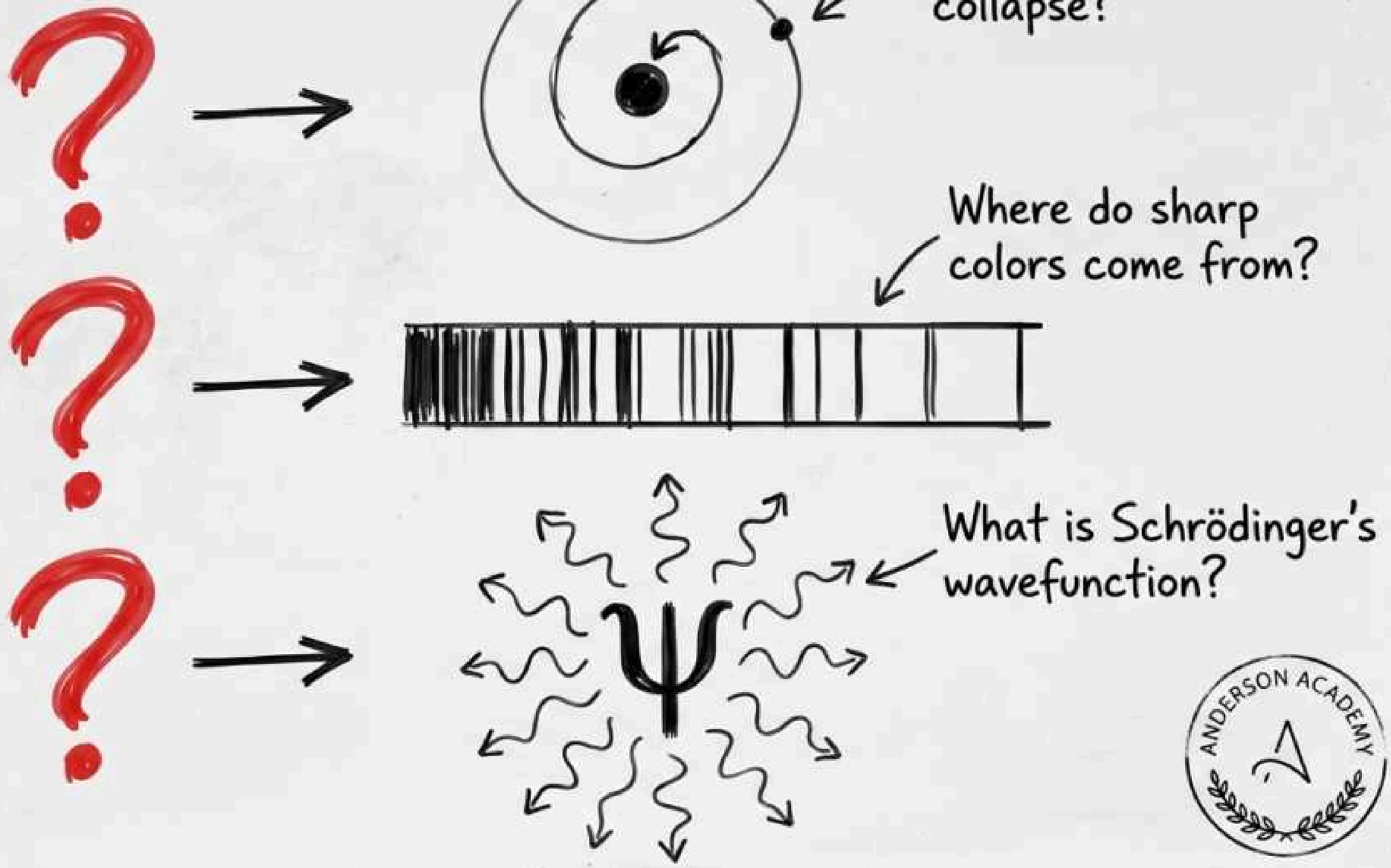
PHYS 220 B Lecture 3: From the Bohr Atom to the Wavefunction

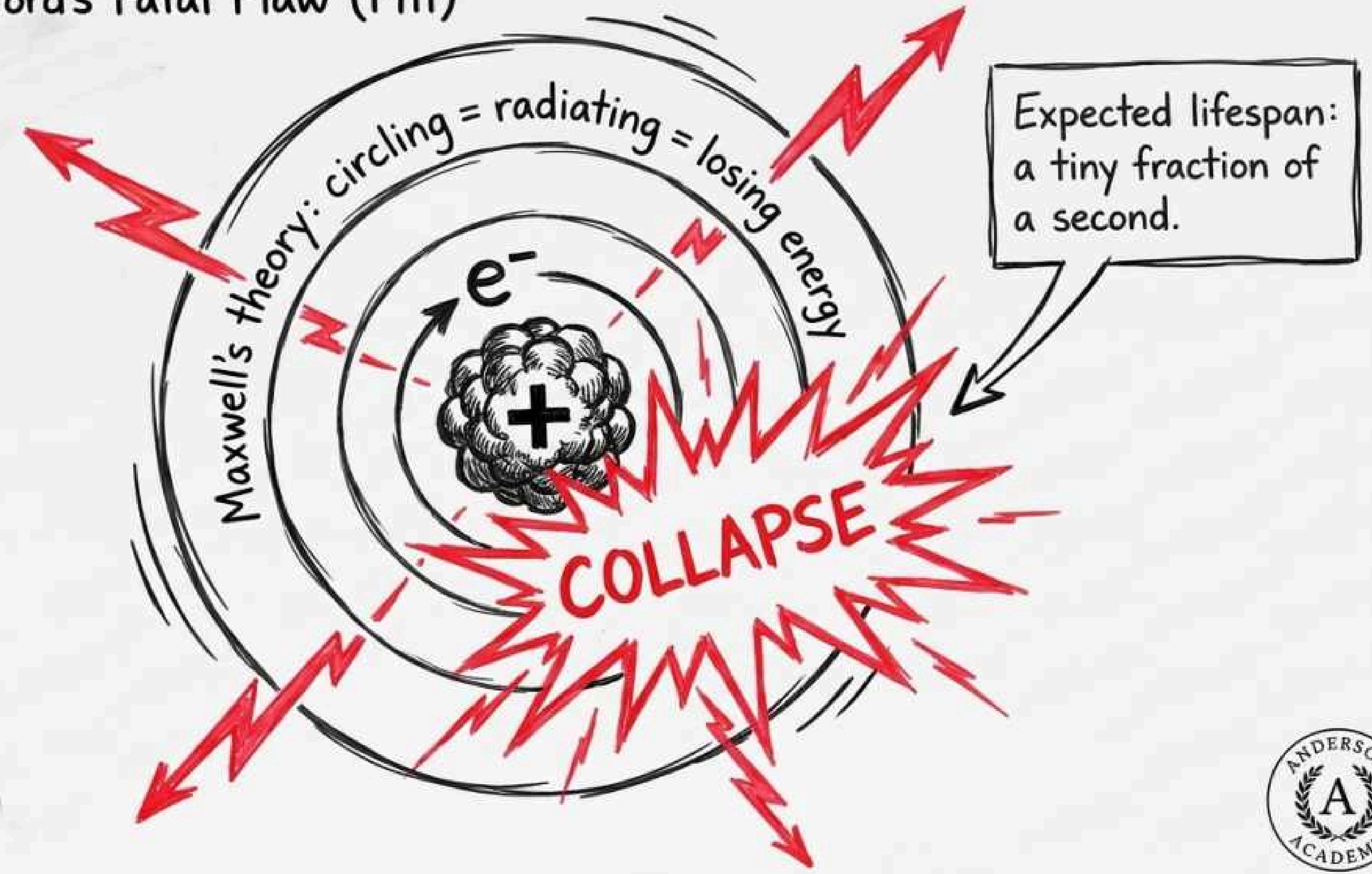
Anderson Academy Extracurricular Physics



3.1 - The Three Questions

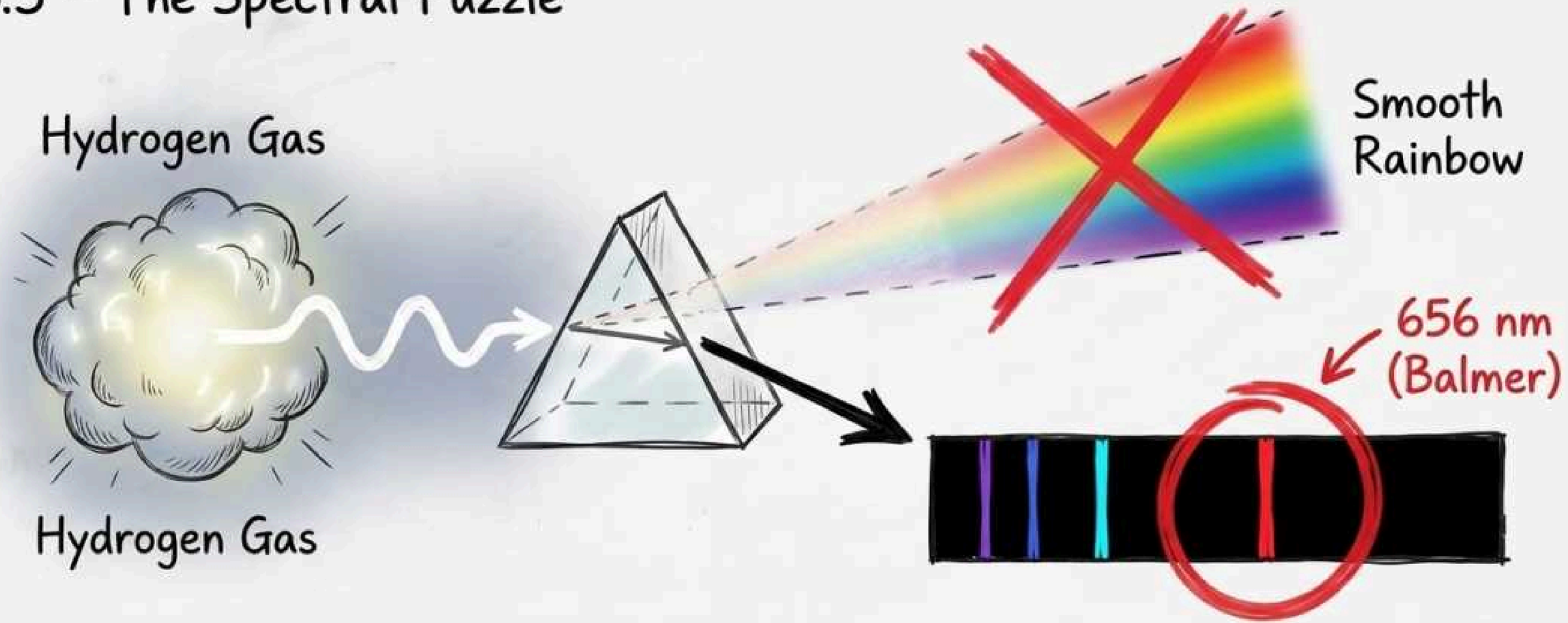


3.2 – Rutherford's Fatal Flaw (1911)



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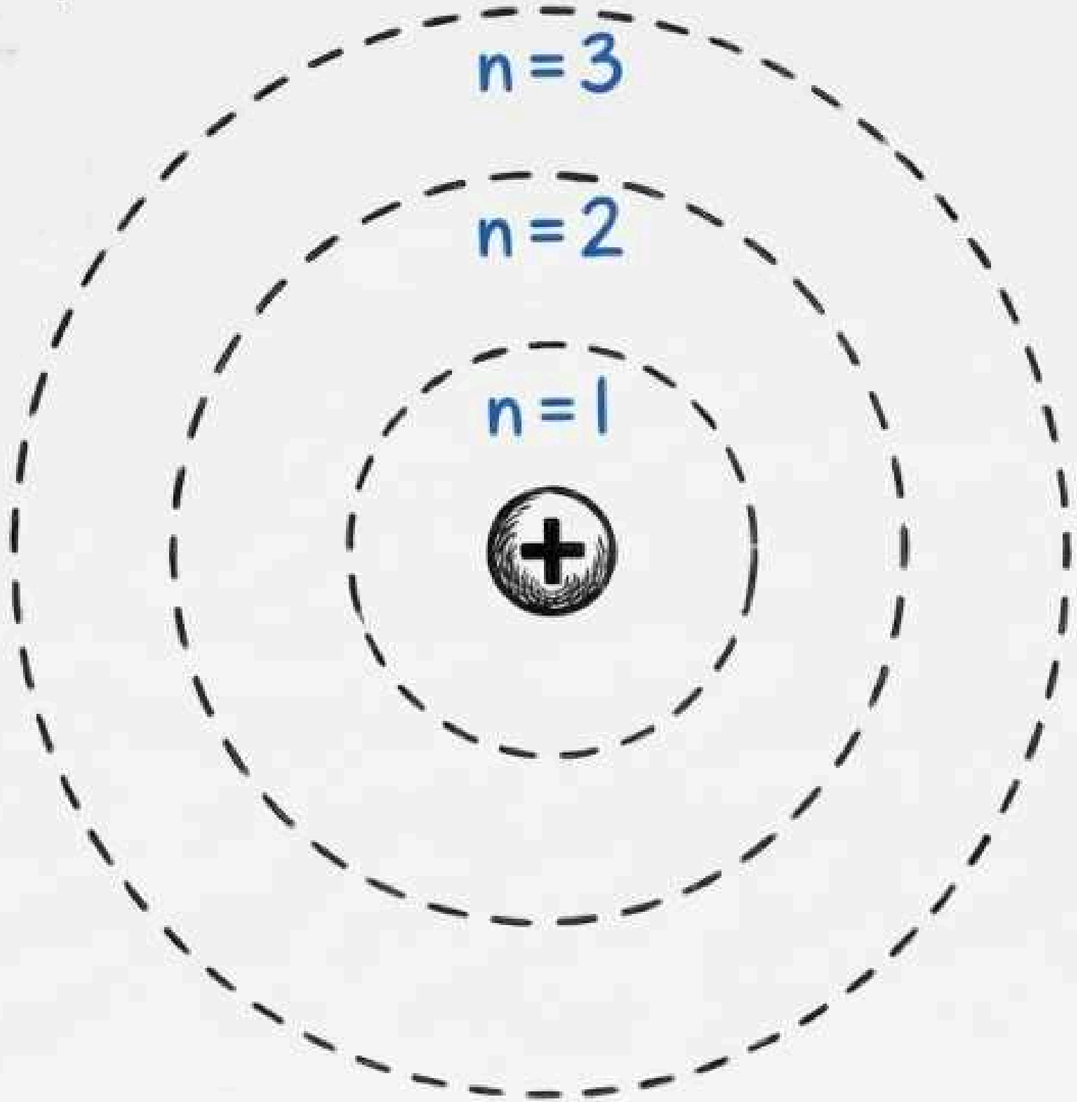
3.3 - The Spectral Puzzle



Atoms emit specific sharp colors, like a fingerprint—not a smooth rainbow.

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3.4 - Bohr's Bold Proposal (1913)

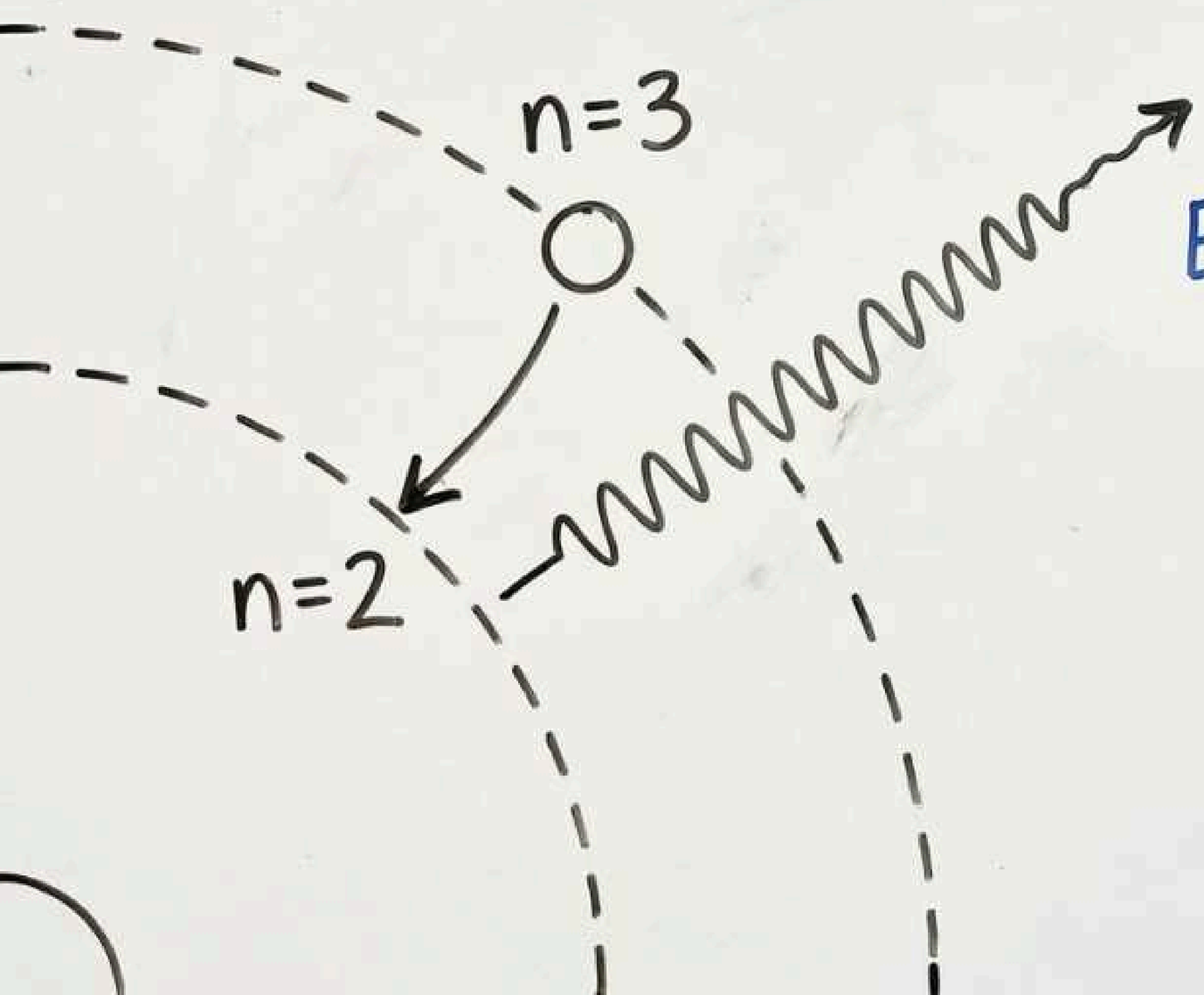


$$E_n = \frac{-13.6\ eV}{n^2}$$

- Allowed orbits ONLY.
- No radiation while in orbit.
- Light ONLY emitted during jumps.

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3.5 - The Math of the Red Line



$$hf = E_{upper} - E_{lower}$$

$$E_3 - E_2 = 13.6 \times \left(\frac{1}{4} - \frac{1}{9}\right) eV$$

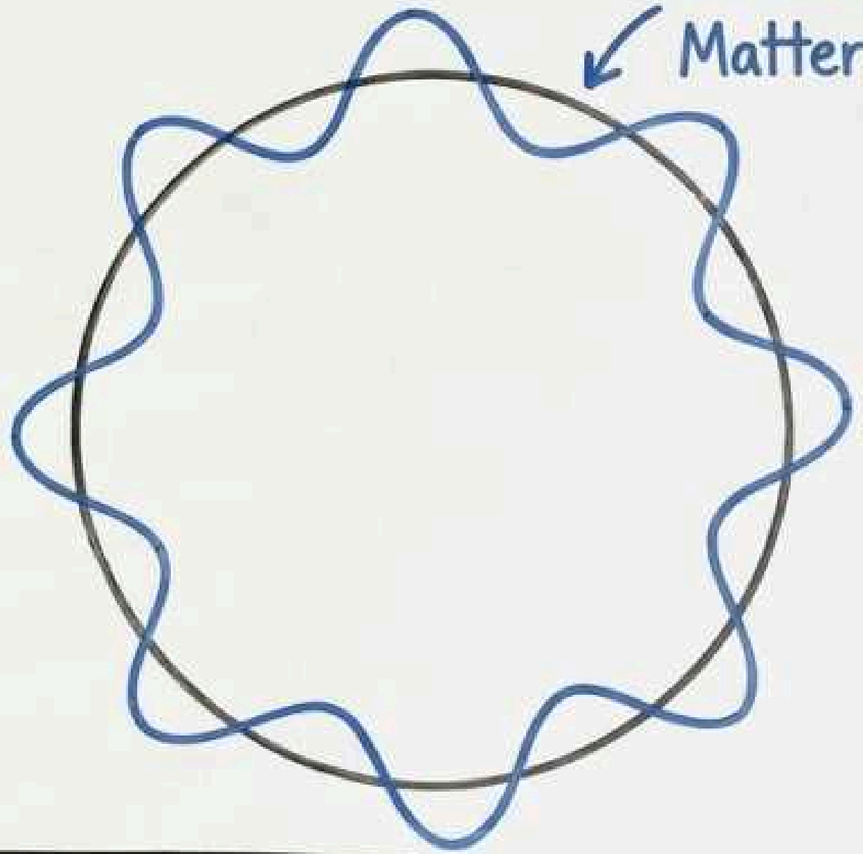
$$= 1.89\ eV$$

$$\rightarrow 656\ nm$$

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3.6 - An Uneasy Triumph



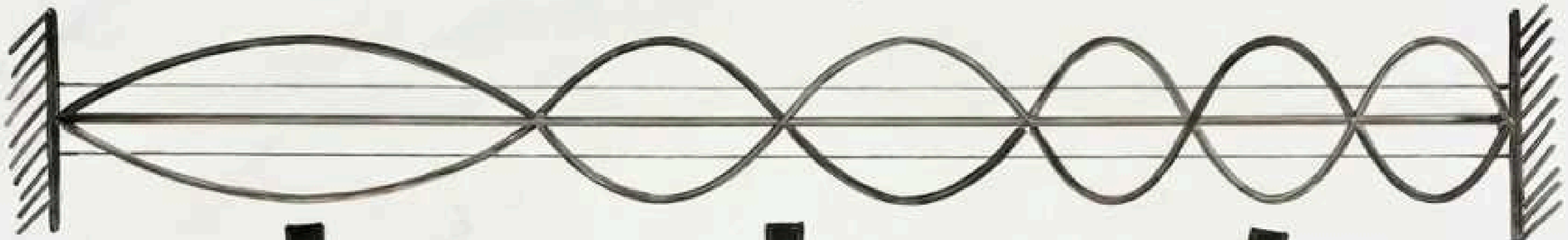


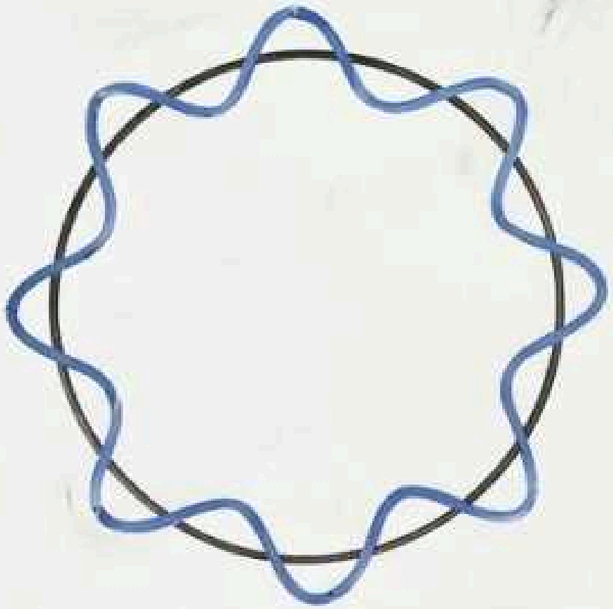
An electron wave must fit an exact whole number of wavelengths around the path. Otherwise, the wave cancels itself out.

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3.7 - Standing Waves

Macro: Fixed String = Harmonics





Confinement = Quantization.
The pull of the nucleus confines the electron wave, allowing only certain "standing wave" patterns to survive.

3.8 - Two Complete Theories (1925-1926)

$$\begin{bmatrix} \Sigma_i & 0 & \Sigma_j & \Sigma_k & \cdots & \Sigma_m \\ 0 & \Sigma_i & \Sigma_j & \Sigma_l & \cdots & \Sigma_m \\ 1 & j & k & m & \cdots & \Sigma_n \\ \Sigma_i & \Sigma_k & \Sigma_l & \Sigma_k & \cdots & H_p \\ \Sigma_i & \Sigma_k & \Sigma_l & xn_l & \cdots & H_m \\ \vdots & \vdots & \vdots & \vdots & \ddots & \vdots \\ \Sigma_n & \Sigma_m & X_{n2} & X_{nn} & \cdots & H_{rc} \end{bmatrix}$$

Heisenberg's Matrix Mechanics
(Observables only)

=

$$i\hbar\frac{\partial\psi}{\partial t} = -\frac{\hbar^2}{2m}\nabla^2\psi + V\psi$$

Schrödinger's Wave Equation
(Changes in space/time)

Completely different math.
Identical predictions.

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3.9 - Schrödinger's Hydrogen

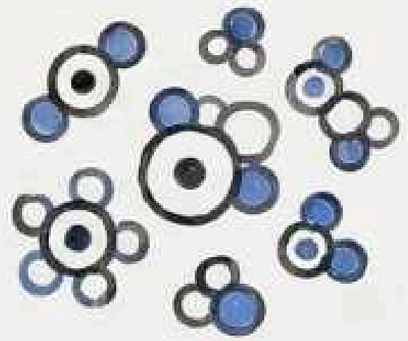
Schrödinger's Wave Equation



Predicts Spectral Brightness

Bohr's Exact Energy Levels

Emerge naturally as standing waves, no ad-hoc rules needed.

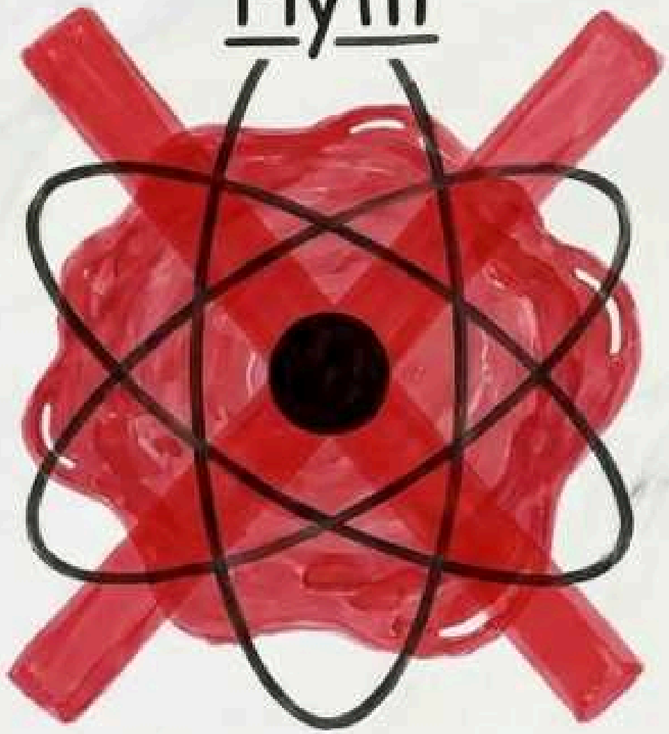


Works for ALL Atoms and Molecules

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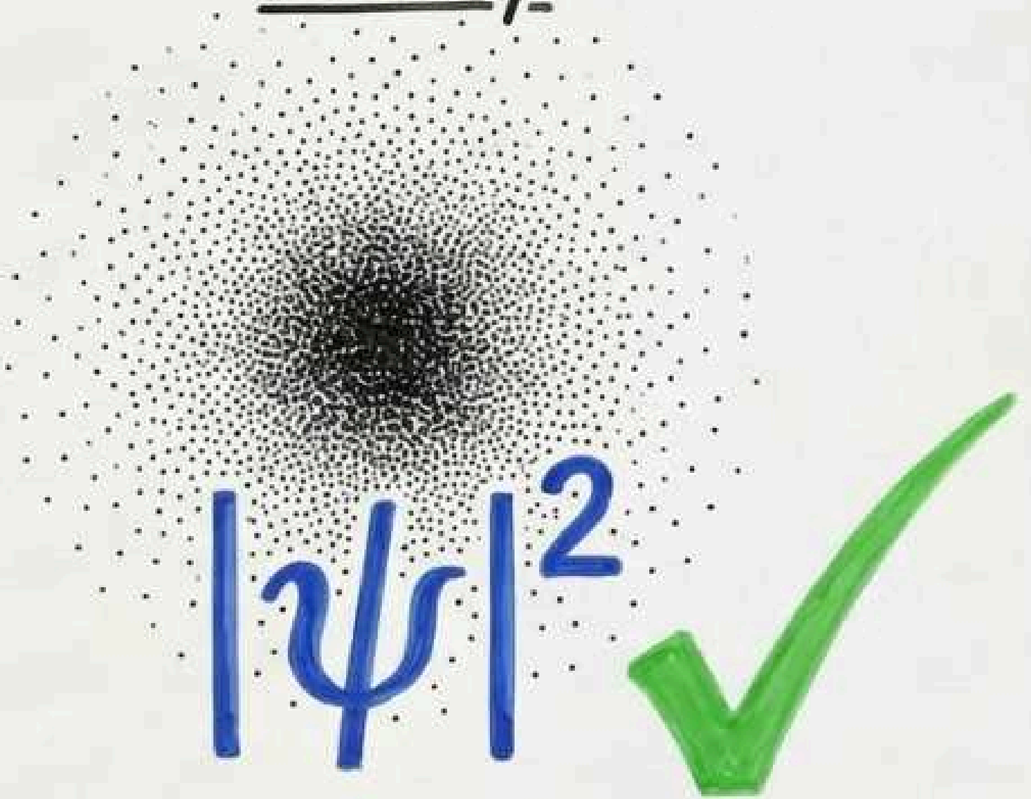
3.10 - The Born Rule (1926)

Myth



Not a smeared cloud of matter.

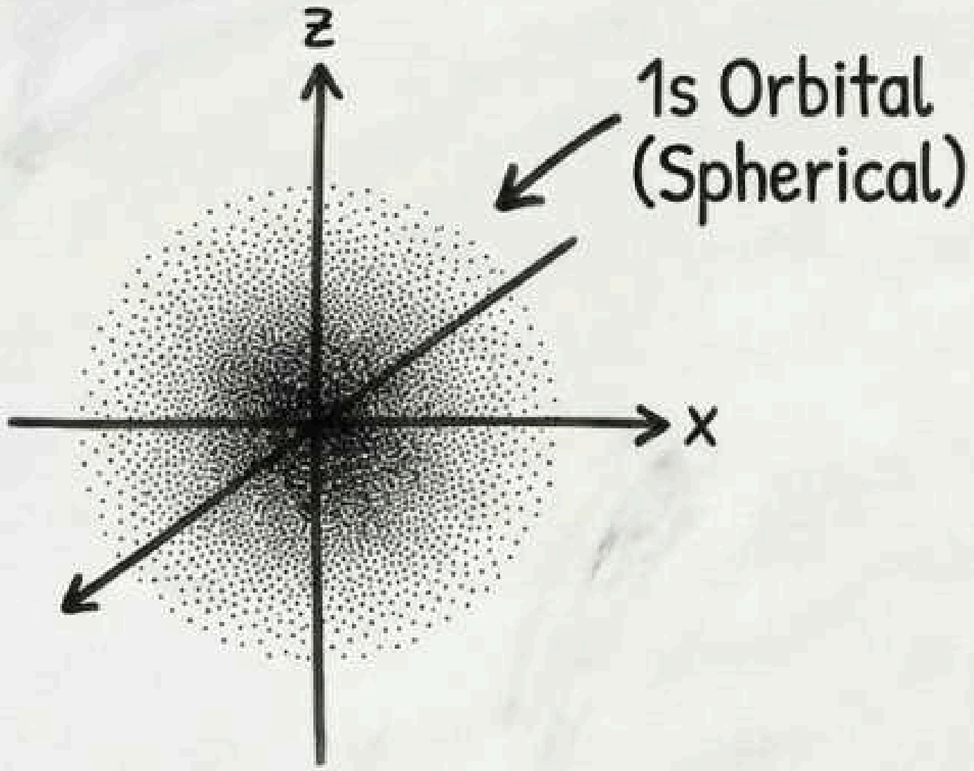
Reality



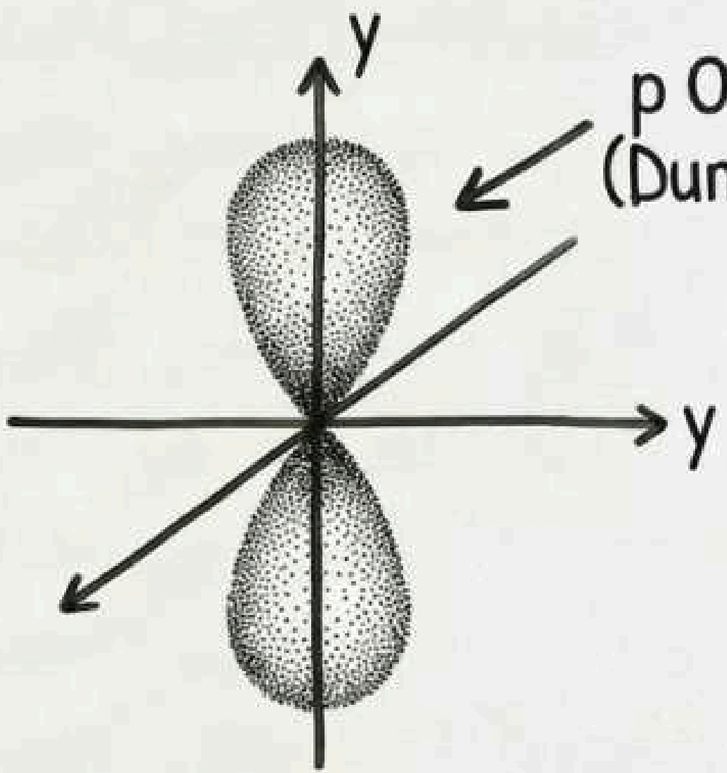
The wavefunction is not physical. Its square ($|\psi|^2$) is the probability of finding the electron if you look.

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3.11 - From Orbits to Orbitals



1s Orbital
(Spherical)



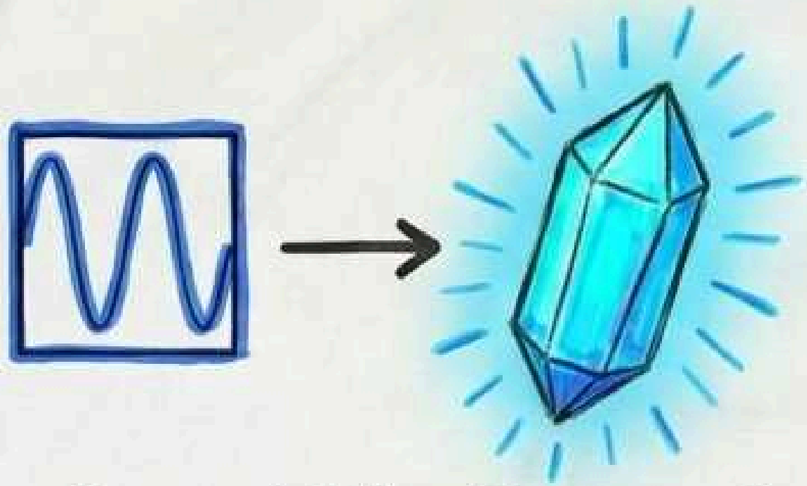
p Orbital
(Dumbbell)

- Orbitals = standing-wave probability patterns.
- Dictated by quantum numbers.
- These shapes determine how atoms bond and create all molecular structure.

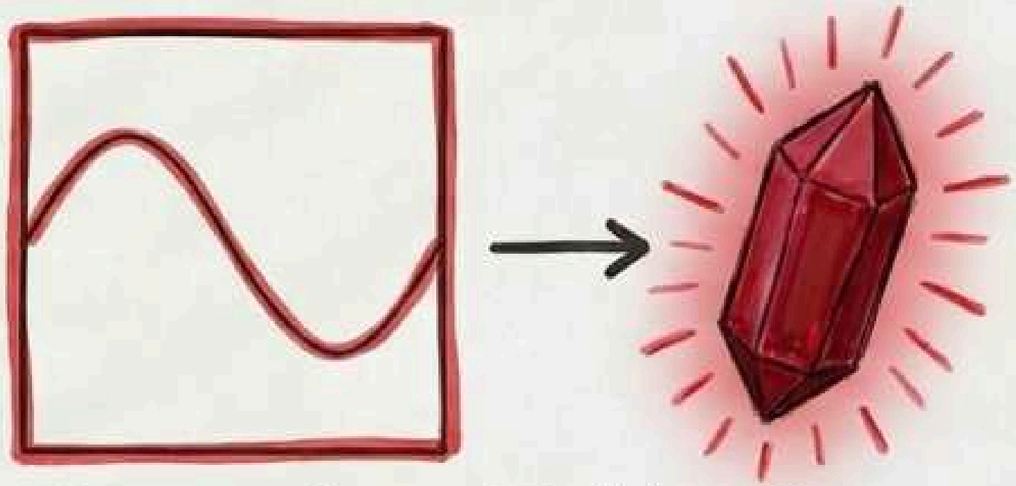
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3.12 - Particle in a Box

$$E_n = \frac{n^2h^2}{8mL^2}$$



Smaller Box = Wider Energy Gaps
= Bluer Light

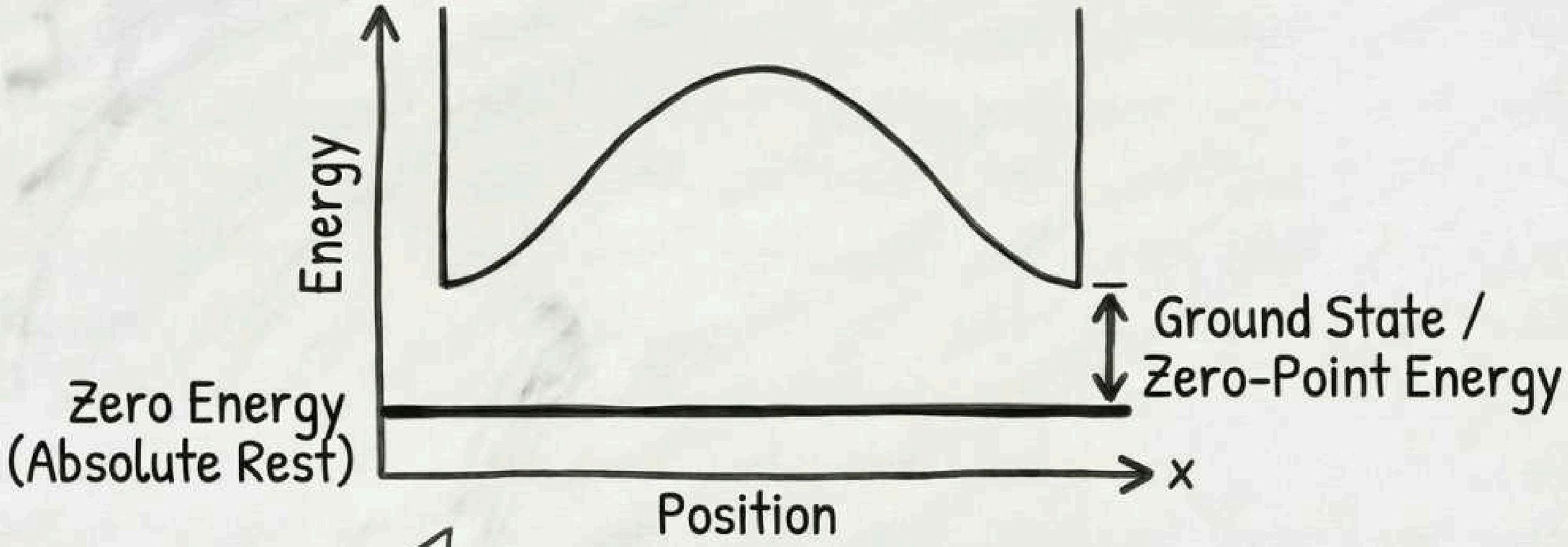


Larger Box = Tighter Gaps
= Redder Light

2023 Nobel Prize in Chemistry:
Quantum Dots use size to control color.

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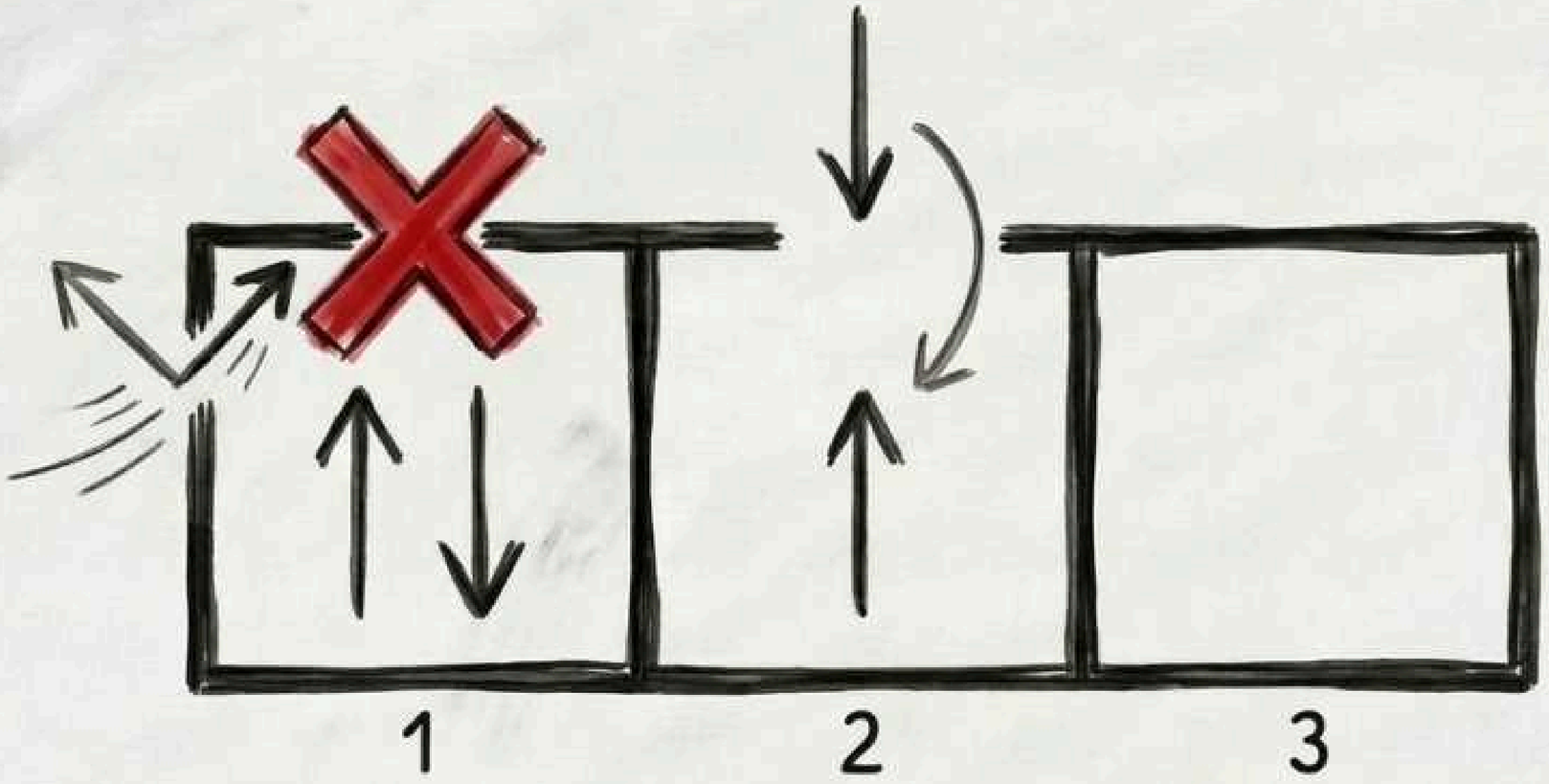
3.13 - Zero-Point Energy



A confined wave must have at least one hump.
It can never be perfectly still.
Atoms keep moving even at absolute zero.
(This is why electrons can't fall into the nucleus!)

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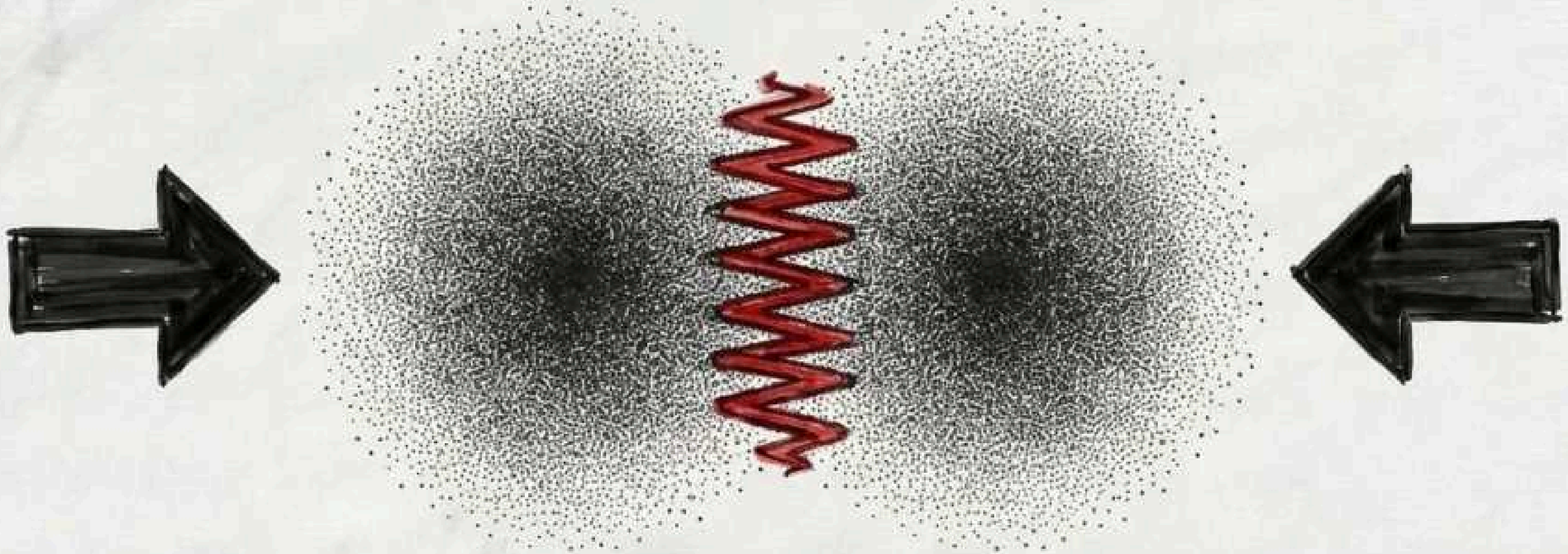
3.14 - Pauli Exclusion Principle (1925)



No two electrons can occupy the same quantum state.

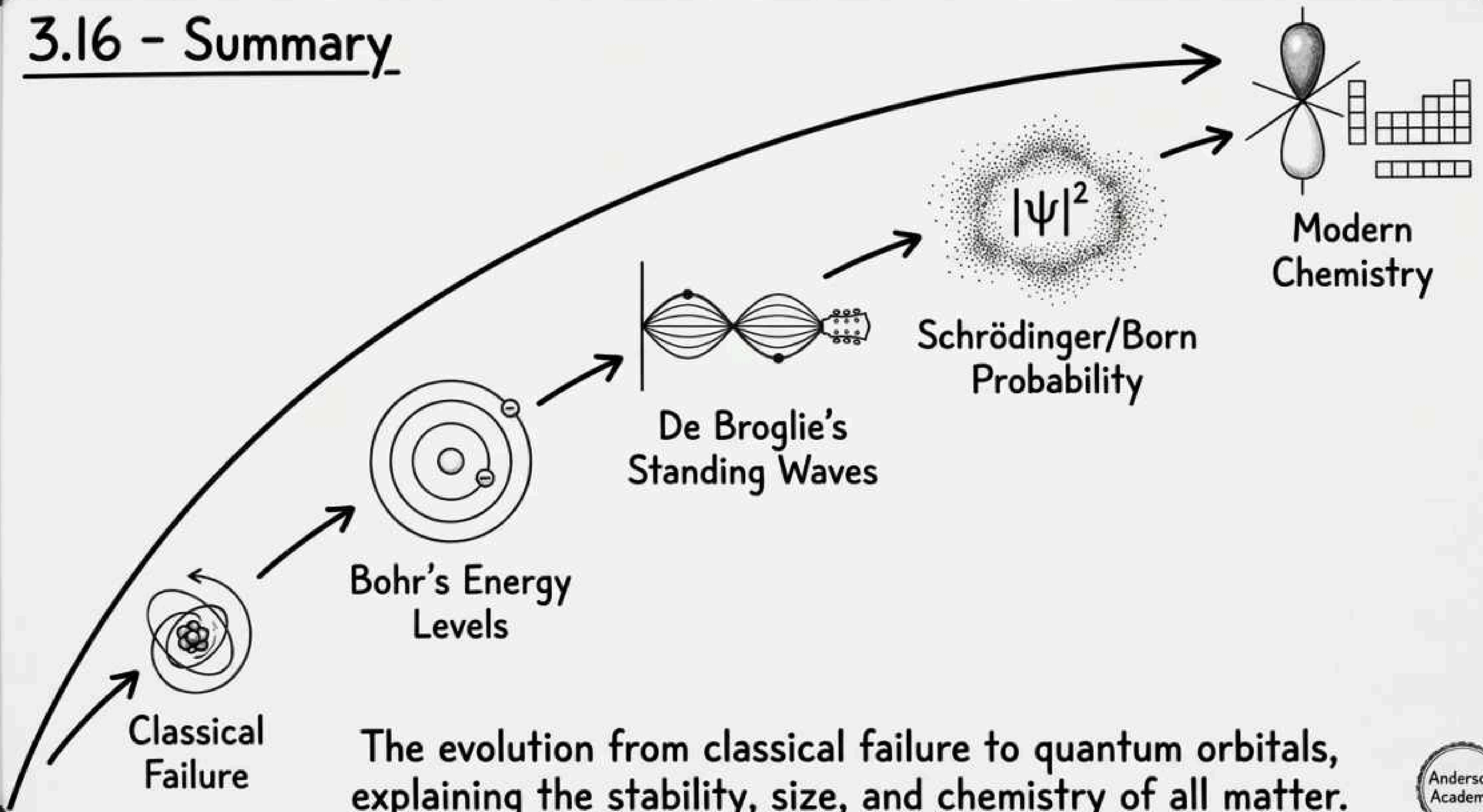
Electrons build up in "shells", creating the rows and columns of the periodic table.

3.15 - Why Matter is Solid



Pressing atoms together forces electrons into higher states (because of Pauli Exclusion). This costs energy and pushes back. The firmness of a table under your hand is a quantum effect.

3.16 - Summary



The evolution from classical failure to quantum orbitals, explaining the stability, size, and chemistry of all matter.

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