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Introduction

Overview

- Dalton's atomic theory (1808): Atom is the ultimate particle of matter.
- Explains laws: conservation of mass, constant composition, multiple proportions.
- Fails to explain electrical charging of rubbed substances (glass, ebonite).

Solved Examples

- **Ex 1:** Explain why Dalton's atomic theory could not explain electrical charging of rubbed substances.
 - Dalton's theory did not consider subatomic particles or electric charge.

Practice Set

- **Q1:** Dalton's atomic theory regards the atom as:
 - A. Divisible particle
 - B. Ultimate particle
 - C. Made of electrons
 - D. Electrically charged particle
- **Q2:** Which law is NOT explained by Dalton's atomic theory?
 - A. Law of conservation of mass
 - B. Law of constant composition
 - C. Law of multiple proportions
 - D. Electrical charging of rubbed substances
- **Q3:** Dalton's atomic theory was proposed in:
 - A. 1908
 - B. 1808
 - C. 2008
 - D. 1708

Answer Key

- **Q1:** B - Atom is considered the ultimate particle in Dalton's theory.

- **Q2:** D - Dalton's theory could not explain electrical charging phenomena.
- **Q3:** B - Dalton proposed the atomic theory in 1808.

Memory Aids

- Dalton's theory = Atom is indivisible and ultimate particle.
- Fails for electrical phenomena → leads to discovery of subatomic particles.

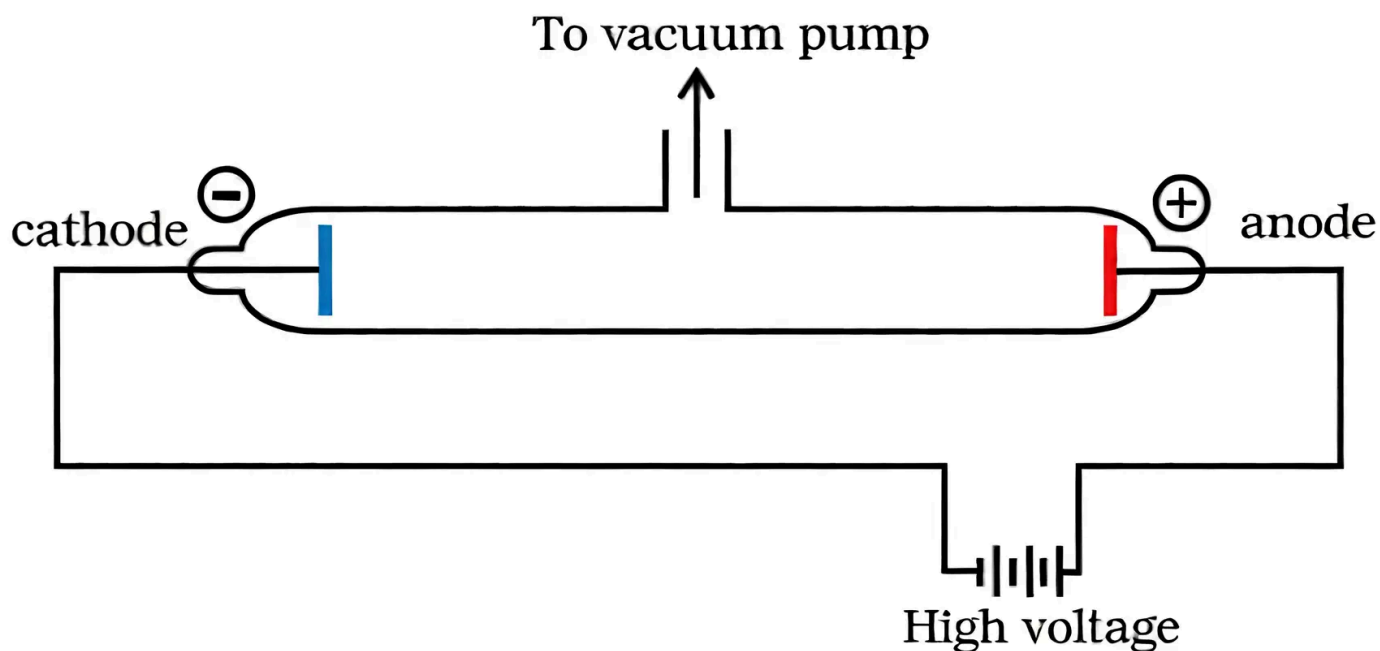
Discovery of Sub-atomic Particles

Fundamental Particles

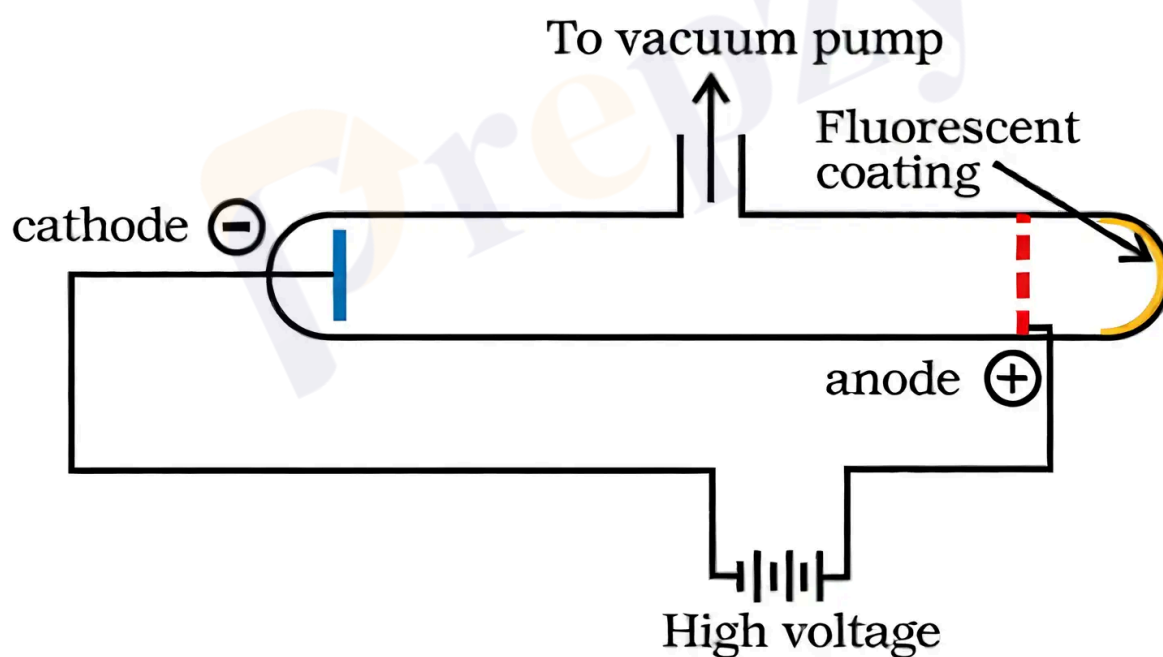
- Atoms consist of electrons, protons, and neutrons.
- These are called fundamental particles.

Discovery of Electron

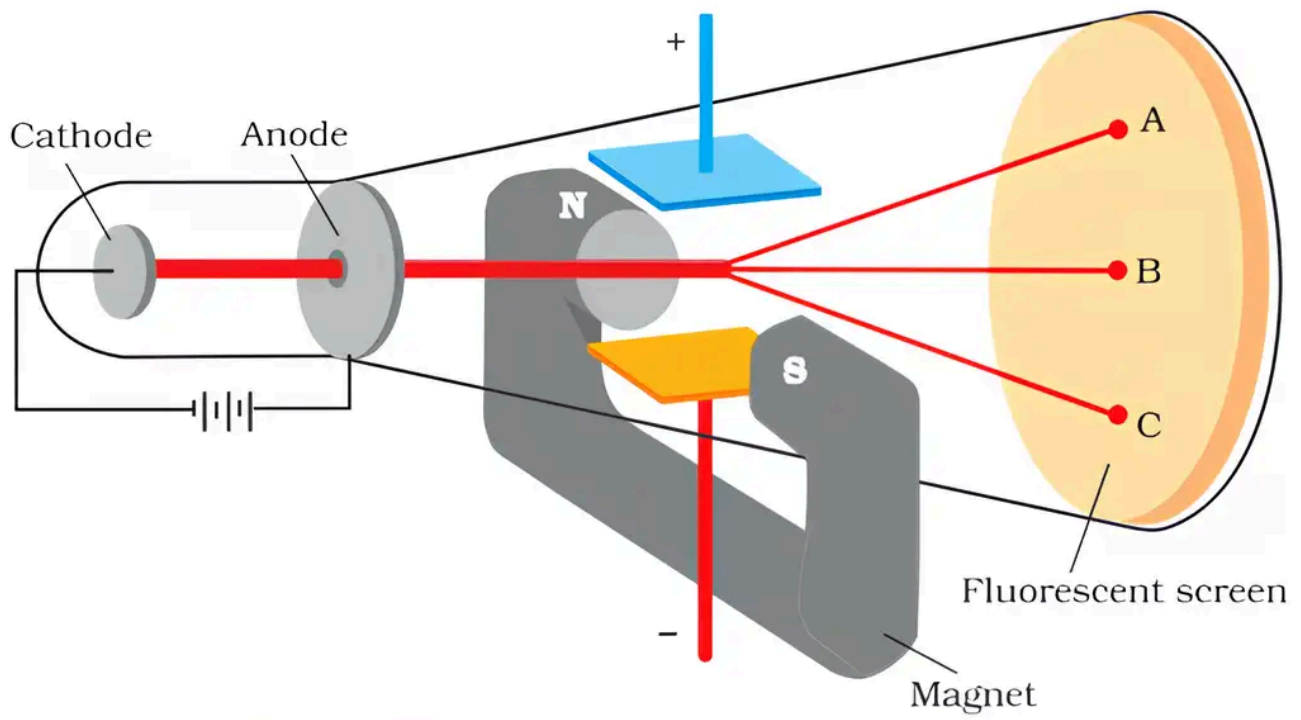
- Discovered by J.J. Thomson using cathode ray discharge tube.
- Cathode rays: stream of negatively charged particles (electrons) from cathode to anode.
- Visible via fluorescent coating glow when rays strike it.
- Characteristics:
 - Travel in straight lines without electric/magnetic field.
 - Deflected by electric/magnetic fields like negatively charged particles.
 - Independent of electrode material and gas type.
- Charge to mass ratio (e/m) = $1.758820 \times 10^{11} \text{ C kg}^{-1}$ (Thomson).
- Charge on electron = $-1.6 \times 10^{-19} \text{ C}$ (Millikan oil drop experiment).
- Mass of electron = $9.1094 \times 10^{-31} \text{ kg}$.



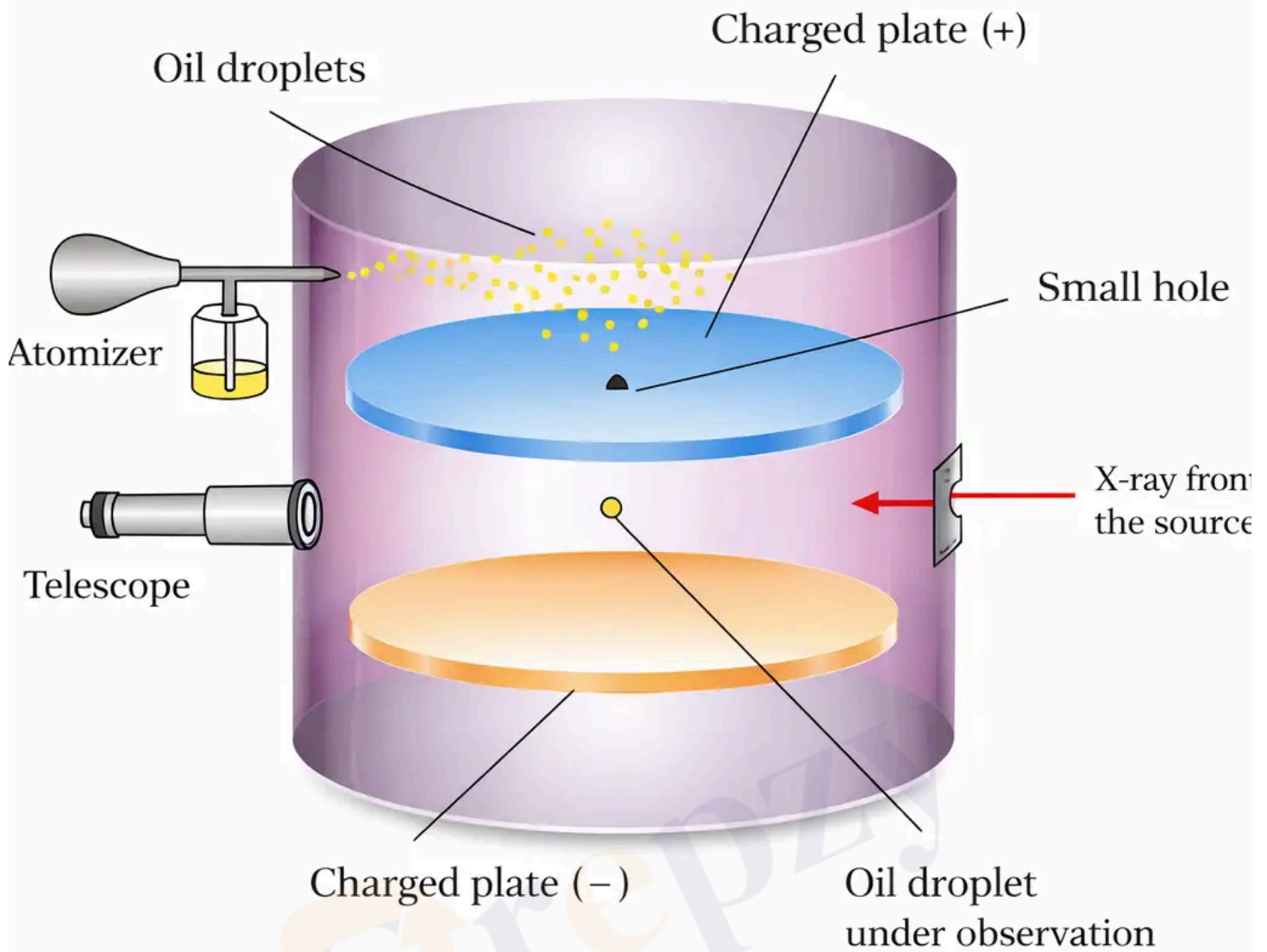
A cathode ray discharge tube



A cathode ray discharge tube with perforated anode



The apparatus to determine the charge to the mass ratio of electron



The Millikan oil drop apparatus for measuring charge 'e'

Discovery of Protons and Canal Rays

- Protons discovered by E. Goldstein using modified cathode ray tube.
- Positively charged particles depend on gas type in tube.
- Charge to mass ratio varies with gas origin.
- Some carry multiple units of fundamental charge.
- Behavior opposite to cathode rays in electric/magnetic fields.

Discovery of Neutrons

- Discovered by James Chadwick by bombarding beryllium with α -particles.
- Neutrons are neutral, mass slightly greater than protons.

Properties of Fundamental Particles

- Electron (e): charge -1 , mass 9.1094×10^{-31} kg, relative charge -1 .
- Proton (p): charge $+1$, mass $\sim 1.6726 \times 10^{-27}$ kg, relative charge $+1$.
- Neutron (n): charge 0 , mass $\sim 1.6749 \times 10^{-27}$ kg, relative charge 0 .

Name	Symbol	Absolute charge (C)	Relative charge	Mass (kg)	Mass (u)	Approx. mass (u)
Electron	e	-1.6022×10^{-19}	-1	9.10939×10^{-31}	0.00054	0
Proton	p	$+1.6022 \times 10^{-19}$	$+1$	1.67262×10^{-27}	1.00727	1
Neutron	n	0	0	1.67493×10^{-27}	1.00867	1

Solved Examples

- **Ex 1:** Calculate the mass of an electron given charge $e = 1.6022 \times 10^{-19}$ C and $e/m = 1.758820 \times 10^{11}$ C kg $^{-1}$.
- Mass = $e / (e/m) = 9.1094 \times 10^{-31}$ kg.

Practice Set

- **Q1:** Cathode rays consist of:
A. Protons
B. Neutrons
C. Electrons
D. Alpha particles

- **Q2:** Charge to mass ratio of electron determined by:
A. Millikan
B. Thomson
C. Chadwick
D. Goldstein
- **Q3:** Neutrons were discovered by:
A. J.J. Thomson
B. E. Goldstein
C. James Chadwick
D. R.A. Millikan

Answer Key

- **Q1:** C - Cathode rays are streams of electrons.
- **Q2:** B - Thomson measured e/m ratio using cathode ray tube.
- **Q3:** C - Chadwick discovered neutrons by α -particle bombardment.

Memory Aids

- Electron: negative charge, discovered by Thomson.
- Proton: positive charge, discovered by Goldstein.
- Neutron: neutral, discovered by Chadwick.
- Charge to mass ratio (e/m) key for electron identification.

Atomic Models

Evolution of Atomic Models

- Dalton (1808): Atoms of each element are unique and indivisible.
- Thomson (1898): Atom is a positive sphere with embedded electrons (raisin pudding model).
- Rutherford (1911): Atom has a small, dense, positively charged nucleus; most of atom is empty space.
- Bohr (1913): Electrons move in fixed orbits around nucleus; explained atomic spectra.
- Schrödinger (1926): Electrons form a cloud; exact position uncertain; wave equation predicts electron probability.

History of Atomic Models

John Dalton (1808)

He recognised atoms of a particular element differ from other elements.



J. J. Thomson (1898) (positive and negative charges)

He recognised electrons as components of atoms.



Rutherford (1911) (the nucleus)

He realised positive charge was localised in the nucleus of an atom.



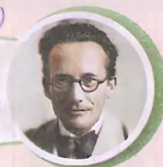
Bohr (1913) (energy levels)

He proposed stable electron orbits and explained the emission spectra of some elements.



Erwin Schrödinger (1926) (electron cloud model)

He proposed electrons don't move around the nucleus in orbits but in clouds where their position is uncertain.



Wave mechanical model :

Schrödinger equation,

$$\frac{d^2\psi}{dx^2} + \frac{d^2\psi}{dy^2} + \frac{d^2\psi}{dz^2} + \frac{8\pi^2m}{h^2}(E - V)\psi = 0$$

The solution of the equation provides data which enables us to calculate the probability of finding an electron of specific energy.

Solved Examples

- Ex 1: Identify the model describing electrons as embedded in a positive sphere.
Answer: Thomson's raisin pudding model.
- Ex 2: Which model introduced the concept of a nucleus?
Answer: Rutherford's nuclear model.

Practice Set

- Q1: Who proposed the electron cloud model?
A. Dalton
B. Thomson
C. Schrödinger
D. Bohr
- Q2: Rutherford's experiment showed that most of the atom is:
A. Solid
B. Empty space
C. Uniformly positive
D. Negatively charged
- Q3: Which model explains the stability of atom by fixed electron orbits?
A. Dalton
B. Thomson

C. Bohr
D. Schrödinger

Answer Key

- Q1: C - Schrödinger proposed the electron cloud model where electron position is probabilistic.
- Q2: B - Rutherford's scattering showed most alpha particles passed through, indicating empty space.
- Q3: C - Bohr's model introduced fixed orbits explaining atomic stability and spectra.

Memory Aids

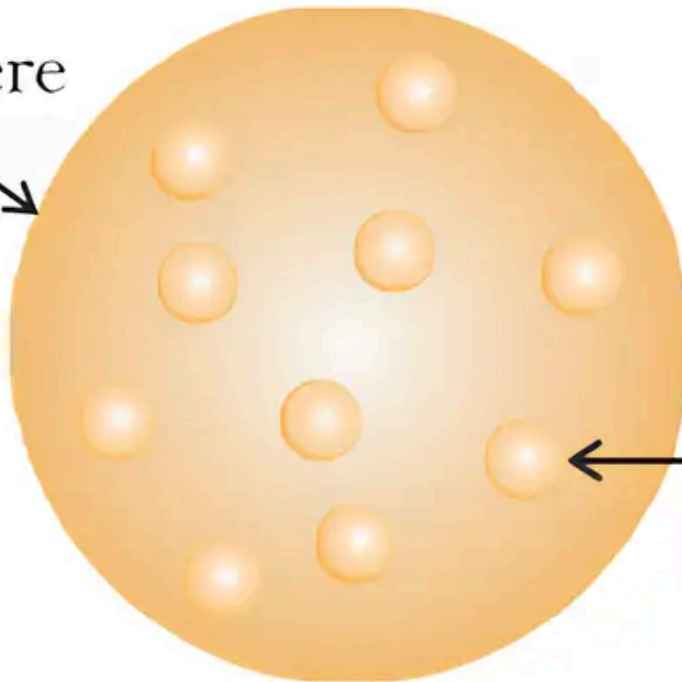
- Dalton: "Distinct atoms"
- Thomson: "Raisin pudding" = electrons in positive sphere
- Rutherford: "Nucleus discovery"
- Bohr: "Fixed orbits"
- Schrödinger: "Electron cloud"

Thomson Model

Key Features

- Atom is a sphere of uniformly distributed positive charge.
- Electrons (negative charges) are embedded like raisins in pudding.
- Mass is uniformly distributed over the atom.
- Explains overall electrical neutrality of atom.
- Fails to explain later experimental results.

Positive sphere



Electron

Thomson model of atom

Solved Examples

- Ex 1: Describe the charge distribution in Thomson's model.
Answer: Positive charge uniformly spread; electrons embedded within.

Practice Set

- Q1: Thomson's model compares electrons to:
 - A. Planets
 - B. Raisins in pudding
 - C. Orbiting satellites
 - D. Nucleus
- Q2: What is assumed about the mass distribution in Thomson's model?
 - A. Concentrated in nucleus
 - B. Uniformly distributed
 - C. Only in electrons
 - D. Only in positive sphere
- Q3: Thomson's model fails to explain:
 - A. Atom's neutrality
 - B. Electron discovery

- C. Later experimental results
- D. Presence of nucleus

Answer Key

- Q1: B - Electrons are like raisins embedded in positive pudding.
- Q2: B - Mass is assumed uniformly distributed over the atom.
- Q3: C - Model inconsistent with later experiments like Rutherford's.

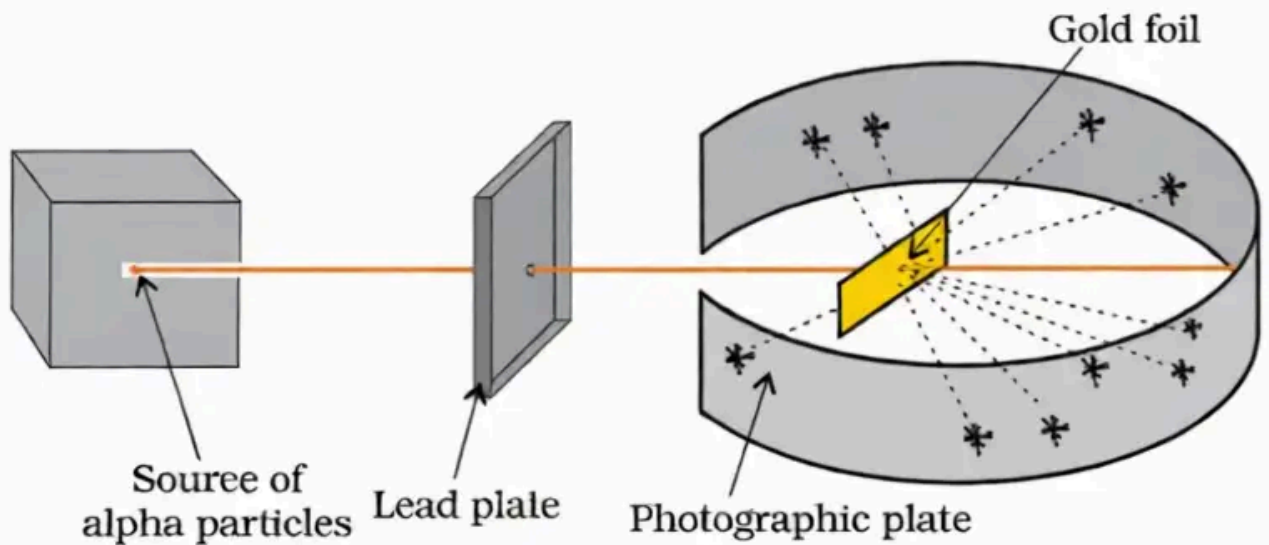
Memory Aids

- "Raisin pudding" = electrons in positive sphere
- Mass uniform, no nucleus concept

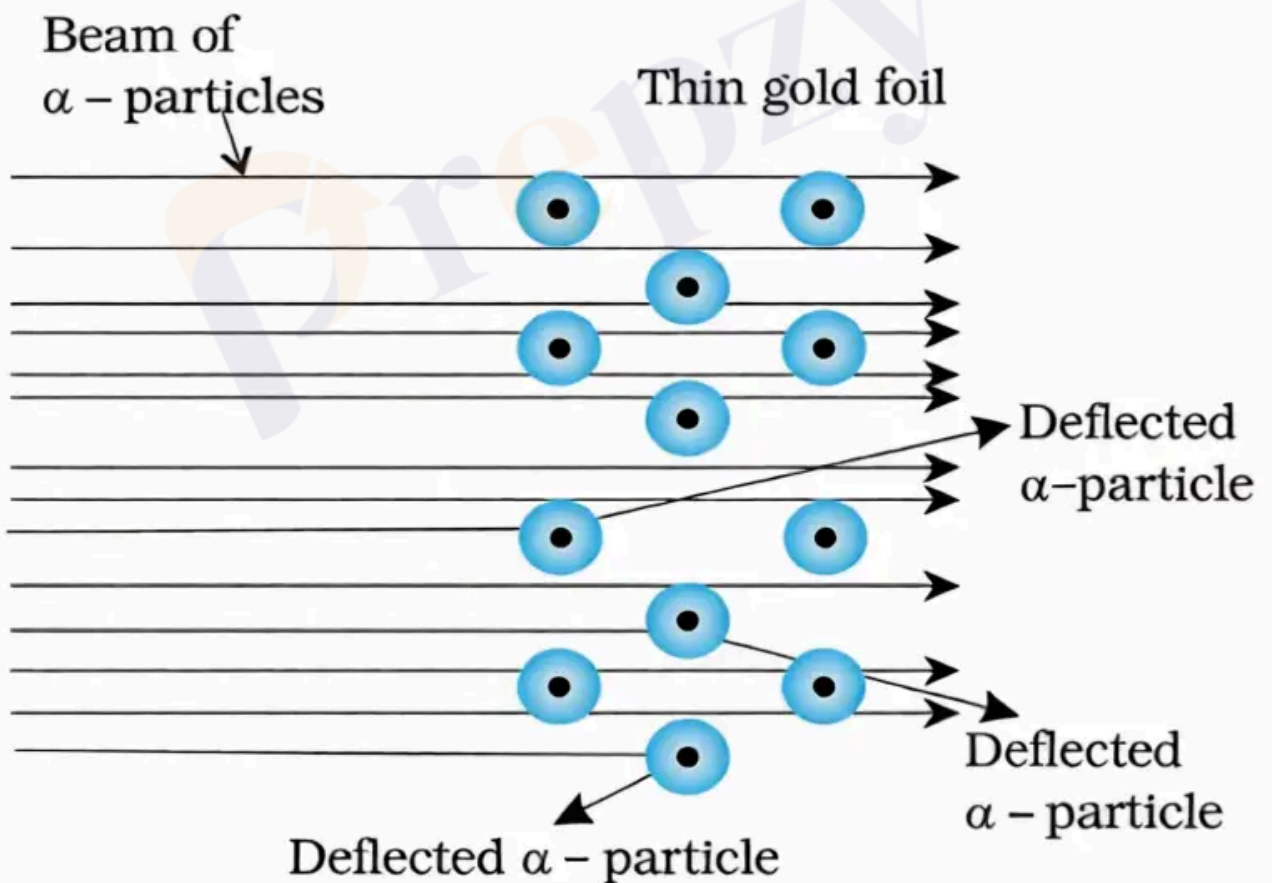
Rutherford Model

Key Observations

- Alpha particles mostly pass through gold foil undeflected → atom mostly empty space.
- Some alpha particles deflected at small and large angles → presence of small, dense, positively charged nucleus.
- Few alpha particles reflected back → nucleus is hard and dense.
- Nucleus radius $\sim 10^{-15}$ m; atom radius $\sim 10^{-10}$ m.
- Electrons revolve around nucleus in circular orbits.
- Electrostatic attraction holds electrons and nucleus together.
- Penetrating power: γ -rays > X-rays > β -rays > α -rays.
- Atomic number (Z) = number of protons = number of electrons in neutral atom.
- Mass number (A) = protons + neutrons.



A. Rutherford's scattering experiment



B. Schematic molecular view of the gold foil

- ▶ **Isotopes** : Different atoms of same element having same atomic number but different mass numbers, *e.g.*, ${}^1_1\text{H}$, ${}^2_1\text{H}$, ${}^3_1\text{H}$; ${}^{35}_{17}\text{Cl}$, ${}^{37}_{17}\text{Cl}$
- ▶ **Isobars** : Atoms of different elements having same mass number but different atomic numbers, *e.g.*, ${}^{40}_{18}\text{Ar}$, ${}^{40}_{19}\text{K}$, ${}^{40}_{20}\text{Ca}$
- **Isotones** : Atoms of different elements containing same number of neutrons, *e.g.*, ${}^{14}_6\text{C}$, ${}^{15}_7\text{N}$, ${}^{16}_8\text{O}$
- **Isodiaphers** : Atoms having same isotopic number (*i.e.*, no. of neutrons – no. of protons = same) *e.g.*, ${}^{235}_{92}\text{U}$, ${}^{231}_{90}\text{Th}$
- **Isosters** : Molecules having same number of atoms and electrons, *e.g.*, CO_2 , N_2O
- **Isoelectronic** : Those species which have same number of electrons, *e.g.*, Na^+ , Mg^{2+} , F^- , Ne

Solved Examples

- Ex 1: What conclusion was drawn from most alpha particles passing through gold foil?
Answer: Atom is mostly empty space.
- Ex 2: Define atomic number.
Answer: Number of protons in nucleus = number of electrons in neutral atom.

Practice Set

- Q1: What does Rutherford's experiment prove about atom?
 - A. Solid sphere
 - B. Mostly empty space
 - C. Uniform positive charge
 - D. No nucleus
- Q2: Atomic number represents:
 - A. Number of neutrons
 - B. Number of protons
 - C. Number of electrons only
 - D. Mass number
- Q3: Which ray has the highest penetrating power?
 - A. α -rays
 - B. β -rays
 - C. X-rays
 - D. γ -rays

Answer Key

- Q1: B - Most alpha particles passed through, showing atom is mostly empty space.
- Q2: B - Atomic number is number of protons in nucleus.

- Q3: D - γ -rays have highest penetrating power.

Memory Aids

- "Gold foil experiment" → nucleus discovery
- Penetrating power order: $\gamma > X > \beta > \alpha$
- Atomic number = protons = electrons (neutral atom)

Drawbacks of Rutherford Model

Limitations

- Electrons revolving in orbits should emit radiation and lose energy → atom collapse (not observed).
- Calculated collapse time $\sim 10^{-8}$ s, but atoms are stable.
- No explanation of electron energy levels or distribution.

Solved Examples

- Ex 1: Why is Rutherford's model unable to explain atomic stability?
Answer: Accelerating electrons should emit radiation and spiral into nucleus.

Practice Set

- Q1: According to Rutherford, electrons in orbits should:
 - A. Remain stable
 - B. Emit radiation and lose energy
 - C. Not move
 - D. Gain energy
- Q2: Rutherford's model fails to explain:
 - A. Nucleus existence
 - B. Electron energy levels
 - C. Atomic number
 - D. Alpha particle scattering

- Q3: Time for electron to spiral into nucleus is approximately:
 - A. 10^8 s
 - B. 10^{-8} s
 - C. 1 s
 - D. Infinite

Answer Key

- Q1: B - Accelerating electrons should emit radiation and lose energy.
- Q2: B - Model does not explain electron energy levels or distribution.
- Q3: B - Calculated collapse time is about 10^{-8} seconds.

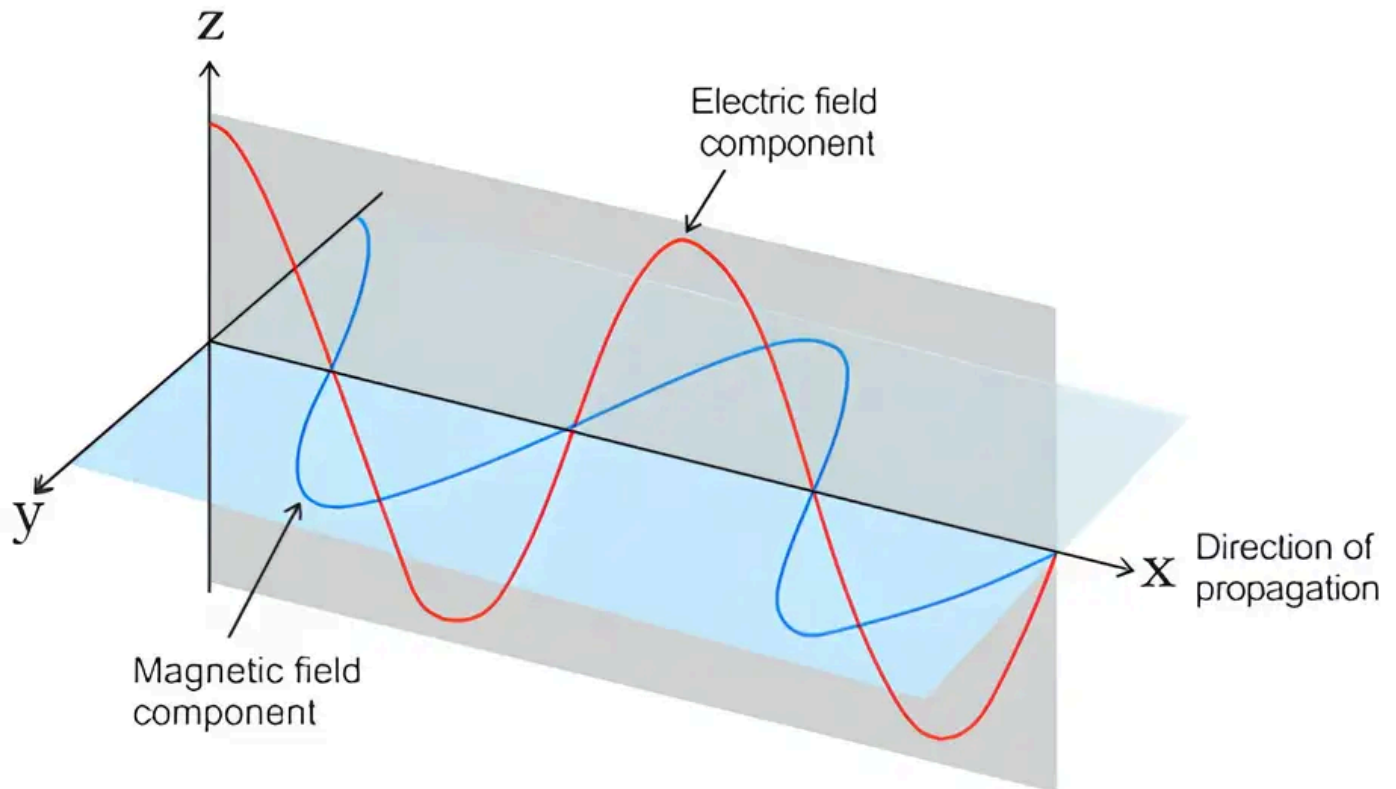
Memory Aids

- "Electron radiation loss" → instability in Rutherford model
- Need for Bohr model to explain stability

Developments Leading to Bohr's Model

Wave Nature of Electromagnetic Radiations

- Energy emitted continuously as electromagnetic waves with perpendicular electric and magnetic fields.
- Electric and magnetic fields vibrate in mutually perpendicular planes.
- Wave travels in one direction with electric and magnetic fields oscillating perpendicular to it.



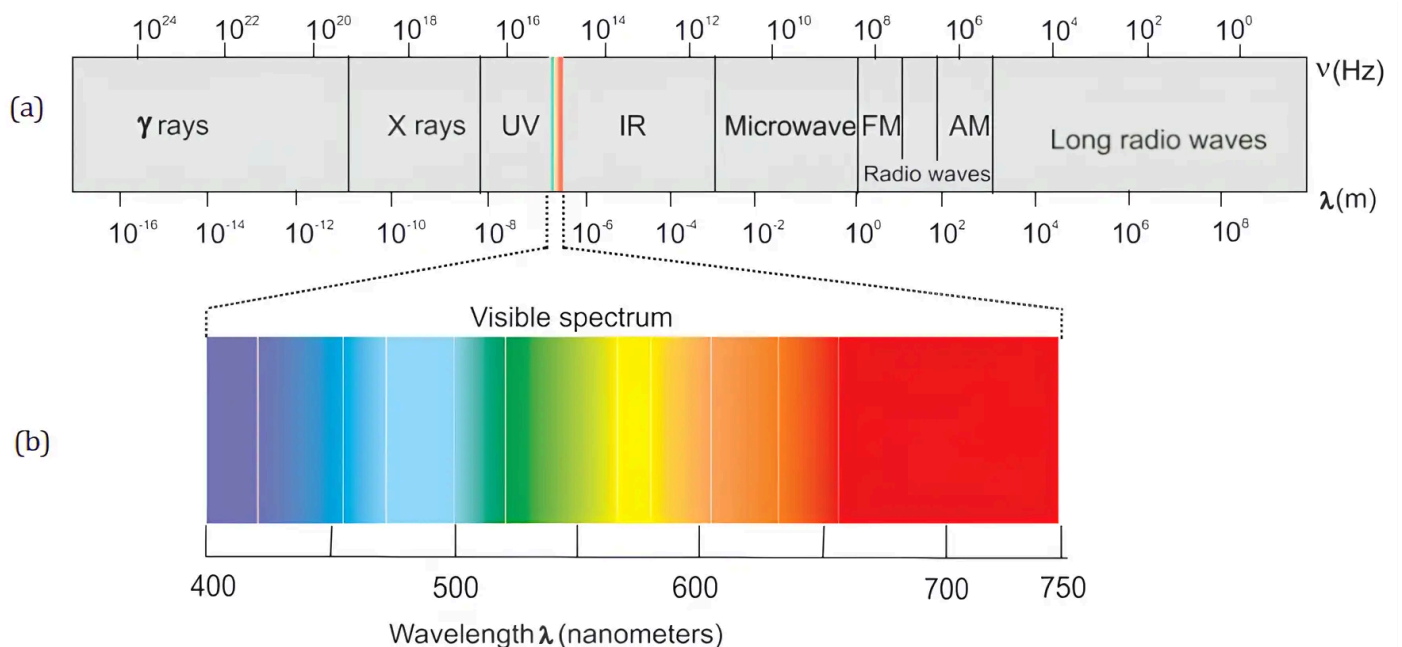
The electric and magnetic field components of an electromagnetic wave.

Characteristic Properties of Electromagnetic Waves

- **Frequency (ν):** Number of waves passing a point per second (Hz or s^{-1}).
- **Wavelength (λ):** Distance between two consecutive crests or troughs (m, Å, pm, nm).
- **Velocity (c):** Speed of wave in vacuum = 3.0×10^8 m/s; $c = \nu\lambda$.
- **Wavenumber ($\bar{\nu}$):** Number of wavelengths per unit length (m^{-1} , commonly cm^{-1}).
- **Amplitude:** Height of crest or depth of trough; determines light intensity.
- **Relations:** $c = \nu\lambda$, $\nu = c/\lambda$, $\bar{\nu} = 1/\lambda$.

Electromagnetic Spectrum

- Arrangement of electromagnetic radiations by increasing wavelength or decreasing frequency.
- Includes gamma rays, X-rays, UV, visible light, infrared, microwaves, and radio waves.
- Visible spectrum ranges from violet (short wavelength) to red (long wavelength).



(a) The spectrum of electromagnetic radiation. (b) Visible spectrum. (The visible region is only A small part of the entire spectrum.)

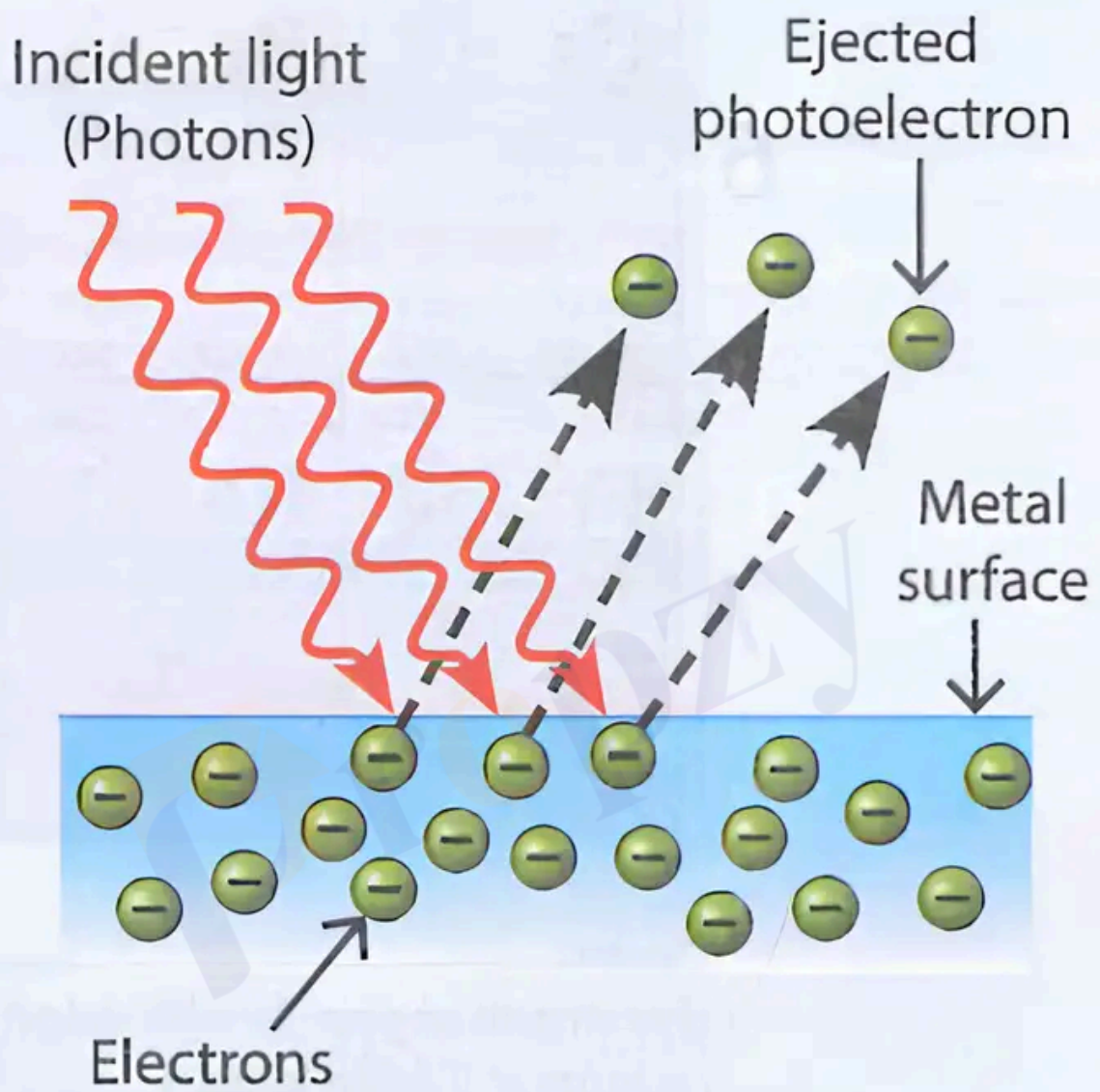
Black Body Radiation and Planck's Quantum Theory

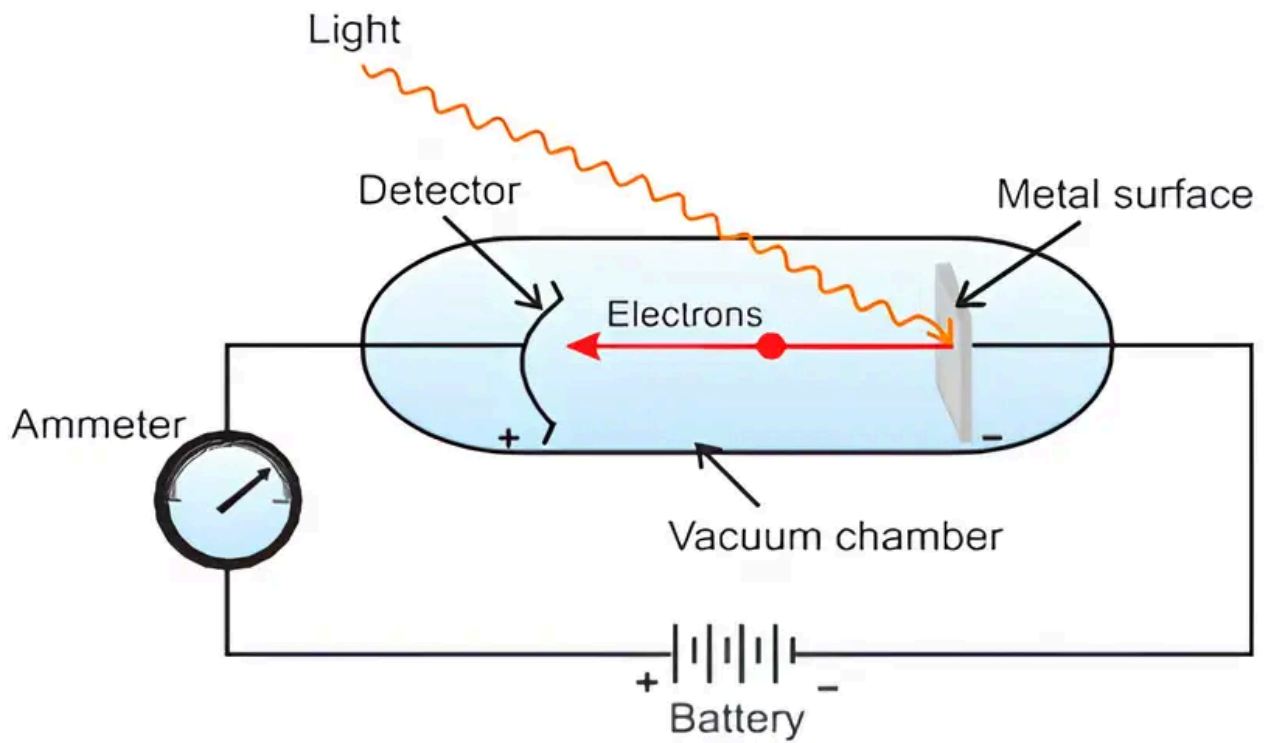
- **Black body:** Ideal body absorbing and emitting all frequencies.
- **Planck's theory:** Energy emitted/absorbed in discrete packets called quanta.
- Energy of quantum: $E = h\nu$, where $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$.
- Total energy: $E = nh\nu$, $n = \text{integer } (0, 1, 2, \dots)$.

Photoelectric Effect

- Emission of electrons from metal surface when light of minimum frequency (threshold frequency ν_0) shines on it.
- Light consists of photons transferring energy to electrons.
- Electrons ejected are called photoelectrons.
- Kinetic energy of ejected electron: $h\nu = h\nu_0 + \frac{1}{2} m_e v^2$.
- Work function ($h\nu_0$): Minimum energy to eject electron.
- Demonstrates particle nature of light.

Photoelectric Effect





Equipment for studying the photoelectric effect.

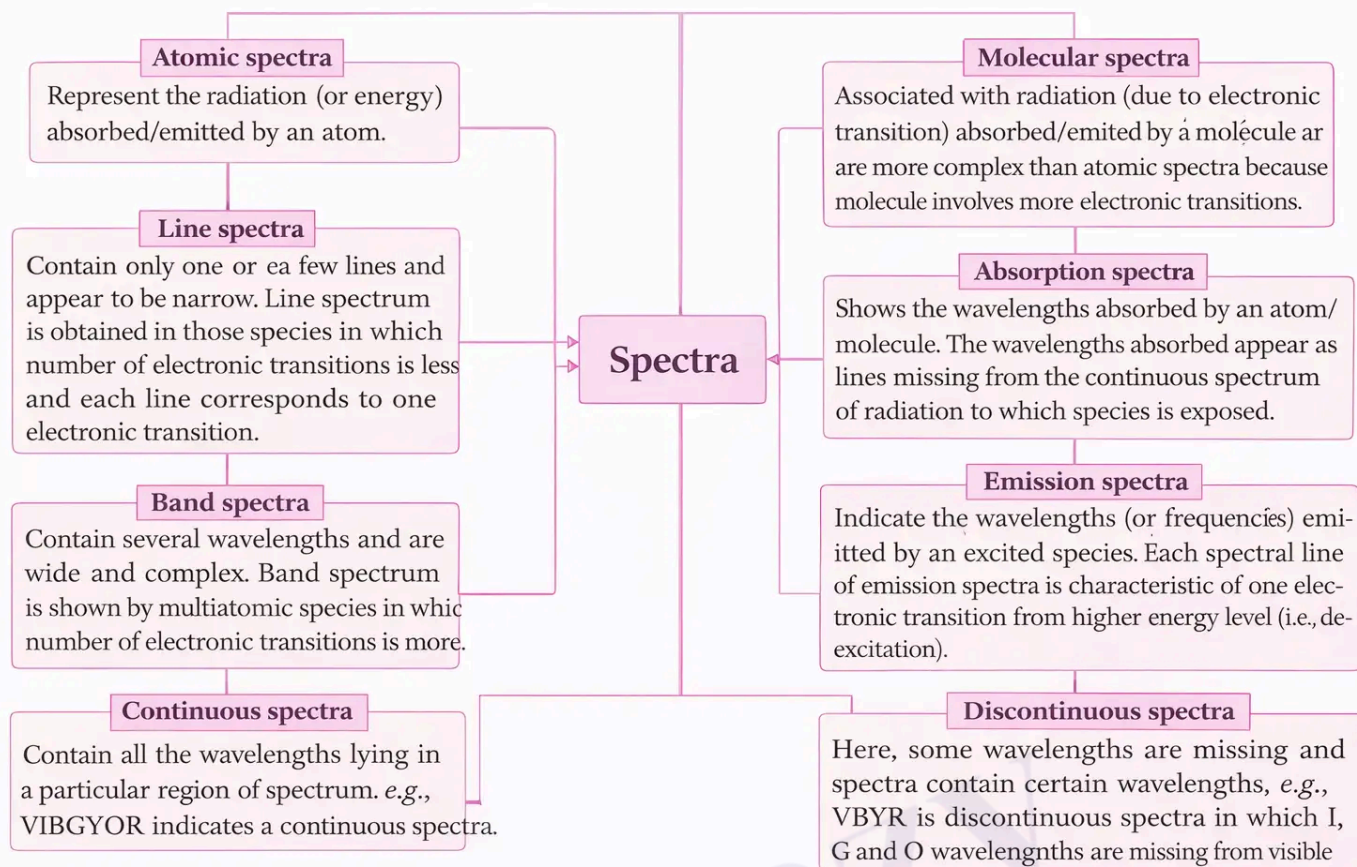
Dual Nature of Electromagnetic Radiation

- Light behaves as wave during propagation and as particles during interaction with matter.

Spectrum Types

- Combination of lights of different wavelengths analyzed through a prism.
- Atomic spectra: line, band, continuous.
- Molecular spectra: absorption, emission, discontinuous.

Types of spectra:



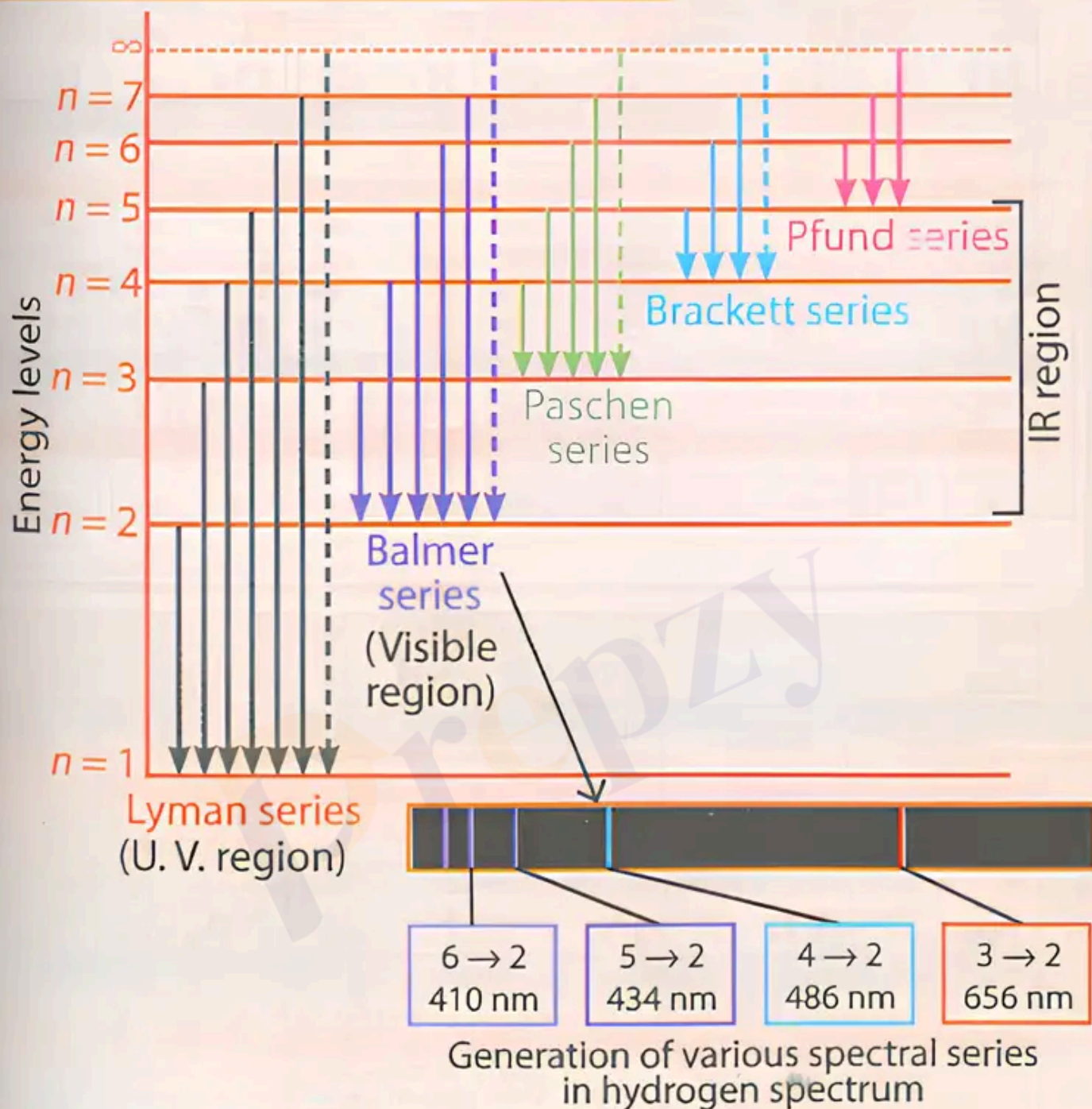
Hydrogen Atomic Spectra

- Spectral lines grouped into series: Lyman (UV), Balmer (visible), Paschen, Brackett, Pfund (infrared).
- Electron transitions between energy levels n_1 and n_2 produce these lines.
- Maximum number of lines from n to 1: $n(n-1)/2$.
- Lines from n_2 to n_1 : $(n_2 - n_1)(n_2 - n_1 + 1)/2$.

The Spectral Lines for Atomic Hydrogen

Series	n_1	n_2	Spectral Region
Lyman	1	2,3....	Ultraviolet
Balmer	2	3,4....	Visible
Paschen	3	4,5....	Infrared
Brackett	4	5,6....	Infrared
Pfund	5	6,7....	Infrared

Line Spectrum of Hydrogen



The emission spectra of atomic hydrogen is divided into a number of spectral series, with wavelength given by Rydberg formula :

$$\frac{1}{\lambda} = RZ^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

These observed spectral lines are due to the electron making transitions between two energy levels in an atom.

Solved Examples

- **Ex 1:** Calculate wavelength if frequency is 6×10^{14} Hz.
Solution: $\lambda = c/\nu = (3 \times 10^8 \text{ m/s}) / (6 \times 10^{14} \text{ Hz}) = 5 \times 10^{-7} \text{ m}$ (visible light).
- **Ex 2:** Find kinetic energy of photoelectron if $h\nu = 5 \text{ eV}$ and work function $h\nu_0 = 2 \text{ eV}$.
Solution: $\text{KE} = h\nu - h\nu_0 = 5 - 2 = 3 \text{ eV}$.
- **Ex 3:** Number of spectral lines when electron falls from $n=4$ to $n=1$.
Solution: $n(n-1)/2 = 4 \times 3 / 2 = 6$ lines.

Practice Set

- **Q1:** What is the speed of electromagnetic waves in vacuum?
A. $3 \times 10^6 \text{ m/s}$
B. $3 \times 10^8 \text{ m/s}$
C. $3 \times 10^{10} \text{ m/s}$
D. $3 \times 10^5 \text{ m/s}$
- **Q2:** Which series of hydrogen spectrum lies in the visible region?
A. Lyman
B. Balmer
C. Paschen
D. Brackett
- **Q3:** Calculate the wavelength of light if frequency is $5 \times 10^{14} \text{ Hz}$.
A. $6 \times 10^{-7} \text{ m}$
B. $5 \times 10^{-7} \text{ m}$
C. $4 \times 10^{-7} \text{ m}$
D. $3 \times 10^{-7} \text{ m}$

Answer Key

- **Q1:** B. $3 \times 10^8 \text{ m/s}$ - Speed of light in vacuum.

- **Q2:** B. Balmer - Visible region series of hydrogen.
- **Q3:** B. 5×10^{-7} m - Using $\lambda = c/\nu$.

Memory Aids

- Speed of light $c = 3 \times 10^8$ m/s.
- Energy quantum $E = h\nu$; Planck's constant $h = 6.626 \times 10^{-34}$ J·s.
- Visible spectrum: VIBGYOR (Violet shortest λ , Red longest λ).
- Photoelectric effect metals: K, Rb, Cs.
- Max spectral lines from n : $n(n-1)/2$.

Bohr's Model for Hydrogen Atom

Key Concepts

- Electrons move in fixed circular orbits (energy shells) around the nucleus with fixed radius and energy.
- Energy of an electron in a given orbit remains constant over time.
- Electron absorbs energy to jump to a higher orbit and releases energy when falling back to a lower orbit.
- Energy levels follow: $E_1 < E_2 < E_3 \dots$
- Energy difference between levels: $\Delta E = E_2 - E_1 = h\nu$, with ΔE decreasing for higher orbits.
- Angular momentum quantization: $mvr = nh/2\pi$, where $n = 1, 2, 3\dots$
- Energy of n th orbit: $E_n = -R_H (Z^2/n^2) = -2.18 \times 10^{-18} \times Z^2/n^2$ J.
- Rydberg constant $R_H = 2.18 \times 10^{-18}$ J.
- Radius of n th orbit for hydrogen: $r_n = 0.529 \times n^2$ Å.
- For one-electron species with atomic number Z : $r_n = 0.529 \times n^2/Z$ Å.

$$E_n = \frac{-2\pi^2 m e^4 Z^2}{n^2 h^2}; E_n = \frac{-1312 Z^2}{n^2} \text{ kJ mol}^{-1}; E_n = -\frac{Z^2}{n^2} \times 13.6 \text{ eV}$$

Here, Z = nuclear charge for H-atom and single electron species like He^+ , Li^{2+} , Be^{3+} , etc.

- **Ionisation energy of hydrogen like particles :**

$$I.E. = E_{\infty} - E_n = 0 - \left(-\frac{Z^2}{n^2} \times 2.18 \times 10^{-18} \text{ J} \right) = \frac{Z^2}{n^2} \times 2.18 \times 10^{-18} \text{ J}$$

- **Velocity of electron in n^{th} shell of hydrogen like particles :**

$$v_n = \frac{2\pi e^2 Z}{nh}, v_n = \frac{Z}{n} \times 2.188 \times 10^8 \text{ cm/sec}$$

- $r_n = \frac{n^2 h^2}{4\pi^2 m e^2 Z}$, r_n = radius of n^{th} orbit, m = mass of electron, e = charge on electron

Solved Examples

- **Ex 1:** Calculate the radius of the 3rd orbit of hydrogen.

- Use $r_n = 0.529 \times n^2 \text{ \AA}$
- $r_3 = 0.529 \times 3^2 = 0.529 \times 9 = 4.761 \text{ \AA}$

- **Ex 2:** Find the energy of the 2nd orbit of He^+ ($Z=2$).

- $E_n = -2.18 \times 10^{-18} \times Z^2/n^2 \text{ J}$
- $E_2 = -2.18 \times 10^{-18} \times 4/4 = -2.18 \times 10^{-18} \text{ J}$

Practice Set

- Q1: What is the angular momentum of an electron in the 1st orbit?
 - A. 0
 - B. $h/2\pi$
 - C. $2h/2\pi$
 - D. nh
- Q2: The radius of the 4th orbit of hydrogen atom is:
 - A. $0.529 \text{ \AA}$
 - B. $8.464 \text{ \AA}$
 - C. $0.529 \times 16 \text{ \AA}$
 - D. $2.116 \text{ \AA}$
- Q3: Energy difference between 2nd and 1st orbit is:
 - A. $h\nu$
 - B. $E_2 - E_1$

- C. Both A and B
- D. None

Answer Key

- Q1: B - Angular momentum is quantized as $mvr = nh/2\pi$, for $n=1$ it is $h/2\pi$.
- Q2: C - $r_4 = 0.529 \times 4^2 = 0.529 \times 16 \text{ \AA}$.
- Q3: C - Energy difference $\Delta E = E_2 - E_1 = h\nu$.

Memory Aids

- Angular momentum quantization: "n times h over 2 pi".
- Radius scales as n squared: "Radius $\propto n^2$ ".
- Energy levels are negative and inversely proportional to n squared.

Rydberg Formula

Key Concepts

- Used for hydrogen and hydrogen-like atoms to calculate spectral lines.
- Formula: $1/\lambda = \nu = R [1/n_1^2 - 1/n_2^2] Z^2$
- $R = 109,677 \text{ cm}^{-1}$ (Rydberg constant in wave numbers).
- Z is the atomic number.

Solved Examples

- **Ex 1:** Calculate wavelength for transition $n_2=3$ to $n_1=2$ in hydrogen ($Z=1$).
 - $1/\lambda = R [1/2^2 - 1/3^2] = 109,677 [1/4 - 1/9]$
 - $1/\lambda = 109,677 \times (5/36) = 15,243 \text{ cm}^{-1}$
 - $\lambda = 1/15,243 \text{ cm} = 6.56 \times 10^{-5} \text{ cm} = 656 \text{ nm}$

Practice Set

- Q1: What is the value of Rydberg constant R?
 - A. $109,677 \text{ cm}^{-1}$
 - B. $2.18 \times 10^{-18} \text{ J}$
 - C. $0.529 \text{ \AA}$
 - D. $6.626 \times 10^{-34} \text{ Js}$
- Q2: For hydrogen, the spectral line corresponds to transition from $n=4$ to $n=2$. Which formula is used?
 - A. $E_n = -R_H Z^2/n^2$
 - B. $1/\lambda = R [1/n_1^2 - 1/n_2^2] Z^2$
 - C. $mvr = nh/2\pi$
 - D. $r_n = 0.529 n^2/Z$
- Q3: What does Z represent in the Rydberg formula?
 - A. Principal quantum number
 - B. Atomic number
 - C. Energy level
 - D. Wavelength

Answer Key

- Q1: A - Rydberg constant $R = 109,677 \text{ cm}^{-1}$.
- Q2: B - Rydberg formula calculates spectral lines for transitions.
- Q3: B - Z is the atomic number.

Memory Aids

- Rydberg formula relates inverse wavelength to energy level differences.
- Remember $R = 109,677 \text{ cm}^{-1}$ for hydrogen.
- Z squared factor accounts for nuclear charge effect.

Limitations of Bohr's Theory

Key Points

- Fails to explain spectra of multi-electron atoms.
- Cannot explain electron distribution in atoms.
- Does not account for splitting of spectral lines under magnetic (Zeeman effect) and electric fields (Stark effect).
- Ignores wave-particle duality and Heisenberg's uncertainty principle.

Solved Examples

- **Ex 1:** Why can't Bohr's model explain Zeeman effect?
 - Because it ignores electron spin and magnetic interactions.

Practice Set

- Q1: Which phenomenon is NOT explained by Bohr's model?
 - A. Hydrogen spectrum
 - B. Zeeman effect
 - C. Energy quantization
 - D. Radius of orbits
- Q2: Bohr's model ignores which principle?
 - A. Conservation of energy
 - B. Uncertainty principle
 - C. Pauli exclusion principle
 - D. Aufbau principle
- Q3: Bohr's model is limited to atoms with how many electrons?
 - A. One
 - B. Two
 - C. Many
 - D. Zero

Answer Key

- Q1: B - Zeeman effect involves magnetic splitting not explained by Bohr's model.
- Q2: B - It ignores the uncertainty principle.
- Q3: A - Applicable only to one-electron atoms/ions.

Memory Aids

- Bohr's model = "One electron only" model.
- Does not explain magnetic/electric field effects on spectra.
- Ignores wave nature and uncertainty principle.

Towards Quantum Mechanical Model of the Atom

Wave-particle duality and de Broglie's equation

- Electrons exhibit dual nature: particle and wave.
- de Broglie equation: $\lambda = \frac{h}{p} = \frac{h}{mv}$, where λ is wavelength, h Planck's constant, p momentum.
- Relation: $E = mc^2 = h\nu = \frac{hc}{\lambda}$.
- Every moving object has wave character; de Broglie waves are associated with electrons only.
- Wavelength decreases with increasing mass; negligible for heavy particles.
- Applicable mainly to subatomic particles like electrons.
- Alternate form: $\lambda = \frac{h}{\sqrt{2mE_k}} = \frac{h}{\sqrt{2mqV}}$, where E_k kinetic energy, q charge, m mass, V accelerating potential.

Solved Examples

- **Ex 1:** Calculate wavelength of electron accelerated through 100 V.
 - Use $\lambda = \frac{h}{\sqrt{2mqV}}$.
 - Substitute values: $h = 6.626 \times 10^{-34} Js$, $m = 9.11 \times 10^{-31} kg$, $q = 1.6 \times 10^{-19} C$, $V = 100V$.
 - Calculate λ in meters.

Practice Set

- Q1: What is the de Broglie wavelength of a particle with momentum $2 \times 10^{-24} \text{ kg m/s}$?
 - A) $3.3 \times 10^{-10} \text{ m}$
 - B) $6.6 \times 10^{-34} \text{ m}$
 - C) $1.0 \times 10^{-10} \text{ m}$
 - D) $1.2 \times 10^{-9} \text{ m}$
- Q2: Which particle has the largest de Broglie wavelength?
 - A) Electron
 - B) Proton
 - C) Neutron
 - D) Alpha particle
- Q3: The wavelength of a particle is inversely proportional to:
 - A) Mass
 - B) Velocity
 - C) Momentum
 - D) All of the above

Answer Key

- Q1: A - $\lambda = \frac{h}{p} = \frac{6.626 \times 10^{-34}}{2 \times 10^{-24}} = 3.3 \times 10^{-10} \text{ m}$.
- Q2: A - Electron has smallest mass, hence largest wavelength.
- Q3: C - Wavelength inversely proportional to momentum $p = mv$.

Memory Aids

- de Broglie wavelength $\lambda = \frac{h}{mv}$ remember as "wavelength inversely proportional to momentum".
- Wave nature significant only for very small mass particles.
- Heavier particles have negligible wave properties.

Heisenberg's Uncertainty Principle

Principle and significance

- It is impossible to simultaneously know exact position (Δx) and momentum (Δp) of an electron.
- Mathematically: $\Delta x \times \Delta p \geq \frac{h}{4\pi}$ or $\Delta x \times m\Delta v \geq \frac{h}{4\pi}$.
- Uncertainty principle rules out definite electron paths or trajectories.
- Significant only for microscopic particles; negligible for macroscopic objects.

Solved Examples

- **Ex 1:** Calculate minimum uncertainty in position if uncertainty in momentum is $1 \times 10^{-24} \text{ kg m/s}$.
 - Use $\Delta x \geq \frac{h}{4\pi\Delta p}$.
 - Substitute $h = 6.626 \times 10^{-34} \text{ Js}$, $\Delta p = 1 \times 10^{-24} \text{ kg m/s}$.
 - Calculate Δx in meters.

Practice Set

- Q1: Heisenberg's uncertainty principle states that:
 - A) Position and momentum can be known exactly.
 - B) Position and momentum cannot be known simultaneously with exact precision.
 - C) Only position can be known exactly.
 - D) Only momentum can be known exactly.
- Q2: If uncertainty in position decreases, uncertainty in momentum:
 - A) Decreases
 - B) Increases
 - C) Remains same
 - D) Becomes zero
- Q3: The uncertainty principle is most significant for:
 - A) Macroscopic objects
 - B) Microscopic particles
 - C) Both equally
 - D) None

Answer Key

- Q1: B - Exact simultaneous knowledge of position and momentum is impossible.
- Q2: B - Decreasing uncertainty in position increases uncertainty in momentum.
- Q3: B - Principle applies mainly to microscopic particles like electrons.

Memory Aids

- Remember: $\Delta x \times \Delta p \geq \frac{h}{4\pi}$ means position and momentum cannot both be exact.
- Uncertainty principle forbids exact electron paths.
- Significant only at atomic/subatomic scale.

Quantum Mechanical Model of Atom

Key Concepts

- Schrodinger wave equation: $\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2m}{h^2}(E - V)\psi = 0$
- ψ is the wave function; ψ^2 gives the probability density of finding an electron in a small region.
- Atomic orbital: 3D space around nucleus with maximum electron probability.
- Energy of electrons is quantized due to wave nature and solutions of Schrodinger equation.
- Exact position and velocity of electron cannot be known simultaneously (Heisenberg uncertainty principle).
- Probability density ψ^2 is always positive and proportional to electron presence.

Solved Examples

- **Ex 1:** Write the Schrodinger wave equation for an electron in an atom.

Answer: $\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2m}{h^2}(E - V)\psi = 0$

Practice Set

- Q1: What does ψ^2 represent in the quantum mechanical model?
 - A. Electron velocity
 - B. Electron position
 - C. Probability density of electron
 - D. Energy of electron
- Q2: Which principle states that exact position and velocity of electron cannot be known simultaneously?
 - A. Pauli exclusion principle
 - B. Hund's rule
 - C. Heisenberg uncertainty principle
 - D. Aufbau principle
- Q3: An atomic orbital is defined as:
 - A. A point where electron is found
 - B. A 2D space around nucleus
 - C. A 3D space with maximum electron probability
 - D. The path of electron around nucleus

Answer Key

- Q1: C - ψ^2 gives the probability density of finding an electron.
- Q2: C - Heisenberg uncertainty principle limits simultaneous knowledge of position and velocity.
- Q3: C - Atomic orbital is a 3D region with maximum probability of electron presence.

Memory Aids

- Remember: ψ is wave function, ψ^2 is probability density.
- Heisenberg uncertainty = no exact position & velocity simultaneously.
- Atomic orbital = 3D region of max electron probability.

Quantum Numbers

Key Concepts

- Four quantum numbers describe an electron completely: Principal (n), Azimuthal (l), Magnetic (m_l), Spin (m_s).
- Principal quantum number (n): Energy level/shell (1, 2, 3, ...).
- Azimuthal quantum number (l): Subshell shape (0 to $n-1$), $s=0$, $p=1$, $d=2$, $f=3$.
- Magnetic quantum number (m_l): Orbital orientation ($-l$ to $+l$).
- Spin quantum number (m_s): Electron spin ($+1/2$ or $-1/2$).
- Number of subshells in n^{th} shell = n ; orbitals = n^2 ; max electrons = $2n^2$.
- Orbitals in subshell = $2l + 1$; max electrons = $2(2l + 1)$.

Quantum number	Values	Information conveyed
Principal quantum number (n)	Any positive integer value i.e., 1, 2, 3, 4, etc. $n = 1$, K-shell $n = 2$, L-shell $n = 3$, M-shell $n = 4$, N-shell	<ul style="list-style-type: none">• Main shell/orbit of the electron.• Approximate distance from the nucleus.• Energy of shell.• Maximum number of electrons present in the shell ($2n^2$).• Explains the main lines of spectrum.• Angular momentum can be calculated as, $mvr = \frac{nh}{2\pi}$
Azimuthal quantum number (l)	For a particular value of n , $l = 0$ to $n - 1$. $l = 0$, s-subshell $l = 1$, p-subshell $l = 2$, d-subshell $l = 3$, f-subshell	<ul style="list-style-type: none">• No. of subshells present in the main shell.• Relative energies of the subshells.• Shapes of orbitals.• Explains the fine structure of the line spectrum.• Orbital angular momentum of an electron can be calculated as $\mu_l = \sqrt{l(l+1)} \frac{h}{2\pi}$.
Magnetic quantum number (m_l)	For a particular value of l , $m_l = -l$ to $+l$ including zero.	<ul style="list-style-type: none">• Spatial orientation of the orbital with respect to standard set of co-ordinate axis.• No. of orbitals present in any subshell.• Explains the splitting of lines in a magnetic field.
Electron spin quantum number (m_s)	$+\frac{1}{2}$, $-\frac{1}{2}$	<ul style="list-style-type: none">• Direction of electron spin (clockwise or anticlockwise).• Explains the magnetic properties of substances.• Spin angular momentum = $\sqrt{s(s+1)} \frac{h}{2\pi}$

Solved Examples

- **Ex 1:** Find the number of orbitals in the 3rd shell.

Solution: Number of orbitals = $n^2 = 3^2 = 9$.

- **Ex 2:** What is the range of m_l for $l = 2$?

Solution: $m_l = -2, -1, 0, +1, +2$.

Practice Set

- Q1: What is the maximum number of electrons in the 2nd shell?
A. 2
B. 8
C. 18
D. 32
- Q2: The azimuthal quantum number for p-orbital is:
A. 0
B. 1
C. 2
D. 3
- Q3: How many orbitals are there in the d-subshell?
A. 3
B. 5
C. 7
D. 9

Answer Key

- Q1: B - Max electrons = $2n^2 = 2(2)^2 = 8$.
- Q2: B - p-orbital corresponds to $l = 1$.
- Q3: B - Number of orbitals in d-subshell = $2l + 1 = 5$.

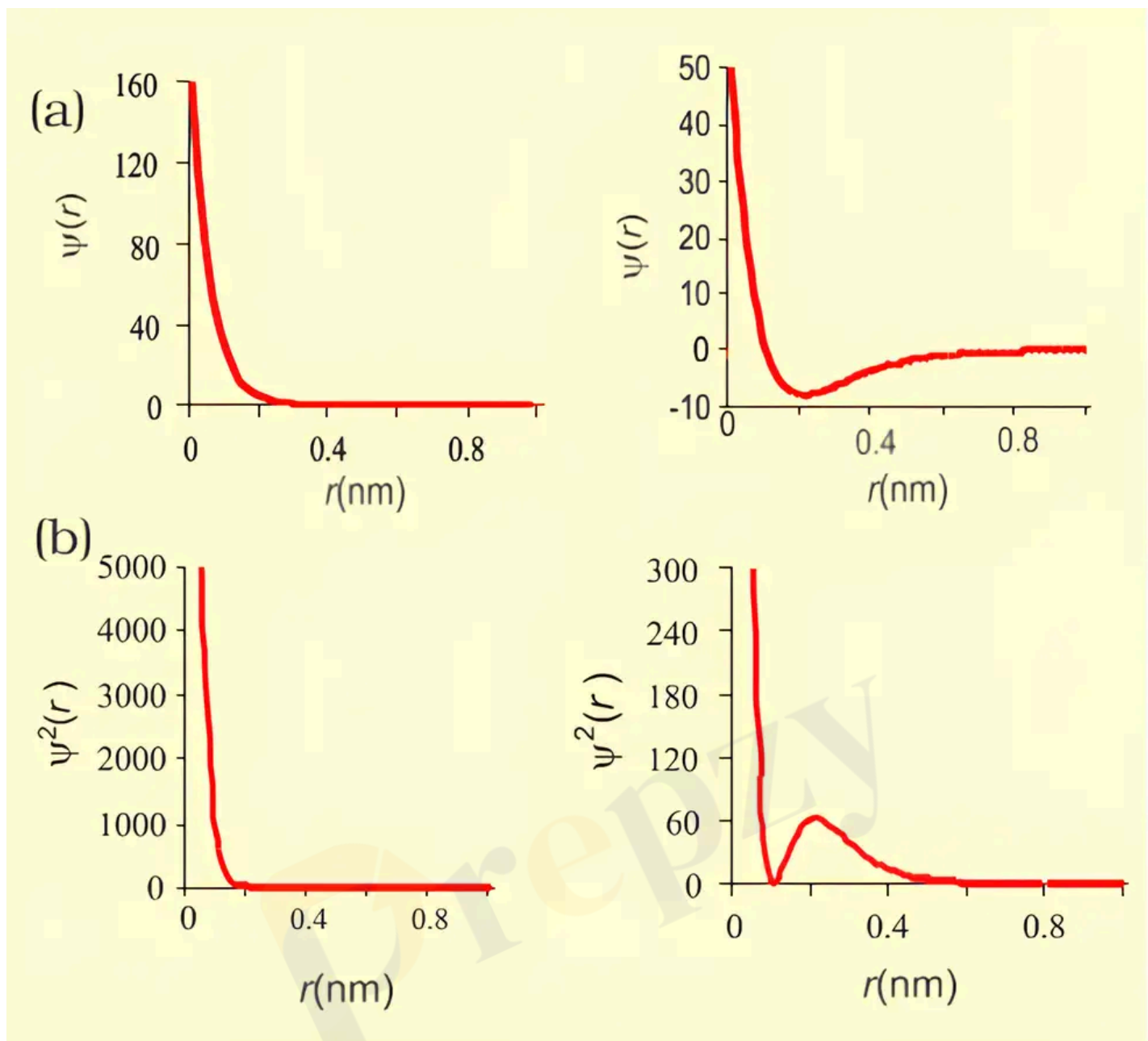
Memory Aids

- n = shell number; l = subshell shape ($s=0$, $p=1$, $d=2$, $f=3$).
- m_l ranges from $-l$ to $+l$.
- $m_s = +1/2$ or $-1/2$ (spin up/down).

Shapes of Atomic Orbitals

Key Concepts

- Orbital shape given by total probability density ψ^2 (radial + angular parts).
- 1s orbital: spherical shape, max probability at nucleus, decreases outward.
- 2s orbital: has radial node where probability density is zero.
- Nodes: regions where probability density is zero; radial nodes = $n - l - 1$, angular nodes = l , total nodes = $n - 1$.
- p-orbitals: dumbbell shape with two lobes along x, y, z axes; one nodal plane through nucleus.
- d-orbitals: clover-leaf shape; five types; d_{z^2} has doughnut shape; two nodal planes each.



The plots of (a) the orbital wave function $\psi(r)$; (b) the variation of probability density $\psi^2(r)$ as a function of distance r of the electron from the nucleus for 1s and 2s orbitals.

Solved Examples

- **Ex 1:** Calculate radial nodes for 3p orbital.
Solution: Radial nodes = $n - l - 1 = 3 - 1 - 1 = 1$.

- **Ex 2:** Describe shape of p-orbital.

Solution: Dumbbell shape with two lobes opposite nucleus along an axis.

Practice Set

- Q1: Number of total nodes in 4d orbital?
A. 2
B. 3
C. 4
D. 5
- Q2: Shape of s-orbital is:
A. Dumbbell
B. Spherical
C. Clover-leaf
D. Doughnut
- Q3: How many nodal planes does a d-orbital have?
A. 1
B. 2
C. 3
D. 4

Answer Key

- Q1: C - Total nodes = $n - 1 = 4 - 1 = 3$.
- Q2: B - s-orbitals are spherical.
- Q3: B - Each d-orbital has two nodal planes.

Memory Aids

- s = spherical, p = dumbbell, d = clover-leaf.
- Nodes total = $n - 1$; radial = $n - l - 1$; angular = l .
- d_{z^2} orbital has doughnut shape.

Energy of Orbitals

Key Concepts

- In hydrogen atom, energy depends only on principal quantum number (n).
- Energy order: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < \dots$
- Orbitals with same energy are degenerate.
- $1s$ orbital is ground state; higher orbitals are excited states.
- In multi-electron atoms, energy depends on n and l .
- Lower $(n + l)$ means lower energy; if equal, lower n means lower energy.
- Energy decreases with increasing atomic number (Z_{eff}) in same subshell.

Solved Examples

- **Ex 1:** Arrange orbitals in increasing energy for hydrogen: $3p, 2s, 1s$.
Answer: $1s < 2s < 3p$.
- **Ex 2:** Which orbital has lower energy: $4s$ or $3d$?
Answer: $4s$ has lower energy than $3d$.

Practice Set

- Q1: Which orbital is degenerate with $2p$ in hydrogen atom?
A. $2s$
B. $3s$
C. $1s$
D. $3p$
- Q2: In multi-electron atoms, energy depends on:
A. Only n
B. Only l
C. Both n and l
D. Neither n nor l
- Q3: Which has the lowest energy?
A. $3d$
B. $4s$
C. $4p$
D. $3p$

Answer Key

- Q1: A - 2s and 2p are degenerate in hydrogen.
- Q2: C - Energy depends on both n and l in multi-electron atoms.
- Q3: B - 4s has lower energy than 3d and 4p.

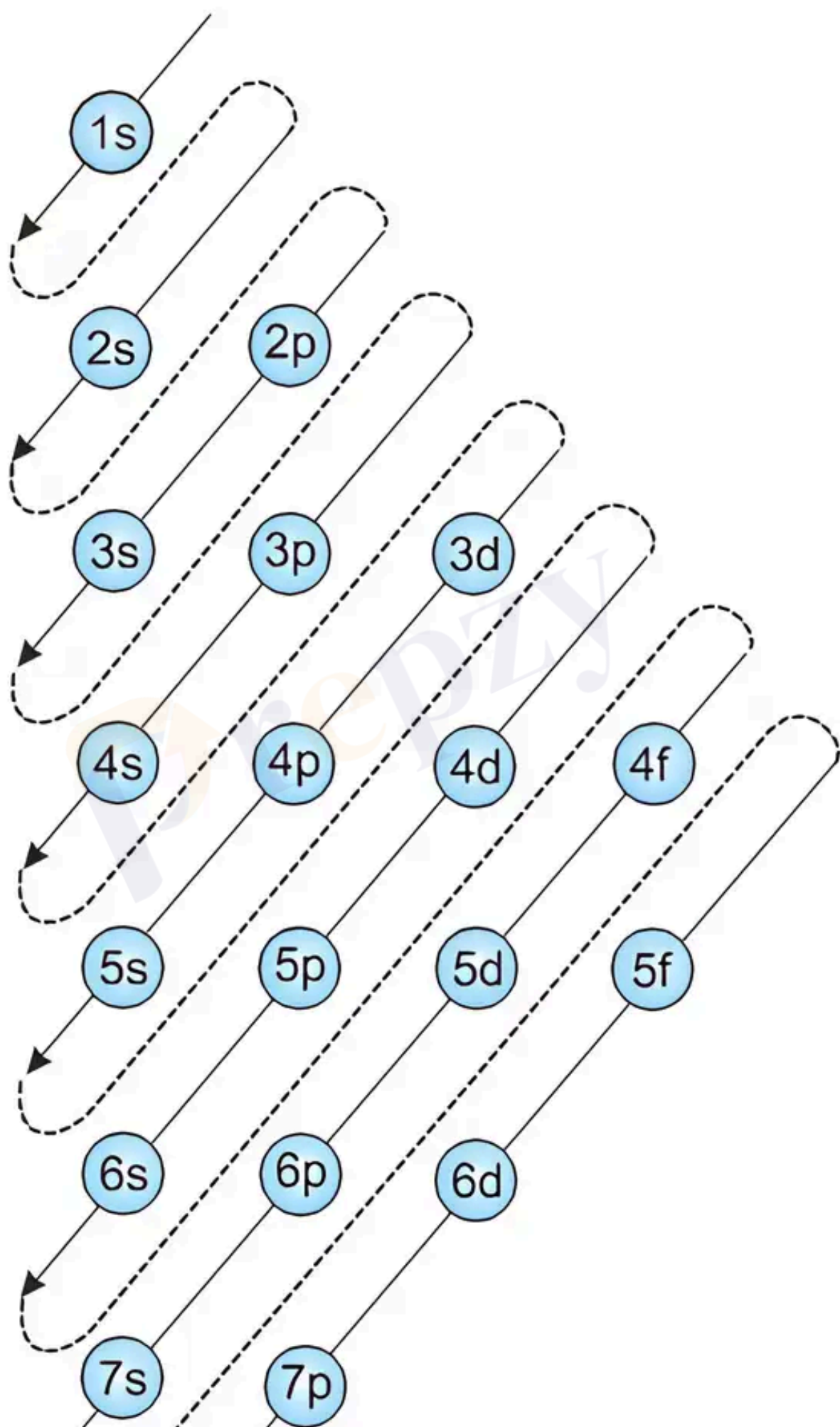
Memory Aids

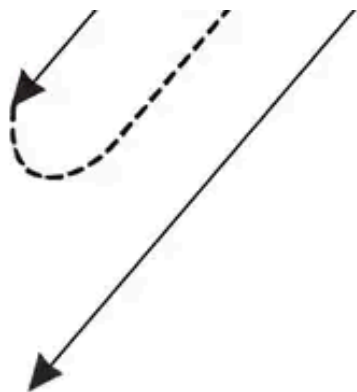
- Energy order follows $(n + l)$ rule.
- Degenerate orbitals have same energy in hydrogen.
- 4s fills before 3d due to lower energy.

Filling of Orbitals in Atom

Key Concepts

- Aufbau principle: Orbitals fill in order of increasing energy (1s, 2s, 2p, 3s, 3p, 4s, 3d, ...).
- Pauli exclusion principle: No two electrons have same four quantum numbers; max two electrons per orbital with opposite spins.
- Hund's rule: Electrons fill orbitals singly first, then pair up.





Order of filling of orbitals

Solved Examples

- **Ex 1:** Write the order of filling for first 10 electrons.

Answer: $1s < 2s < 2p < 3s < 3p$.

- **Ex 2:** How many electrons can occupy one orbital?

Answer: Maximum 2 with opposite spins.

Practice Set

- Q1: According to Pauli principle, maximum electrons in one orbital are:
A. 1
B. 2
C. 3
D. 4
- Q2: Hund's rule states that:
A. Electrons pair first
B. Orbitals fill randomly
C. Orbitals fill singly before pairing
D. Electrons have same spin
- Q3: Aufbau principle means:
A. Electrons fill highest energy orbitals first
B. Electrons fill orbitals in increasing energy order
C. Electrons fill orbitals randomly
D. Electrons fill orbitals based on spin

Answer Key

- Q1: B - Max 2 electrons per orbital with opposite spins.
- Q2: C - Orbitals fill singly before pairing (Hund's rule).
- Q3: B - Orbitals fill in increasing energy order (Aufbau).

Memory Aids

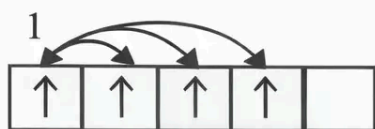
- Remember: Pauli = 2 electrons max per orbital, opposite spins.
- Hund = fill singly first.
- Aufbau = fill lowest energy orbitals first.

Electronic Configuration of Atoms

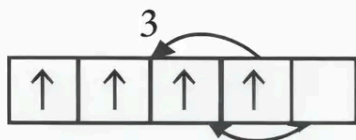
Key Concepts

- Electronic configuration: distribution of electrons in orbitals following filling rules.
- Half-filled and fully-filled orbitals are more stable due to symmetry and exchange energy.
- Symmetry: symmetrical electron distribution increases stability.
- Exchange energy: energy released when electrons exchange positions in same subshell; max in half/full-filled subshells.

In case of d^4 configuration,



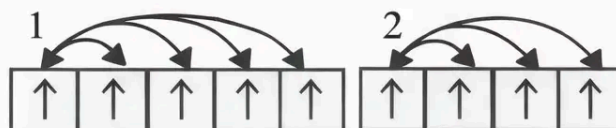
Three exchanges by 1st electron



One exchange by 3rd electron

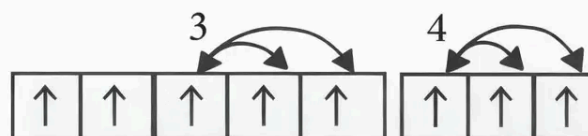
Total number of exchanges =
 $3 + 2 + 1 = 6$

In case of d^5 configuration,



Four exchanges
by 1st electron

Three exchanges
by 2nd electron



Two exchanges by
3rd electron

One exchange
4th electron

Total number of exchanges =
 $4 + 3 + 2 + 1 = 10$

Solved Examples

- **Ex 1:** Why is d^5 more stable than d^4 ?
Answer: More exchange energy due to more electron exchanges.
- **Ex 2:** What causes stability in half-filled orbitals?
Answer: Symmetry and maximum exchange energy.

Practice Set

- Q1: Which configuration is more stable?
A. d^4
B. d^5
C. d^3
D. d^2
- Q2: Exchange energy is maximum in:
A. Half-filled subshell
B. Empty subshell

- C. Partially filled subshell
- D. Fully filled subshell
- Q3: Stability due to symmetry is because:
 - A. Unequal electron distribution
 - B. Symmetrical electron distribution
 - C. Random electron spins
 - D. No electron pairing

Answer Key

- Q1: B - d^5 has maximum exchange energy and stability.
- Q2: A - Exchange energy max in half-filled subshells.
- Q3: B - Symmetry leads to greater stability.

Memory Aids

- Half-filled and full-filled orbitals = max stability.
- Exchange energy increases with electron exchanges.
- Symmetry = stability.