Atomic Structure Overview

Lecture Notes on Bohr's Hydrogen Spectrum

Bohr's Hydrogen Spectrum

Definition:

- The hydrogen spectrum consists of a series of lines that represent the transitions of electrons between energy levels in a hydrogen atom.

Key Concepts:

1. **Emission Spectrum:**
    - When a hydrogen atom transitions from a higher energy level to a lower energy level, it emits a photon of specific energy, resulting in discrete spectral lines.
2. **Absorption Spectrum:**
    - When a hydrogen atom absorbs energy, an electron moves to a higher energy level, producing an absorption spectrum with dark lines.
3. **Series in Hydrogen Spectrum:**
    - **Lyman Series:** Transitions to $n = 1$ (Ultraviolet region).
    - **Balmer Series:** Transitions to $n = 2$ (Visible region).
    - **Paschen Series:** Transitions to $n = 3$ (Infrared region).
    - **Brackett Series:** Transitions to $n = 4$ (Infrared region).
    - **Pfund Series:** Transitions to $n = 5$ (Infrared region).

Postulates of Bohr's Atomic Model

1. **Quantized Orbits:**
    - Electrons orbit the nucleus in specific, fixed orbits with quantized angular momentum.
2. **Stable Orbits:**
    - While in these quantized orbits, electrons do not radiate energy and thus remain stable.
3. **Energy Quantization:**
    - Electrons can only gain or lose energy by jumping from one allowed orbit to another, absorbing or emitting a photon with energy $E = hv$.

4. **Angular Momentum:**
    - The angular momentum of an electron in a permitted orbit is quantized and given by $L = n\hbar$, where $\hbar = \frac{h}{2\pi}$.

Energy of an Electron in an Orbit

Formula:

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

Explanation:

- E_n is the energy of the electron in the n-th orbit.
- The negative sign indicates that the energy is lower (more negative) than that of a free electron at rest (which is zero).

Key Points:

- The energy levels are quantized.
- The energy becomes less negative as n increases, meaning higher orbits have higher (less negative) energy.

Radius of Electron Orbit

Formula:

$$r_n = n^2 \times 0.529\,\text{Å}$$

Explanation:

- r_n is the radius of the n-th orbit.
- The radius increases with the square of n, meaning electrons in higher orbits are farther from the nucleus.

Limitations of Bohr's Model

1. **Applies Only to Hydrogen-like Atoms:**
    - The model accurately describes only single-electron systems (e.g., $\text{H}, \text{He}^+, \text{Li}^{2+}$).
2. **No Explanation for Spectral Line Intensities:**
    - Bohr's model does not explain the relative intensities of spectral lines.
3. **Zeeman Effect:**
    - It cannot explain the splitting of spectral lines in a magnetic field.

4. **Fine Structure:**
    - The model fails to account for the fine structure of spectral lines, arising from relativistic corrections and electron spin.
5. **Electron Transition Probability:**
    - The model does not incorporate the probability of electron transitions between energy levels.
6. **Quantum Mechanics:**
    - Bohr's model was eventually superseded by the more accurate and comprehensive quantum mechanical model, which explains atomic structure and behavior using wave functions and probability distributions.

Summary

- **Bohr's Model**: Proposed quantized orbits for electrons with specific energies and radii, explaining the hydrogen atom's emission spectrum.
- **Energy Levels and Radii**: Provided formulas for calculating the energy and radius of electrons in these orbits.
- **Limitations**: Recognized the model's inability to explain multi-electron atoms, spectral line intensities, Zeeman effect, and fine structure, paving the way for quantum mechanics.