The Nature of the Atom

Hydrogen Spectral Lines:

Lyman Series: \[ \frac{1}{\lambda} = R \left( \frac{1}{1^2} - \frac{1}{n^2} \right) \quad n = 2, 3, 4, \ldots \]

Balmer Series: \[ \frac{1}{\lambda} = R \left( \frac{1}{2^2} - \frac{1}{n^2} \right) \quad n = 3, 4, 5, \ldots \quad R = 1.097 \times 10^7 \text{ m}^{-1} \]

Paschen Series: \[ \frac{1}{\lambda} = R \left( \frac{1}{3^2} - \frac{1}{n^2} \right) \quad n = 4, 5, 6, \ldots \]

The Bohr Model:

Bohr Radius: \( r_n = (5.29 \times 10^{-11} m) \frac{n^2}{Z} \quad n = 1, 2, 3, \ldots \)

Bohr Energy: \( E_n = -(2.18 \times 10^{-18} J) \frac{Z^2}{n^2} \quad n = 1, 2, 3, \ldots \)

Energy of Hydrogen Atom: \( E_n = \frac{-13.6 eV}{n^2} \)

⇒ when an electron jumps from a higher energy state to a lower energy state, it emits a photon of energy:

\[ E_{\text{photon}} = E_i - E_f \quad \text{(photon emission)} \]

\[ hf = \frac{hc}{\lambda} = E_i - E_f \]

⇒ an electron can also absorb a photon and jump from a lower to a higher energy state

\[ E_{\text{photon}} = E_f - E_i \quad \text{(photon absorption)} \]

\[ hf = \frac{hc}{\lambda} = E_f - E_i \]

Quantum Mechanical Picture of Hydrogen Atom:

**Principle Quantum Number:** \( n = 1, 2, 3, \ldots \)

**Orbital Quantum Number:** \( l = 0, 1, 2, \ldots, (n - 1) \)

**Magnetic Quantum Number:** \( m_l = -l, \ldots, -2, -1, 0, +1, +2, \ldots, +l \)

**Spin Quantum Number:** \( m_s = +\frac{1}{2} \) or \( m_s = -\frac{1}{2} \)

Pauli Exclusion Principle: No two electrons in an atom can have the same set of values for quantum numbers \( n, l, m_l, \) and \( m_s. \)