Atomic Line Spectra: the Bohr model

Niels Bohr's great contribution to science was in building a simple model of the atom. It was based on observations of the SHARP LINE SPECTRA of excited atoms.

- Excited atoms emit light of only certain wavelengths
- The wavelengths of emitted light depend on the element.

Visible lines in the H atom spectrum are called the BALMER series.

Atomic Spectra and Bohr

One (incorrect) view of atomic structure in early 20th century was that an electron (e⁻) traveled about the nucleus in an orbit.

That view:
1. Any orbit should be possible and so should any energy.
2. But a charged particle moving in a circle should emit energy.
   The end result should be destruction!

Bohr said that this classical view was wrong. He saw the need for a new theory — now called QUANTUM or WAVE MECHANICS.

- An e⁻ can only exist in certain discrete orbits — called stationary states.
- An e⁻ is restricted to QUANTIZED (discrete) energy states.
- The energy of a state = - (Rhc)/n² = - (const)/n²
  where n = quantum no. = 1, 2, 3, 4, ..., and the constants R = 1.1×10⁷ m⁻¹, h = 6.6×10⁻³⁴ J·sec and c = 3.0×10⁸ m/sec
Atomic Spectra and Bohr

Energy of a quantized state = $-\frac{Rhc}{n^2}$

- Only orbits where $n = \text{integer number}$ are permitted.
- Radius of allowed orbits = $0.0529 \cdot n^2 \text{ nm}$
- Results can be used to explain atomic spectra, at least for simple atoms.

Calculating $\Delta E$ for $e^{-}$ "falling" from higher energy level ($n = 2$) to lower energy level ($n = 1$).

$\Delta E = E_{\text{final}} - E_{\text{initial}} = -\frac{Rhc}{(1/2^2)} - (1/2^2)$

$\Delta E = -(3/4)Rhc$

The atom’s final energy is less than its initial energy, thus the atom is losing energy.

If $e^{-}$’s are in quantized energy states, then $\Delta E$ between states can have only certain values. This explains sharp line spectra.

$\Delta E = -(3/4)Rhc$

The values of $R$, the Rydberg constant, $h$, the Planck constant and $c$ are known from experiment.

$Rhc = 2.179 \times 10^{-18} \text{ J/atom}$

so, $E$ of emitted light = $(3/4)Rhc = 1.63 \times 10^{-18} \text{ J}$. This corresponds to a frequency $\nu = E/h$ of $2.47 \times 10^{15}$ sec$^{-1}$ and $\lambda = c/\nu = 121.6 \text{ nm}$.

This is exactly in agreement with experiment!
Bohr received the Nobel Prize in 1922 for his theory. However, there are problems —
- theory is successful only for H.
- quantum idea artificially introduced.
- So, we go on to QUANTUM or WAVE MECHANICS to understand the atom.

de Broglie (1924) proposed that all moving objects have wave properties.
For electrons orbiting a nucleus, he theorized that a standing wave is set up in each orbit.
Because each orbit has a predicted radius, the wavelength of a moving particle is \[ \lambda = \frac{h}{mv} \] (v is velocity).

Baseball (115 g) at 100 mph:
\[ \lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34} \text{ J} \cdot \text{sec}}{(0.115 \text{ kg})(45 \text{ m/sec})} = 1.3 \times 10^{-34} \text{ m} = 1.3 \times 10^{-25} \text{ nm} \]

An e⁻ with velocity of 1.9 \times 10^8 \text{ cm/sec:}
\[ \lambda = 3.88 \times 10^{-10} \text{ m} = 0.388 \text{ nm} \]

Schrödinger applied the idea of an e⁻ behaving as a wave to the problem of electrons in atoms.
He developed what is called the WAVE EQUATION. The solution to the wave equation gives a set of mathematical expressions called WAVE FUNCTIONS, \( \Psi \).
Each describes an allowed energy state of an e⁻. Quantization comes naturally from the mathematics.