Chapter 7 - Quantum Model of Atom

Section 2 – Bohr's Atomic Theory

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Bohr's Theory of the Atom

To understand the structure of atom in depth, Neils Bohr, around 1913, set up an experiment where he exposed hydrogen gas to energy. The energy released from the hydrogen was passed through a prism to study the results. A line spectra was seen as a result (*See next slide for this*). What did this line spectrum mean?

First, we need to understand the two kinds of spectra: continuous and line.

Continuous spectrum: White light is continuous spectrum, even when passing through a prism it is continuous as one each color merges into the other.

Line spectrum: Obtained by exciting a gas, passing the energy released through a slit and prism. Only certain radiation at a particular wavelength are observed. This is also called emission spectrum.

On changing the element, Bohr obtained a different line spectrum.

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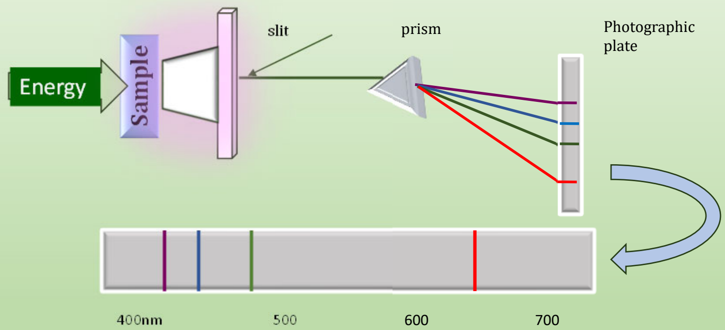
Bohr's Theory of the Atom - Schematic

Neils Bohr set up an experiment as shown below, where he exposed hydrogen gas to energy.



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Bohr's Theory of the Atom - Results

The results of Bohr's experiment could be explained using Planck's theory of quantum along with Einstein's theory of photons.

Here are the key conclusions from Bohr's experiment.

- ***Line spectra***: Emission of light from elements occurring only at specific wavelengths in the visible region.
- Every element gives a unique emission spectrum.
- This line spectrum is generally referred to as "fingerprints" of the element.

What this spectrum indicates is that there are discrete amount of energy absorbed (or emitted) from the atom. This meant that electrons exist in specific location in the atom called "shells". Bohr said the electrons can these "shells" or "orbits", indicated by **n**, can jump to a higher shell by absorbing energy and emit energy when they come back to their ground level shell. This is the energy given in the line spectrum of each element.

See next slide for some examples of the line spectra.

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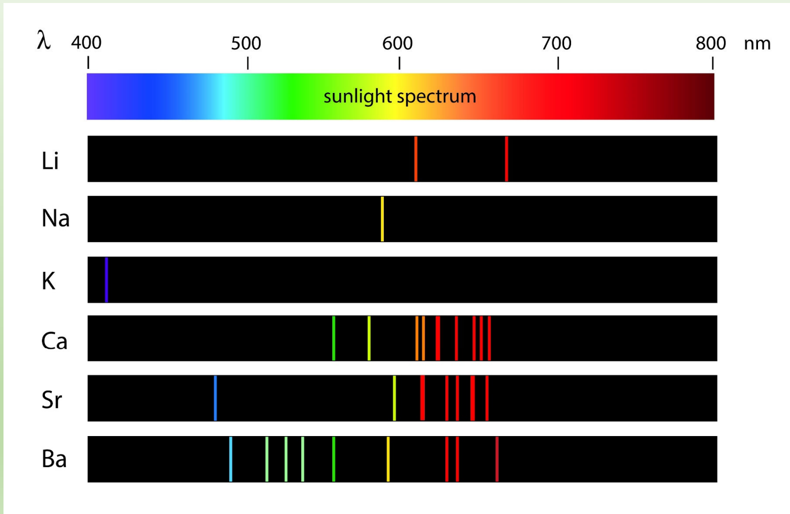
Line Spectrum



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Line Spectrum of Hydrogen

- The *Rydberg equation* is used to calculate the amount of energy required for an electron to transit from one shell to another.

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

- Balmer (initially) and Rydberg (later) developed the equation to calculate all spectral lines in hydrogen.
- Bohr's contribution: showed only valid energies for hydrogen's electron with the following equation

$$E_n = 2.18 \times 10^{-18}\text{J}\left(\frac{1}{n^2}\right)$$

- Bohr's equation can be used to calculate the energy of these transitions within the H atom .

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Electron Transition

Here is what the line spectrum and electron transition mean for any atom.

- Each spectral line corresponds to a specific transition.
- Ground state is the lowest energy state of an atom.
- Excited state is when energy state $n > 1$.
- Energy is required when electrons move from ground state to higher states. ($n_f < n_i$; ΔE will be negative).
- Light/energy is emitted when an electron falls from a higher to a lower state releases energy. ($n_f > n_i$; ΔE will be positive).
- An electron is ejected when $n_f = \infty$.

Solved Problem: Energy calculation in changing shell numbers

What is the wavelength of the light emitted when the electron in a hydrogen atom undergoes a transition from $n = 5$ to $n = 2$?

$n_i = 6$
$n_f = 3$
$R_H = 2.179 \times 10^{-18}$ J

$$\Delta E = -R_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\Delta E = (-2.179 \times 10^{-18} \text{J})\left(\frac{1}{2^2} - \frac{1}{5^2}\right) = -4.576 \times 10^{-19} \text{ J}$$

$|\Delta E| = \frac{hc}{\lambda}$ so $\lambda = \frac{hc}{|\Delta E|}$

$$\lambda = \frac{(6.626 \times 10^{-34} \text{J}\bullet\text{s})(3.00 \times 10^8 \frac{\text{m}}{\text{s}})}{|(-4.576 \times 10^{-19}\text{J})|} = 4.34 \times 10^{-7} \text{ m}$$

Wave Nature of Matter

- Light has properties of both waves and particles (matter). But ***what*** about matter?
- In 1923, **Louis de Broglie**, a French physicist, reasoned that particles (matter) might also have wave properties.
- The wavelength of a particle of mass, *m* (kg), and velocity, *v* (m/s), is given by the de Broglie relation:

$$\lambda = \frac{h}{mv} \quad \textit{where } h = 6.626 \times 10^{-34} \text{J•s}$$

- Building on de Broglie's work, in 1926, **Erwin Schrödinger** devised a theory that could be used to explain the wave properties of electrons in atoms and molecules. Which means that electrons are moving in a wave form and have a particular wavelength and frequency.

Wave Function

In 1927, **Werner Heisenberg** showed how it is impossible to know with absolute precision both the position, *x,* and the momentum, *p,* of a particle such as electron because of its wave nature.

$p = mv$, where is m is mass and v is velocity

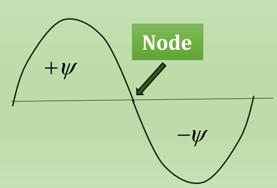
Seeing the equation above, this uncertainty becomes more significant as the mass of the particle becomes smaller.

$$(\Delta x)(\Delta p) \geq \frac{h}{4\pi}$$

- Solving Schrödinger's equation gives us a **wave function**, represented by the Greek letter psi, ψ, which gives information about a particle in a given energy level.

$$\frac{\partial^2 \varphi}{\partial x^2} + \frac{8\,\pi^2 m}{h^2}(E\text{-}V)\,\psi = 0$$

- Psi-squared, ψ^2, gives us the probability of finding the particle in a region of space. One cannot find a particle in the "node".



Key Points

- Niels Bohr’s line spectrum
- Ground state vs excited state
- Heisenberg uncertainty principle
- Schrodinger's wave function