n^2/Z .

Memory Aids

- Radius $r_n = 0.53 \frac{n^2}{Z} \text{ \AA}$
- Energy $E_n = -\frac{13.6}{n^2} \text{ eV}$
- Ground state energy = -13.6 eV

Energy Quantization

Concept of Energy Quantization

- Electron energy in orbits is discrete, not continuous.
- Only certain energy values allowed for electron in atom.
- Explains stability of atoms and spectral lines.

Solved Examples

- **Ex 1:** Electron cannot have energy between -13.6 eV and -3.4 eV in hydrogen atom.

Practice Set

- Q1: Energy levels in atom are:
 - A) Continuous
 - B) Discrete
 - C) Random
 - D) Infinite
- Q2: Energy quantization explains:
 - A) Atomic stability
 - B) Continuous spectra
 - C) Electron emission only
 - D) Nuclear reactions
- Q3: Electron energy between allowed levels is:
 - A) Allowed
 - B) Forbidden
 - C) Random
 - D) Infinite

Answer Key

- Q1: B - Energy levels are discrete, not continuous.
- Q2: A - Quantization explains atomic stability and spectra.
- Q3: B - Electron cannot have energy between allowed levels.

Memory Aids

- Energy levels = discrete values only
- Electron jumps cause photon emission/absorption
- Quantization explains atomic spectra

Line Spectra of Hydrogen

Energy Transitions and Rydberg Formula

- Electron transitions from higher energy level n_i to lower n_f emit photon of frequency ν_{if} .
- Energy difference: $h\nu_{if} = E_{n_i} - E_{n_f} = \frac{me^4}{8\epsilon_0^2 h^2} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$.
- Frequency formula: $\nu_{if} = \frac{me^4}{8\epsilon_0^2 h^3} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$.
- Discrete frequencies arise due to integer quantum numbers n_i, n_f .
- Series: Lyman ($n_f = 1$), Balmer ($n_f = 2$), Paschen ($n_f = 3$), Brackett ($n_f = 4$), Pfund ($n_f = 5$).

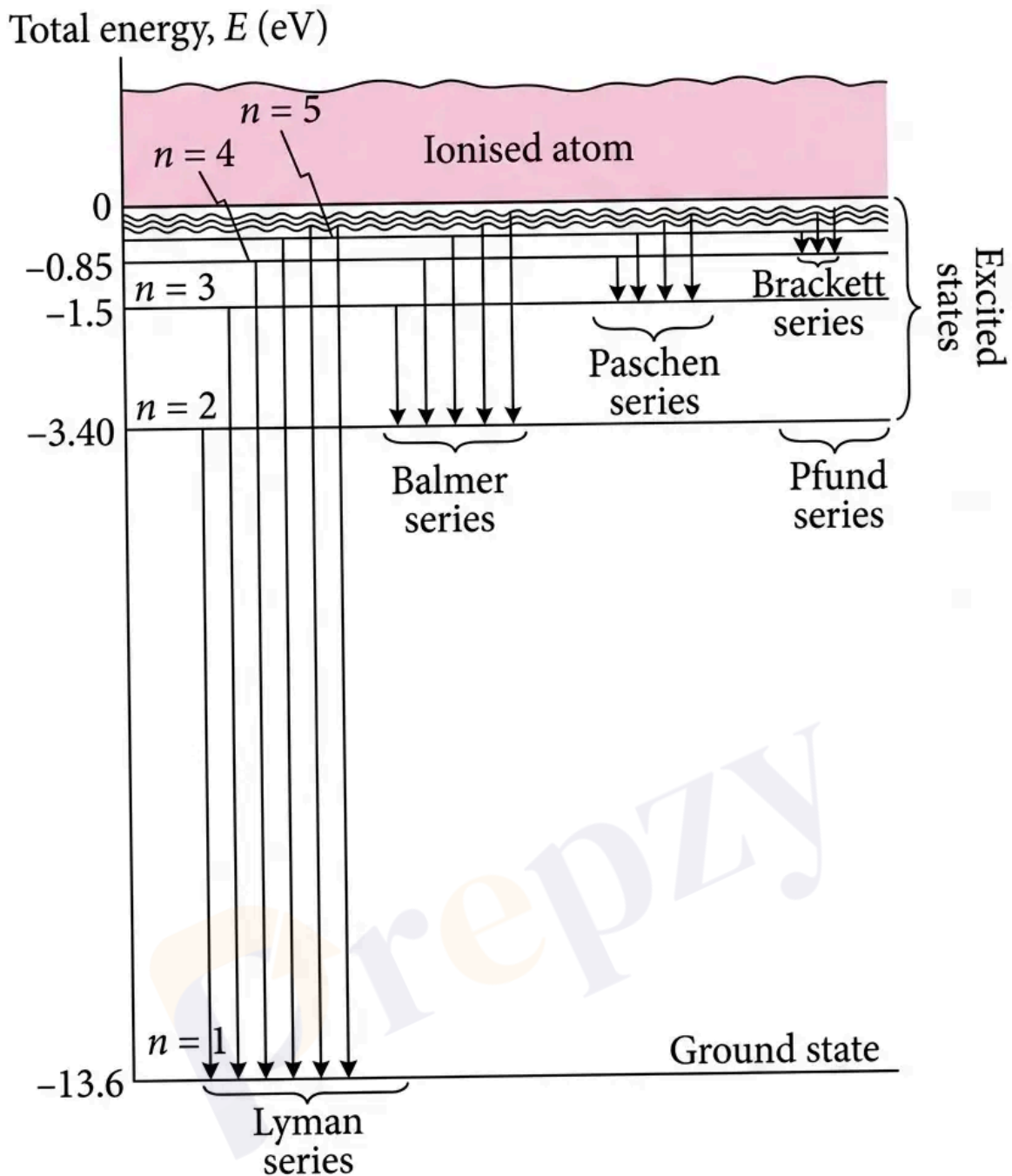


Figure : Line spectra originate in transitions between energy levels.

Solved Examples

- **Ex 1:** Calculate frequency for transition $n_i = 3$ to $n_f = 2$ in hydrogen.
 - Use $\nu = \frac{me^4}{8\epsilon_0^2 h^3} \left(\frac{1}{2^2} - \frac{1}{3^2} \right)$.
 - Substitute constants and compute ν .

Practice Set

- Q1: Which series corresponds to transitions ending at $n = 2$?

- A. Lyman
- B. Balmer
- C. Paschen
- D. Brackett
- Q2: The frequency of emitted photon for transition $n_i = 4$ to $n_f = 1$ is proportional to?
 - A. $\frac{1}{1^2} - \frac{1}{4^2}$
 - B. $\frac{1}{4^2} - \frac{1}{1^2}$
 - C. $\frac{1}{2^2} - \frac{1}{3^2}$
 - D. $\frac{1}{3^2} - \frac{1}{2^2}$
- Q3: For hydrogen, which series involves transitions to $n = 3$?
 - A. Lyman
 - B. Balmer
 - C. Paschen
 - D. Pfund

Answer Key

- Q1: B. Balmer - Transitions ending at $n = 2$ form Balmer series.
- Q2: A. $\frac{1}{1^2} - \frac{1}{4^2}$ - Frequency depends on difference $\frac{1}{n_f^2} - \frac{1}{n_i^2}$.
- Q3: C. Paschen - Transitions to $n = 3$ define Paschen series.

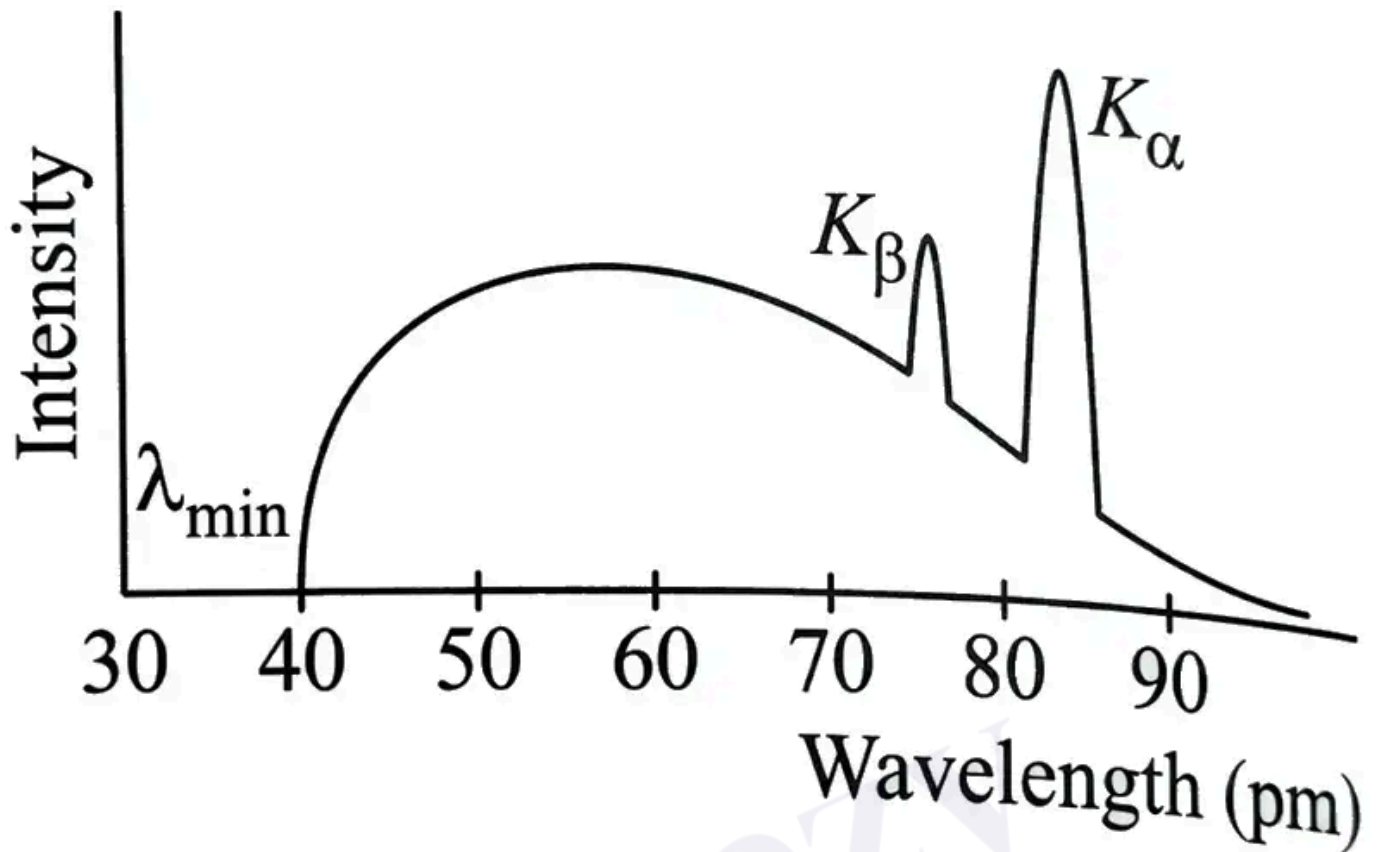
Memory Aids

- Rydberg formula key: $\nu = R \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$.
- Series order: Lyman (1), Balmer (2), Paschen (3), Brackett (4), Pfund (5).
- Higher n_i to lower n_f emits photon.

Continuous and Characteristic X-rays

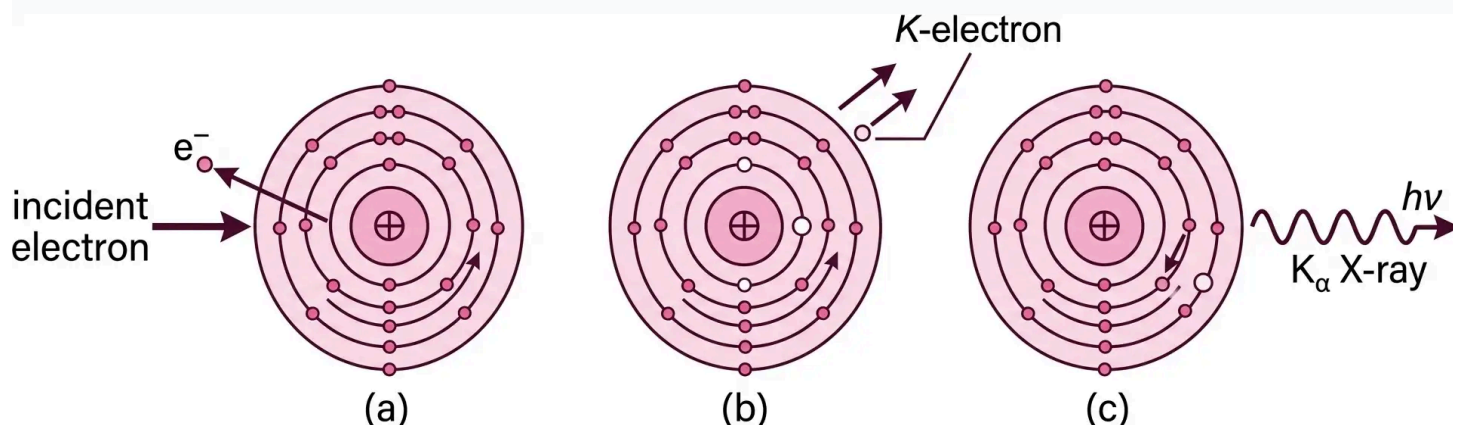
X-ray Production and Cutoff Wavelength

- Electrons accelerated by potential V strike metal target producing X-rays.
- Cutoff wavelength: $\lambda_{min} = \frac{hc}{eV}$ or $\lambda_{min}(nm) = \frac{1242}{V(volt)}$.
- λ_{min} depends only on accelerating voltage, not target material.



Characteristic X-rays and Electron Transitions

- Incident electron knocks out inner shell electron (K-shell), creating vacancy.
- Electron from higher shell fills vacancy, emitting characteristic X-ray photon.
- Common X-ray lines: K_{α} , K_{β} , L_{α} .
- Wavelength formulas:
 - $\lambda = \frac{hc}{E_K - E_L}$ for K_{α} .
 - $\lambda = \frac{hc}{E_K - E_M}$ for K_{β} .
 - $\lambda = \frac{hc}{E_L - E_M}$ for L_{α} .



Solved Examples

- **Ex 1:** Calculate cutoff wavelength for $V = 2000$ V.
 - Use $\lambda_{min} = \frac{1242}{V} = \frac{1242}{2000} = 0.621$ nm.

Practice Set

- Q1: Cutoff wavelength λ_{min} depends on?
 - A. Target material
 - B. Accelerating voltage
 - C. Electron mass
 - D. Atomic number
- Q2: Characteristic X-rays are produced when?
 - A. Electron slows down in target
 - B. Electron knocks out inner shell electron
 - C. Electron accelerates in vacuum
 - D. Electron collides with outer shell electron
- Q3: Wavelength of K_{α} X-ray is given by?
 - A. $\frac{hc}{E_K - E_L}$
 - B. $\frac{hc}{E_L - E_M}$
 - C. $\frac{hc}{E_K - E_M}$
 - D. $\frac{hc}{E_M - E_N}$

Answer Key

- Q1: B. Accelerating voltage - λ_{min} depends only on V .
- Q2: B. Electron knocks out inner shell electron - characteristic X-rays arise from inner shell vacancies.
- Q3: A. $\frac{hc}{E_K - E_L}$ - Formula for K_{α} wavelength.

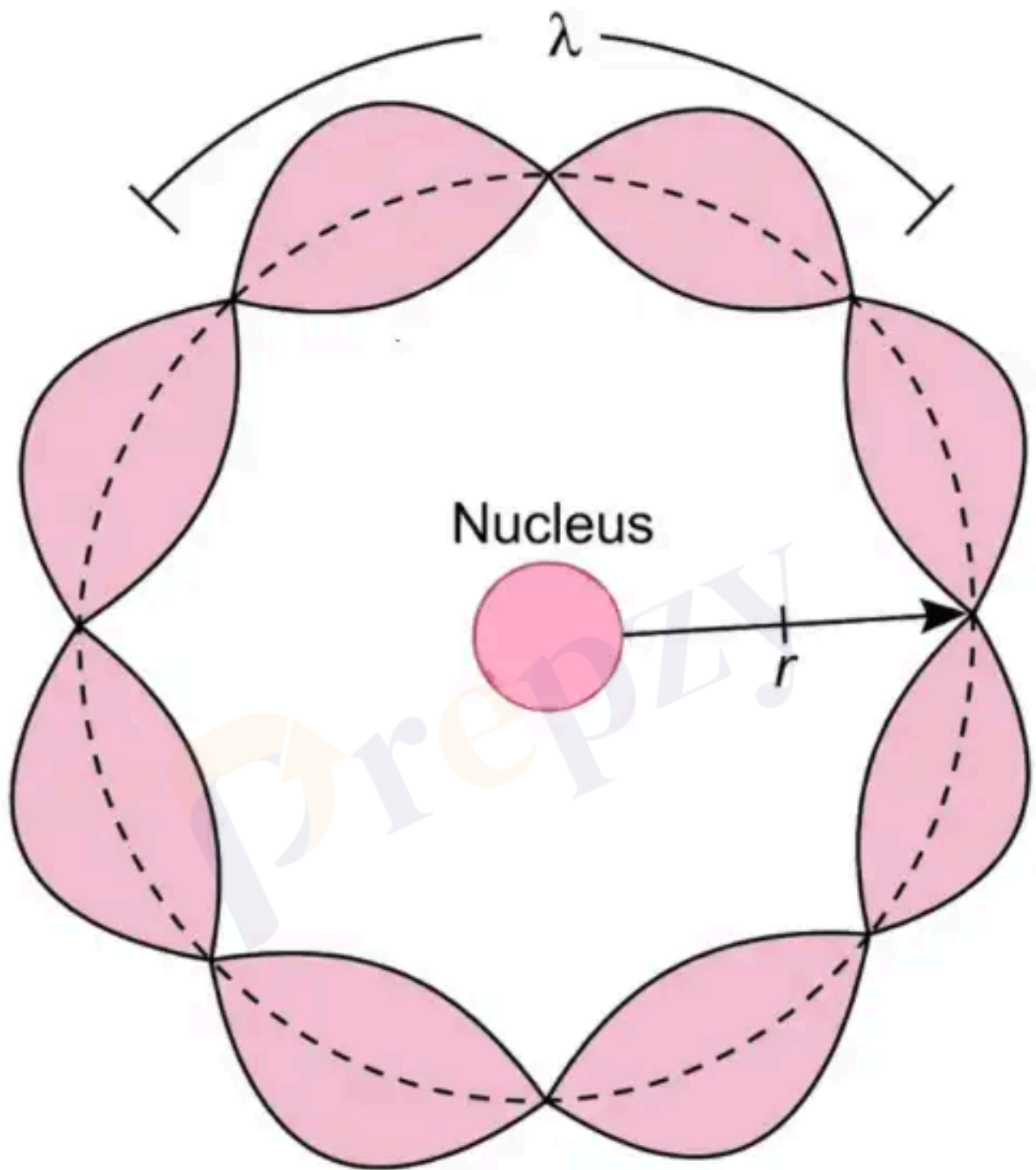
Memory Aids

- $\lambda_{min} = \frac{1242}{V}$ nm for cutoff wavelength.
- Characteristic X-rays: electron transitions to inner shell vacancies.
- K, L, M shells correspond to X-ray lines K_{α} , K_{β} , L_{α} .

De-Broglie Explanation of Bohr's Second Postulate

Concepts

- Bohr's second postulate: Angular momentum quantised as $L_n = \frac{nh}{2\pi}$, $n = 1, 2, 3, \dots$
- De Broglie hypothesis: Electron behaves as a particle wave in circular orbit.
- Standing waves form when orbit circumference equals integral multiples of wavelength:
$$n\lambda = 2\pi r_n$$
- De Broglie wavelength: $\lambda = \frac{h}{mv_n}$
- From $n\lambda = 2\pi r_n$ and $\lambda = \frac{h}{mv_n}$, get $mv_n r_n = \frac{nh}{2\pi}$ (Bohr's quantisation)
- Only orbits with standing waves (resonant condition) are allowed.



Solved Examples

- **Ex 1:** Find angular momentum for $n = 2$ orbit.
 - Use $L_n = \frac{nh}{2\pi}$
 - $L_2 = \frac{2h}{2\pi} = \frac{h}{\pi}$

Practice Set

- Q1: What is the de Broglie wavelength formula?
 - A) $\lambda = \frac{h}{mv}$
 - B) $\lambda = mv$
 - C) $\lambda = \frac{mv}{h}$
 - D) $\lambda = hv$
- Q2: For an electron in 3rd orbit, circumference is:
 - A) $2\pi r_3$
 - B) 3λ
 - C) Both A and B
 - D) None
- Q3: Angular momentum for $n = 1$ orbit is:
 - A) $\frac{h}{2\pi}$
 - B) h
 - C) $2\pi h$
 - D) $\frac{h}{\pi}$

Answer Key

- Q1: A - De Broglie wavelength is $\lambda = \frac{h}{mv}$.
- Q2: C - Circumference equals $n\lambda$, so both are equal.
- Q3: A - For $n = 1$, $L = \frac{h}{2\pi}$.

Memory Aids

- Angular momentum quantised as multiples of $\frac{h}{2\pi}$.
- Standing waves form only if orbit circumference = integral $\times$ wavelength.
- De Broglie wavelength $\lambda = \frac{h}{mv}$.

Limitations of Bohr's Theory

Concepts

- Valid only for single-electron atoms (H, He^+ , Li^{2+}).
- Cannot explain fine structure of spectral lines.
- Fails to explain molecular bonding and stability.
- Cannot fully explain Zeeman and Stark effects.
- Does not explain electron distribution in orbits.
- Ignores electron-electron interactions in multi-electron atoms.
- Cannot explain intensity variations in hydrogen emission spectrum.

Solved Examples

- **Ex 1:** Identify if Bohr's theory applies to He atom.
 - He has 2 electrons → multi-electron atom.
 - Bohr's theory fails here.

Practice Set

- Q1: Bohr's theory is valid for:
 - A) Multi-electron atoms
 - B) Single-electron atoms
 - C) Molecules
 - D) None
- Q2: Which effect is NOT explained by Bohr's theory?
 - A) Zeeman effect
 - B) Stark effect
 - C) Fine structure
 - D) All of the above
- Q3: Bohr's theory ignores:
 - A) Electron-nucleus force
 - B) Electron-electron force
 - C) Nuclear force
 - D) Gravitational force

Answer Key

- Q1: B - Valid only for single-electron atoms.
- Q2: D - Bohr's theory fails to explain all listed effects.
- Q3: B - Ignores electron-electron interactions.

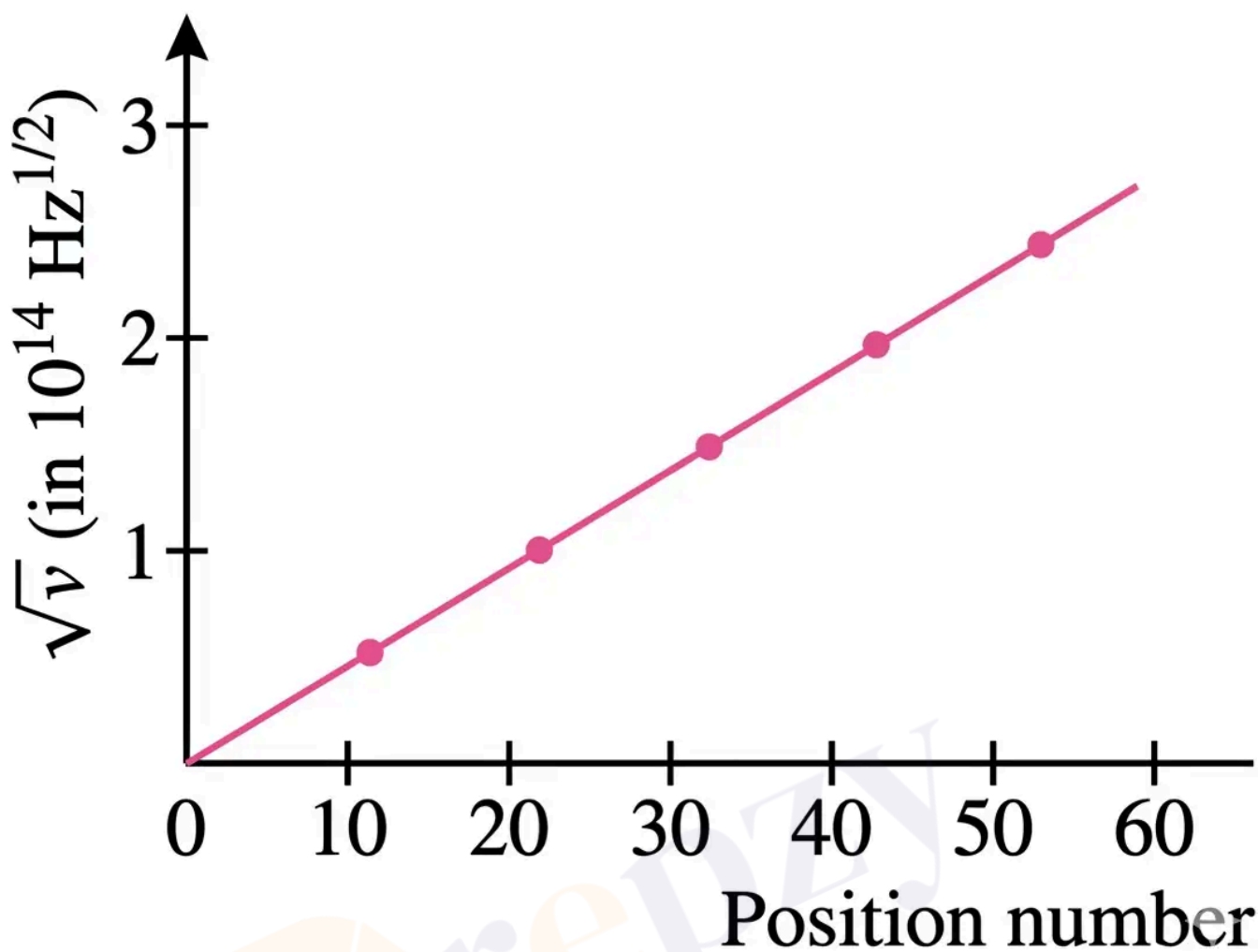
Memory Aids

- Bohr's model = single-electron only.
- Fails for fine structure, Zeeman, Stark effects.
- Ignores electron-electron repulsion.

Moseley's Law

Concepts

- X-ray frequency relation: $\sqrt{\nu} = a(Z - b)$
- Constants a, b ; $b \approx 1$
- Bohr's model fits if Z replaced by $Z - b$
- Energy released for $n = 2 \rightarrow n = 1$: $\Delta E = h\nu = Rhc(Z - b)^2(1 - \frac{1}{4})$



Solved Examples

- **Ex 1:** Calculate $\sqrt{\nu}$ for $Z = 30, b = 1, a = \text{constant}$.
 - $\sqrt{\nu} = a(Z - b) = a(30 - 1) = 29a$

Practice Set

- Q1: Moseley's law relates $\sqrt{\nu}$ to:
 - A) Atomic number Z
 - B) $Z - b$
 - C) Wavelength λ
 - D) Energy E
- Q2: In Moseley's law, constant b is approximately:
 - A) 0
 - B) 1

- C) 2
- D) 3
- Q3: Energy released in transition $n = 2 \rightarrow n = 1$ depends on:
 - A) $(Z - b)^2$
 - B) Z^2
 - C) b^2
 - D) None

Answer Key

- Q1: B - Frequency depends on effective atomic number $Z - b$.
- Q2: B - Constant $b \approx 1$.
- Q3: A - Energy depends on square of $Z - b$.

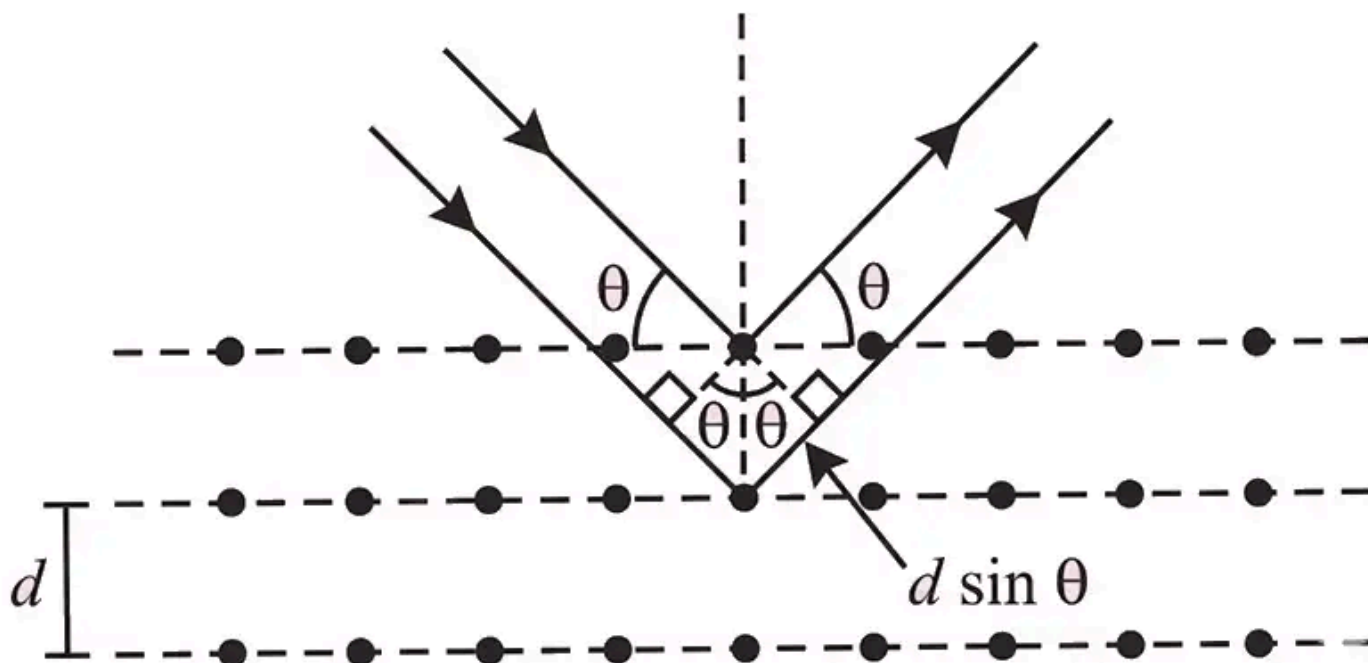
Memory Aids

- $\sqrt{\nu}$ linear with $Z - b$.
- $b \approx 1$ adjusts atomic number.
- Energy $\propto (Z - b)^2$.

Bragg's Law

Concepts

- Condition for X-ray diffraction: $2d \sin \theta = n\lambda$
- n : order of spectrum (1,2,3...)
- d : distance between crystal planes (m)
- θ : glancing angle
- λ : X-ray wavelength (m)
- Path difference between waves: $d \sin \theta$
- Constructive interference occurs when path difference = integral multiples of wavelength.



Solved Examples

- **Ex 1:** Calculate θ for first order diffraction with $d = 0.2 \text{ nm}$, $\lambda = 0.1 \text{ nm}$.
 - Use $2d \sin \theta = n\lambda$
 - $2 \times 0.2 \times \sin \theta = 1 \times 0.1 \Rightarrow \sin \theta = 0.25$
 - $\theta = \sin^{-1} 0.25 = 14.48^\circ$

Practice Set

- Q1: Bragg's law formula is:
 - A) $2d \sin \theta = n\lambda$
 - B) $d = 2\lambda \sin \theta$
 - C) $n = 2d \sin \theta$
 - D) $\lambda = 2d \sin \theta$
- Q2: In Bragg's law, d is:
 - A) Distance between atoms
 - B) Distance between adjacent planes
 - C) Wavelength
 - D) Glancing angle
- Q3: For $n = 2$, $d = 0.3 \text{ nm}$, $\lambda = 0.15 \text{ nm}$, find θ .
 - A) 30°
 - B) 45°
 - C) 60°

- D) 90°

Answer Key

- Q1: A - Bragg's law is $2d \sin \theta = n\lambda$.
- Q2: B - d is distance between adjacent crystal planes.
- Q3: A - $2 \times 0.3 \times \sin \theta = 2 \times 0.15 \Rightarrow \sin \theta = 0.5 \Rightarrow \theta = 30^\circ$.

Memory Aids

- Bragg's law: $2d \sin \theta = n\lambda$.
- d = interplanar spacing.
- Constructive interference requires path difference = integral multiples of wavelength.