

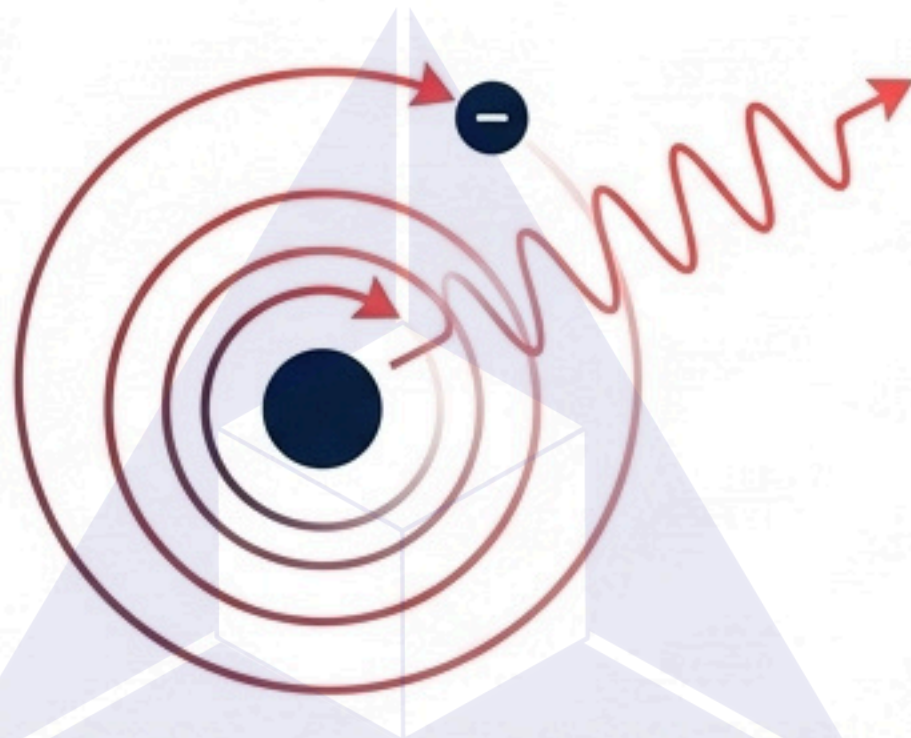
Bohr Model of the Hydrogen Atom

The model that explained atomic stability and discrete spectra



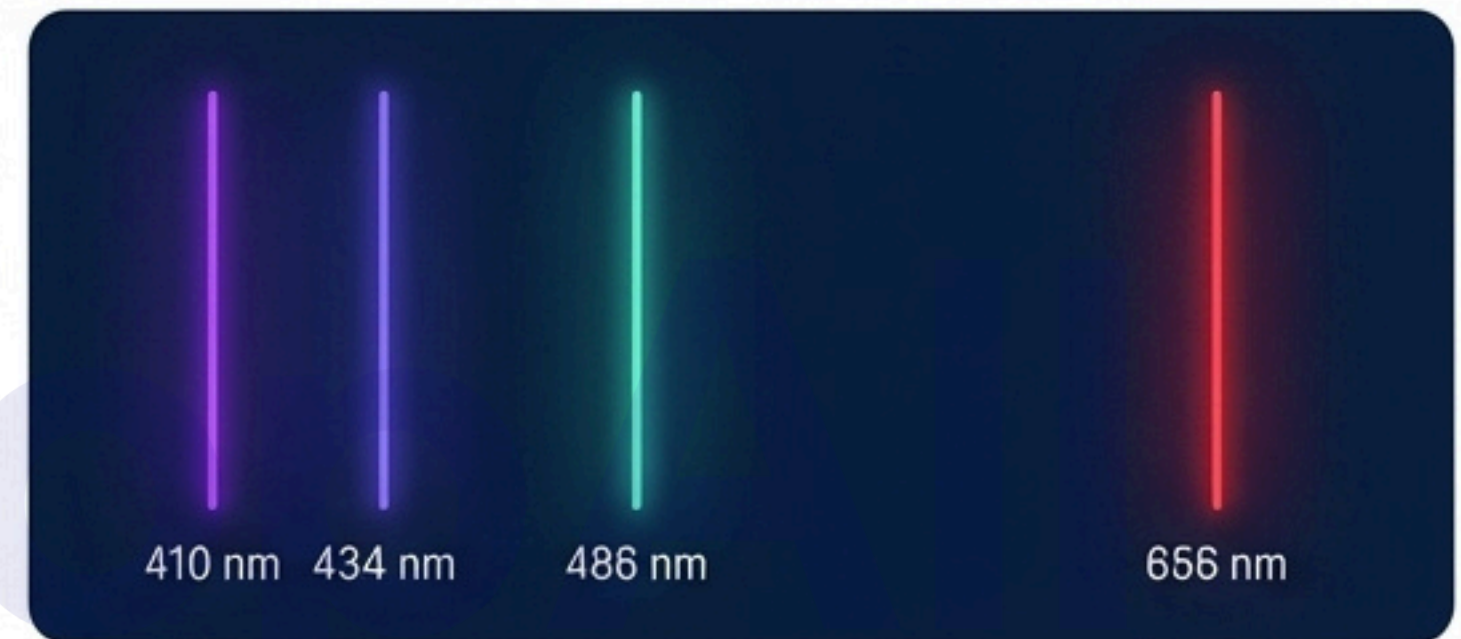
Why the Rutherford Model Was Not Enough

The Instability Problem



According to classical electromagnetic theory, an accelerating charged particle must continuously emit radiation. A revolving electron would lose energy, spiral inward, and collapse into the nucleus. Atoms would not exist.

The Spectral Problem

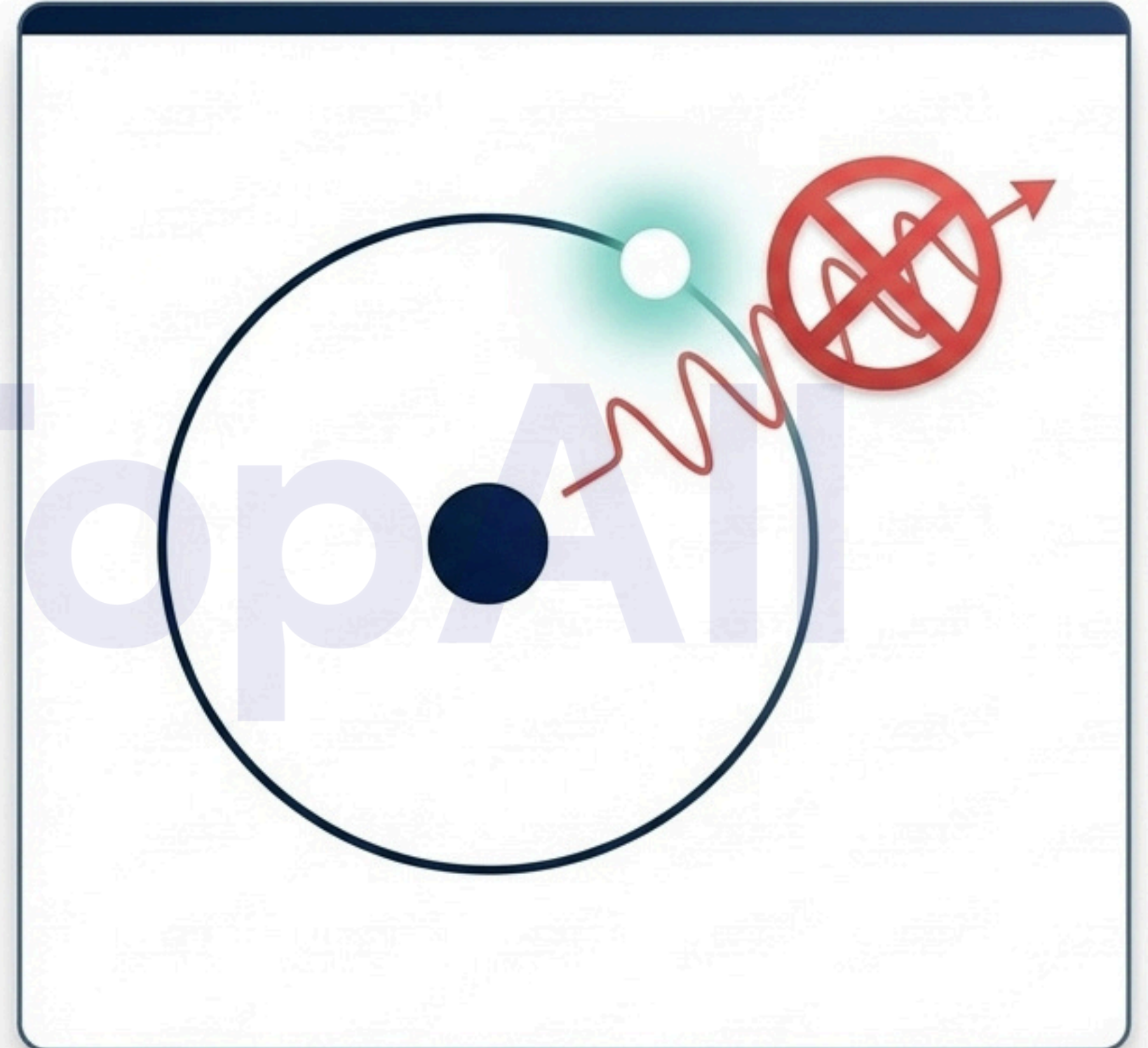


Classical physics predicts that as the electron spirals inward, its frequency changes continuously, emitting a continuous rainbow spectrum. However, real gases emit radiation only at specific, discrete wavelengths (a line spectrum).

Postulate I: Stationary Orbits

Bohr proposed a radical departure from classical mechanics: an electron can revolve in certain stable orbits without emitting radiant energy.

These specific, non-radiating paths are called **Stationary States**.





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$$E=mc^2$$



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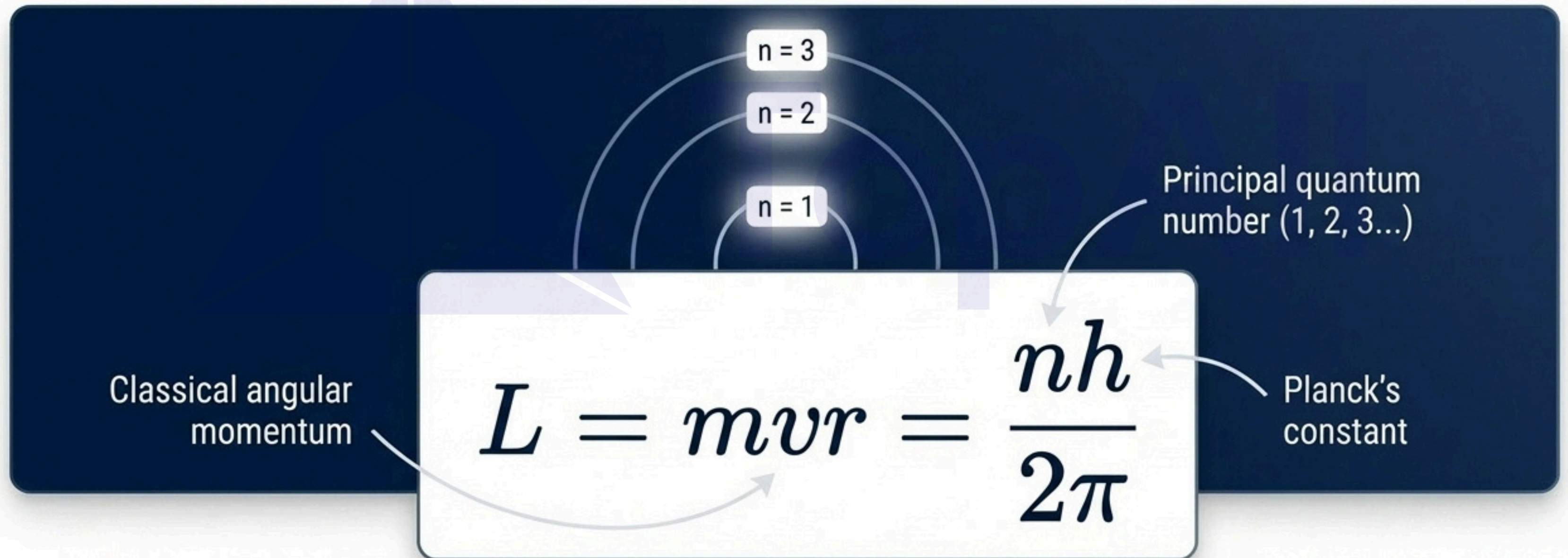
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Postulate II: Quantisation of Angular Momentum

An electron only orbits where its angular momentum (L) is an exact integral multiple of $h/2\pi$.



The diagram illustrates the quantization of angular momentum. It features three concentric circular orbits labeled $n=1$, $n=2$, and $n=3$ from innermost to outermost. A label 'Principal quantum number (1, 2, 3...)' points to these orbits. Below the orbits, a white box contains the equation $L = mvr = \frac{nh}{2\pi}$. An arrow from the text 'Classical angular momentum' points to the mvr term in the equation. Another arrow from the text 'Planck's constant' points to the h in the numerator of the fraction.

Classical angular momentum

$$L = mvr = \frac{nh}{2\pi}$$

Principal quantum number (1, 2, 3...)

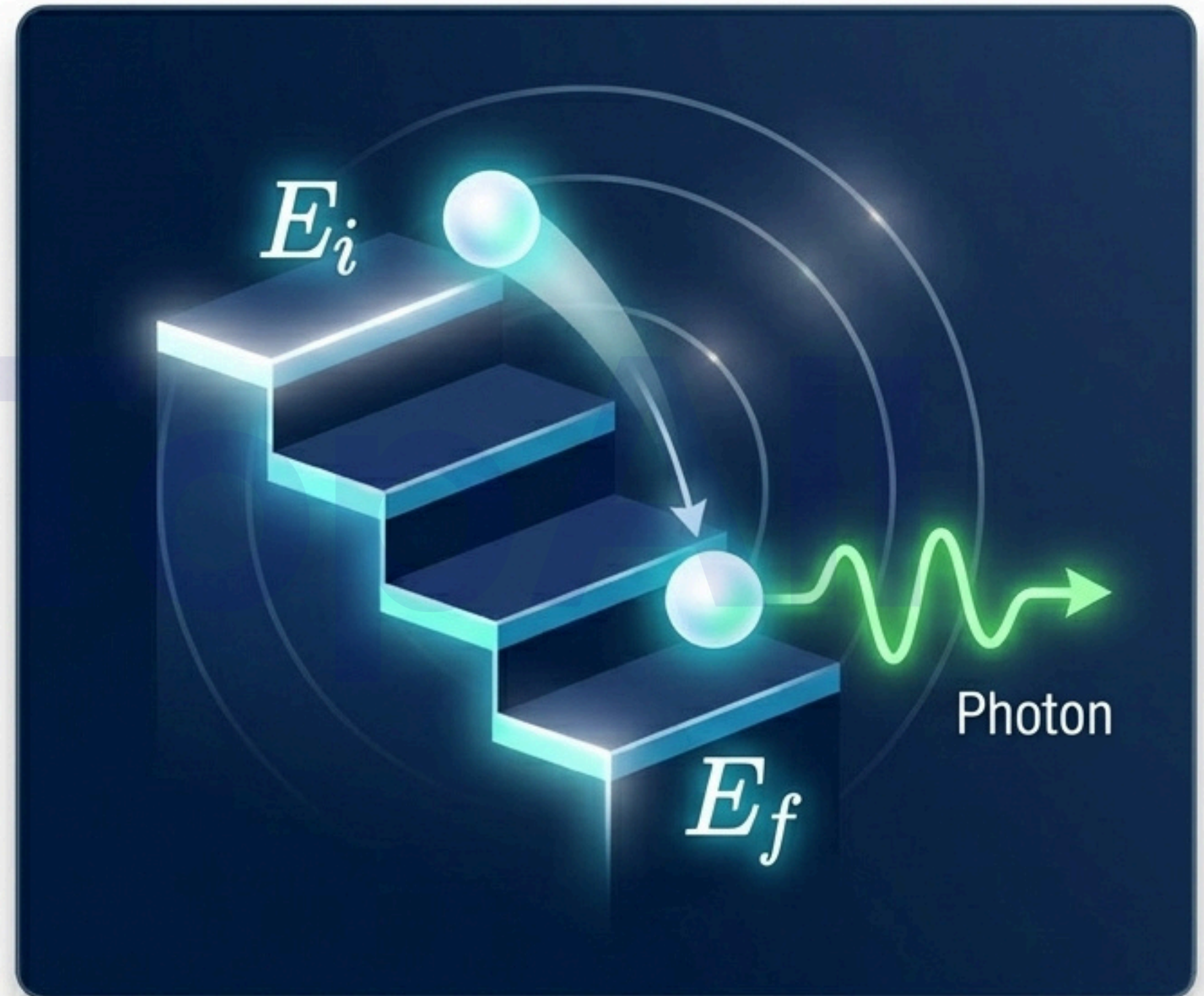
Planck's constant

Postulate III: Energy Transitions

Radiation is only emitted or absorbed when an electron “jumps” between stationary orbits.

$$h\nu = E_i - E_f$$

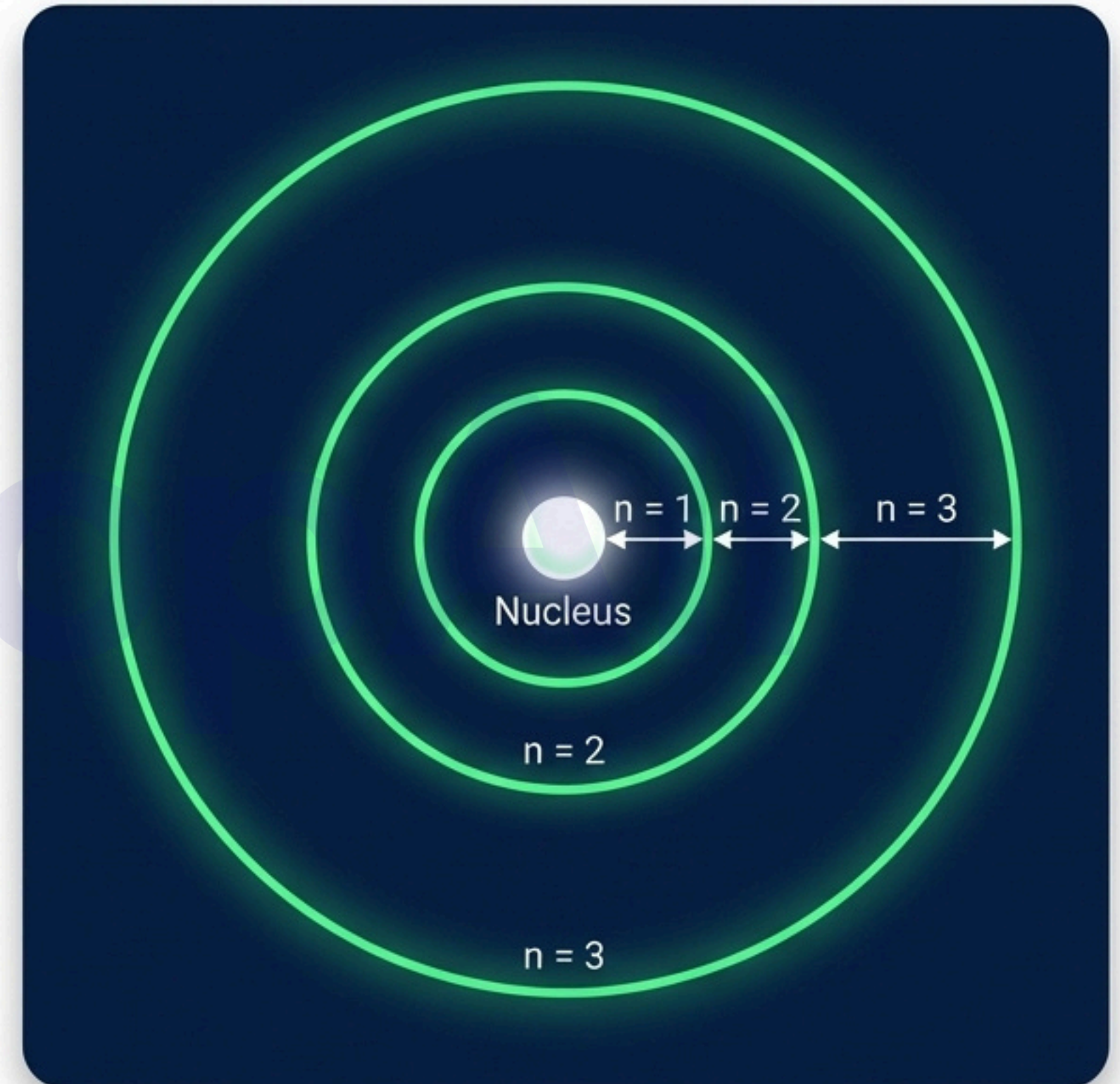
The energy of the emitted photon ($h\nu$) exactly equals the energy difference between the initial and final states.



Deriving Allowed Radii

By balancing classical centripetal force ($F_c = F_e$) with Bohr's angular momentum condition, we prove that only specific orbital distances are allowed.

Key Insight: $r_n \propto n^2$

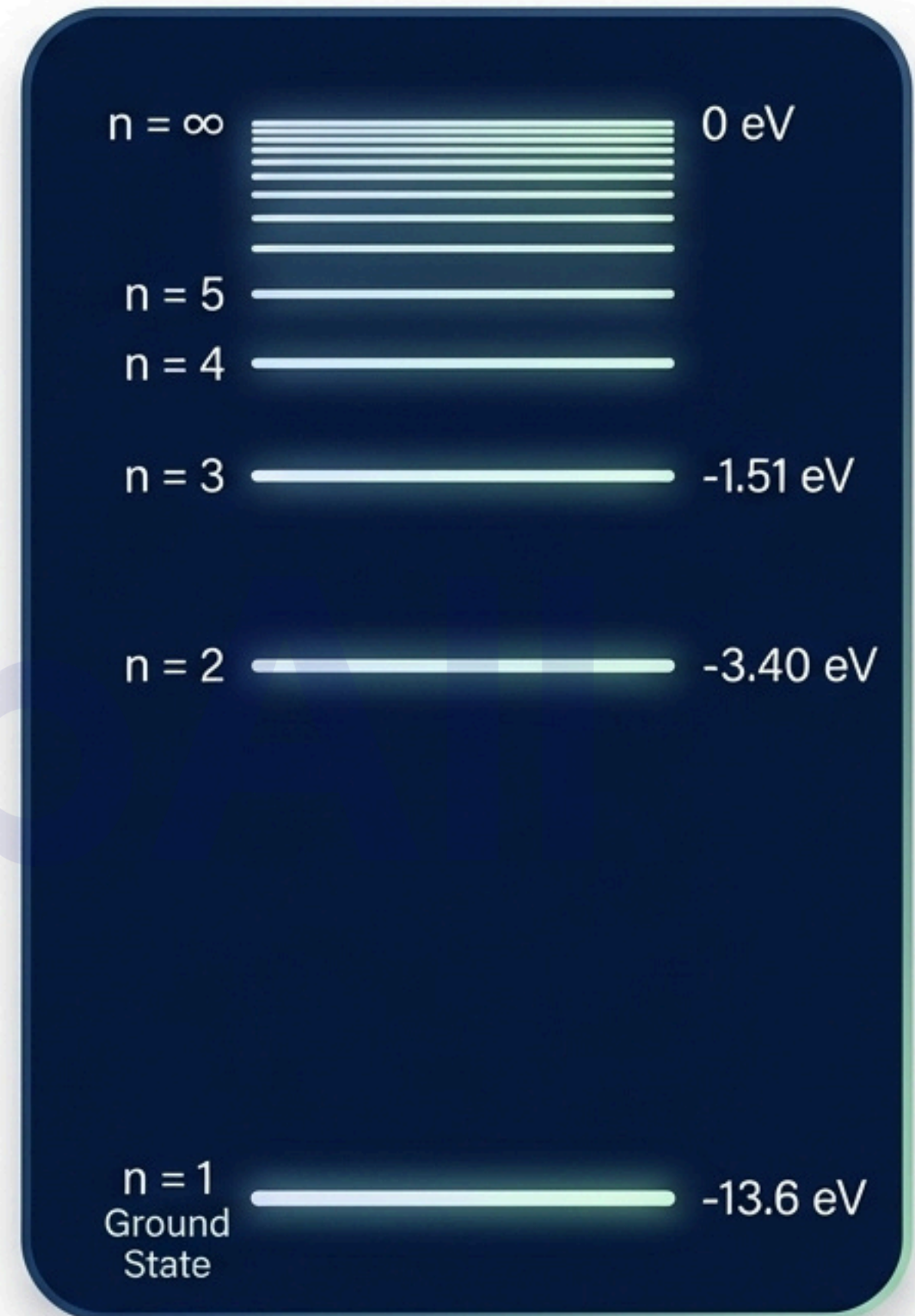


Quantised Energy Levels

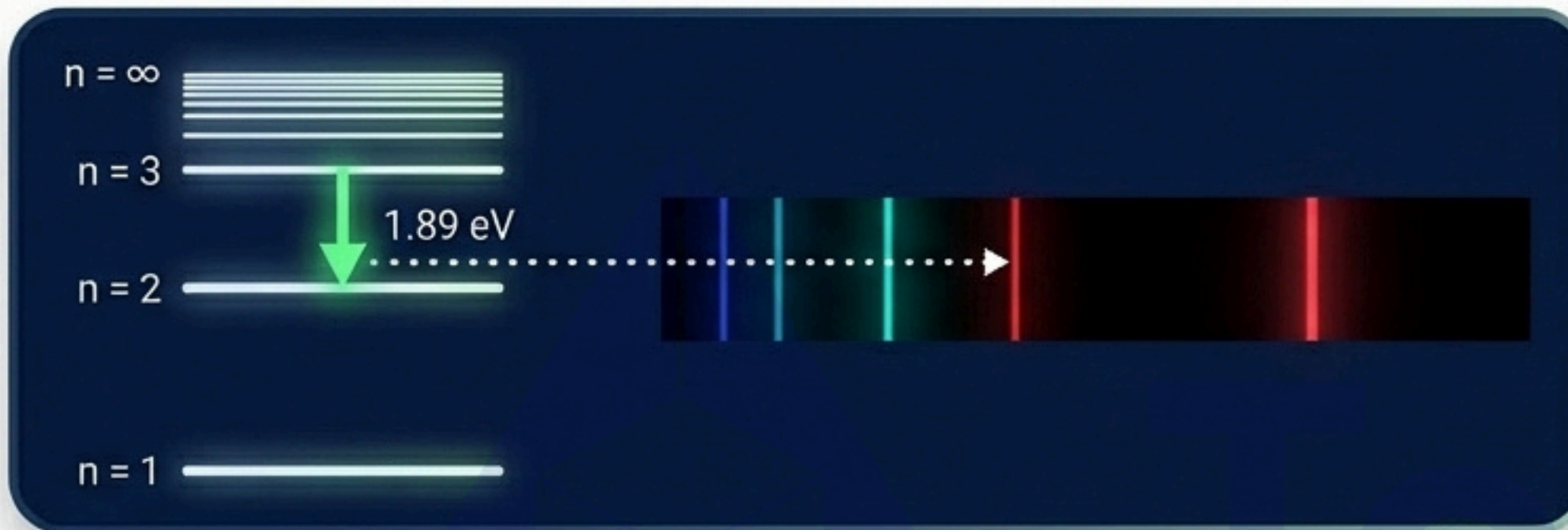
The total energy of the electron is quantised and negative, proving the electron is bound to the nucleus.

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

- $n=1$ (Ground State) = -13.6 eV
- $n=2$ = -3.4 eV
- $n=3$ = -1.51 eV



Solving the Hydrogen Mystery



The specific energy gaps **perfectly match** the **distinct** frequencies of the **hydrogen line spectrum**.

Success / Key Takeaways

- ✓ **Stability Explained:** The electron cannot drop below the $n=1$ ground state. It cannot spiral into the nucleus.
- ✓ **Spectra Explained:** Discrete energy levels mandate discrete photon emissions, creating distinct spectral lines instead of continuous rainbows.

Check Your Understanding

Summary

1. Electrons exist in stable, non-radiating stationary orbits.
2. Angular momentum is quantised:
$$L = nh / 2\pi$$
3. Energy levels scale by inverse square: $E_n = -13.6 / n^2 \text{ eV}$

Self-Check

What happens to the energy gap between adjacent orbits as n increases?

They get closer together

How much energy is required to ionise a Hydrogen atom from its ground state?

+13.6 eV

Does r_n scale linearly with n ?

No, $r_n \propto n^2$

Next: Energy Levels & Spectral Series →



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