

Section 2.4.1

Explanation of Line Spectrum of Hydrogen

Decoding the microscopic leaps that create atomic light.

Electron Transition Between Energy Levels

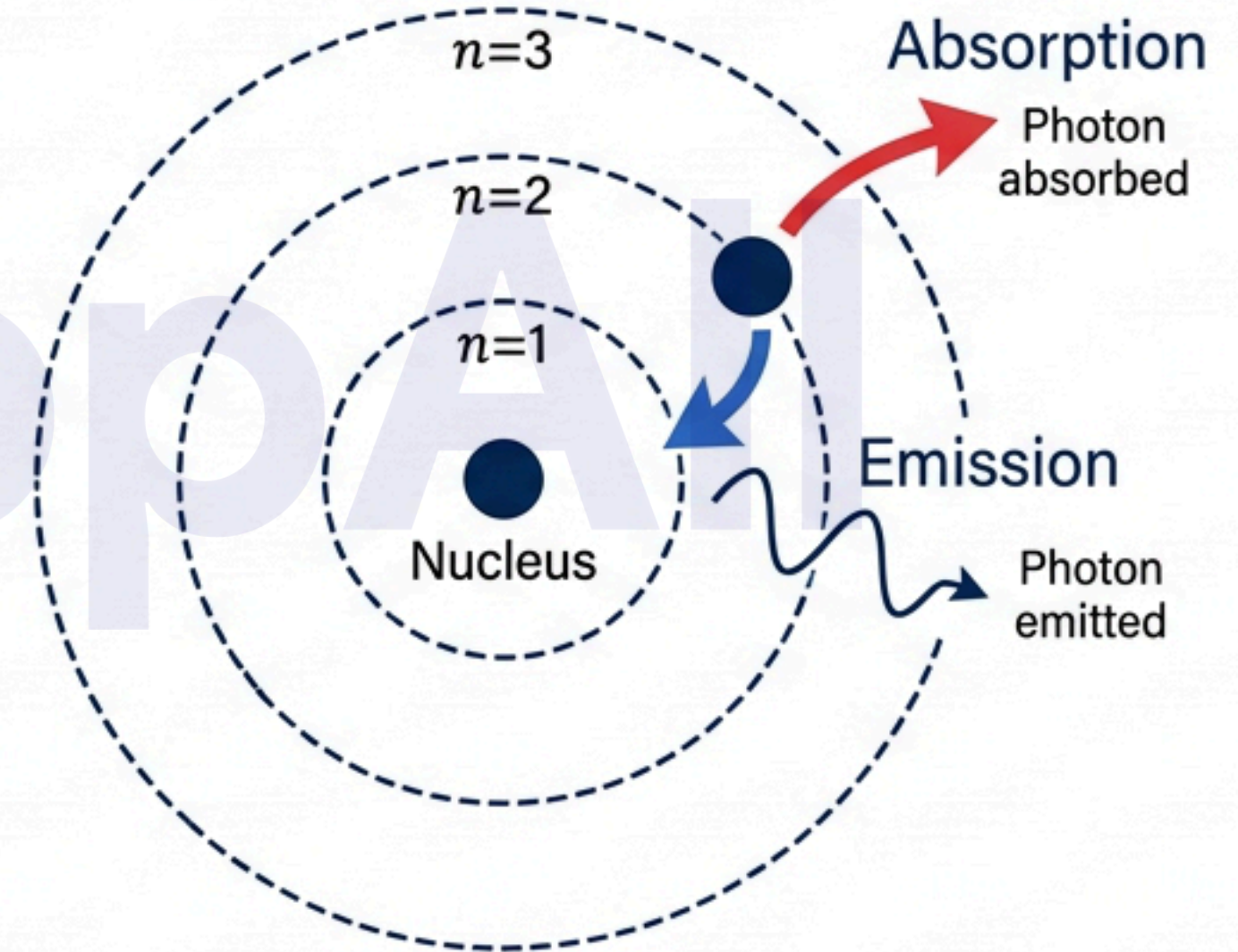
According to Bohr's model, radiation is only absorbed or emitted when an electron jumps between strictly quantized orbits (Principal quantum numbers, n).

Absorption

- Action: Electron moves from a lower n to a higher n .
- Requirement: Energy must be absorbed from the surroundings.
- Math: $n_i < n_f$

Emission

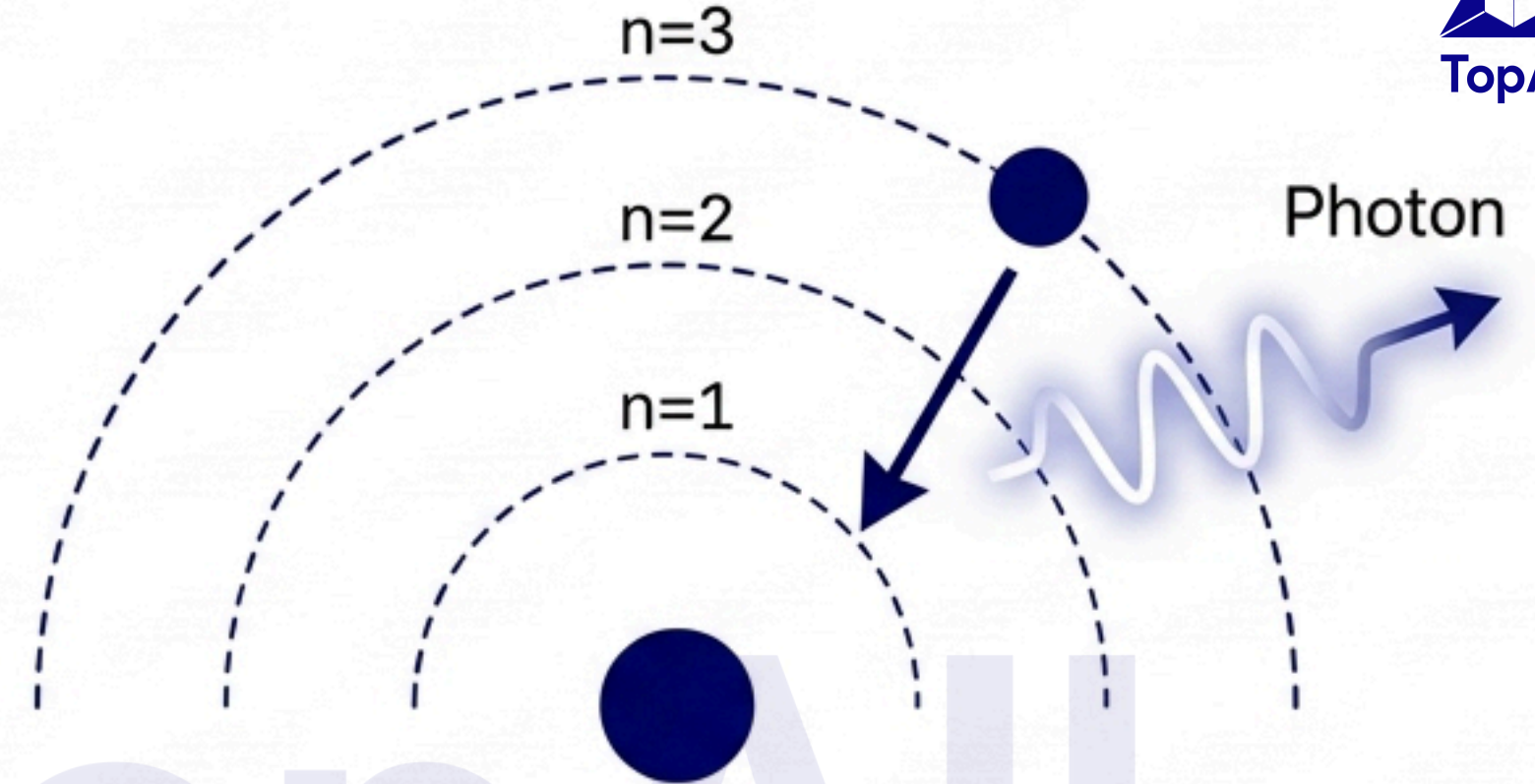
- Action: Electron moves from a higher n to a lower n .
- Result: Energy is emitted as a photon.
- Math: $n_i > n_f$



The precise energy gap between the two specific orbits strictly determines the energy of the resulting radiation.

Emission During Downward Transition

Atoms with electrons in higher, excited states are unstable. To return to a lower energy state, they shed their excess energy.



1. The Drop

The electron transitions from an initial orbit (n_i) down to a final, lower orbit (n_f).

2. The Gap

The exact energy difference is $\Delta E = E_f - E_i$. Because $n_i > n_f$, ΔE is negative, indicating energy release.

3. The Photon

This released energy is instantly ejected as a single photon of electromagnetic radiation.

Note on Intensity: The brightness of a spectral line depends purely on the number of photons of the same frequency being emitted by a large group of atoms simultaneously.



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$$E=mc^2$$



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Translating the Gap into Frequency

By combining Bohr's energy equation with Planck's formula ($\Delta E = h\nu$), we calculate the exact frequency and wavenumber of the emitted light.

Energy Gap (ΔE)

$$\Delta E = 2.18 \times 10^{-18} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \text{ J}$$

◆ $2.18 \times 10^{-18} \text{ J}$ is the Rydberg constant for Hydrogen (R_H).

Frequency (ν)

$$\nu = \frac{\Delta E}{h} = 3.29 \times 10^{15} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \text{ Hz}$$

◆ Derived by dividing ΔE by Planck's constant (h).

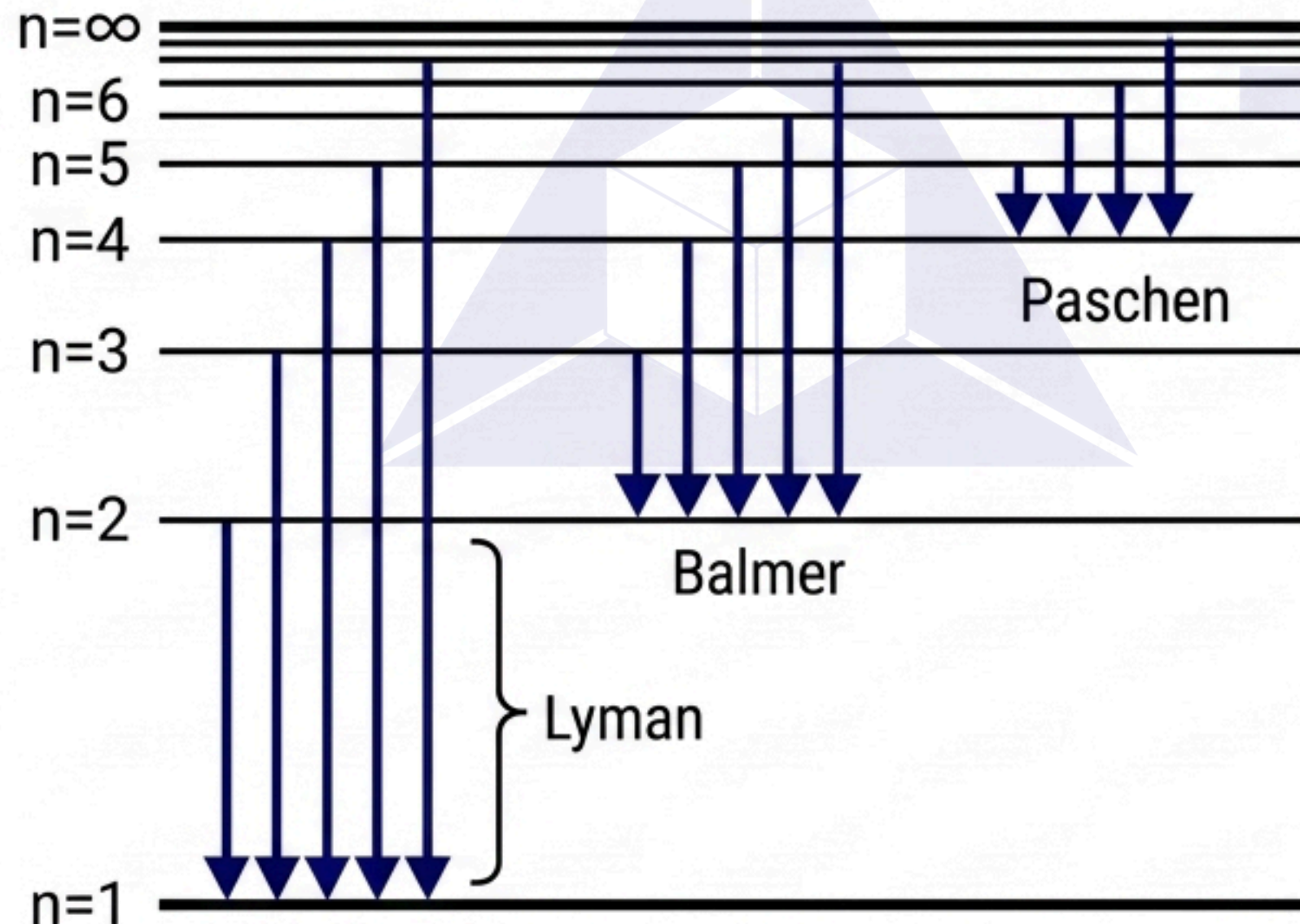
Wavenumber ($\bar{\nu}$)

$$\bar{\nu} = \frac{\nu}{c} = 1.09677 \times 10^7 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \text{ m}^{-1}$$

◆ This theoretical result perfectly matches Rydberg's empirical formula derived from early experiments.

Linking Transitions to Spectral Series

Each distinct series of lines in the hydrogen spectrum corresponds to electrons dropping to one specific baseline orbit.



Lyman Series

Drops to $n_f = 1$ | Originates from $n_i = 2, 3...$ | Region: Ultraviolet

Balmer Series

Drops to $n_f = 2$ | Originates from $n_i = 3, 4...$ | Region: Visible

Paschen Series

Drops to $n_f = 3$ | Originates from $n_i = 4, 5...$ | Region: Infrared

Brackett Series

Drops to $n_f = 4$ | Originates from $n_i = 5, 6...$ | Region: Infrared

Pfund Series

Drops to $n_f = 5$ | Originates from $n_i = 6, 7...$ | Region: Infrared

Final Recap: The Origin of the Spectrum

Why does Hydrogen emit specific lines of light instead of a continuous rainbow?



Energy is strictly locked.

Electrons can only exist in fixed orbits ($n=1, 2, 3\dots$). They cannot hover in between. Therefore, energy is only exchanged in exact, discrete, quantized chunks.



1 Gap = 1 Spectral Line.

Every individual spectral line observed in a lab corresponds to a massive number of electrons making the exact same $n_i \rightarrow n_f$ downward leap.



Quantum math matches reality.

Bohr's theoretical quantum model mathematically proved Rydberg's observational formula, successfully linking microscopic electron behavior to macroscopic visible light.



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