Lecture 2 — Atomic Structure

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- Dalton's Atomic Theory (1803)
  - Key Postulates:
    * All matter is made up of tiny, indivisible particles called atoms
    * Atoms of a given element are identical in size, mass, and properties
    * Atoms of different elements differ in these properties
- Thomson's Plum Pudding Model
  - Cathode Ray Experiment (1897)
    * Observed that cathode rays are streams of negatively charged particles
    * Discovery of the electron, the first subatomic particle
  - Plum Pudding Model
    * Proposed by J.J. Thomson
    * Atoms consist of a positively charged "pudding" with negatively charged electrons embedded within, like plums in a pudding
  - Significance:
    * Challenged Dalton's idea of indivisible atoms
    * Demonstrated that atoms have internal structure
  - Limitations:
    * Could not explain the distribution of charge or atomic structure
- Rutherford's Gold Foil Experiment (1911)
- Bohr's Model (1913)
  - Key Features:
    * Electrons orbit the nucleus in fixed, quantized energy levels

    - Electrons can move between energy levels by absorbing or emitting energy (photons)
    - Orbits correspond to specific allowed energy states, preventing electron collapse into the nucleus
  - Supporting Evidence:
    - Successfully explained the hydrogen emission spectrum
    - Discrete spectral lines correspond to energy transition between levels

- Modern Quantum Mechanical Model (Wave-Particle Duality, de Broglie, Schrödinger)
  - Key Concepts:
    - Electrons exhibit wave-particle duality (de Broglie hypothesis)
    - Electrons exist in orbitals, regions of space with a high probability of finding an electron
    - Atomic behavior described using Schrödinger's equation, which defines the wave function (ψ)
  - Quantum Numbers
    - Describe the unique quantum state of an electron in an atom
    - Define energy, shape, orientation, and spin of electron orbitals
    - Four Numbers:
      1. The Principal Quantum Number (n)
      2. Angular Momentum Quantum Number (l)
      3. Magnetic Quantum Number (m_l)
      4. Spin Quantum Number (m_s)

- Atomic Principles
  - Aufbau Principle:
    - Electrons fill orbitals starting with the lowest energy level first
    - Order of orbital filling: $1s \to 2s \to 2p \to 3s \to 3p \to 4s \to 3d \to 4p$, etc.
    - Visualize the filling sequence with the diagonal rule or energy diagram
  - Pauli Exclusion Principle:
    - No two electrons in an atom can have the same set of all four quantum numbers (n, l, m_l, m_s)
    - Each orbital can hold a maximum of two electrons with opposite spins
  - Hund's Rule:
    - When electrons fill degenerate orbitals (orbitals with the same energy, *e.g.* p,d,f) they maximize unpaired spins before pairing
    - Ensures the lowest-energy arrangement by minimizing electron repulsion

- Significance:
    - Explains electron configurations of elements
    - Influences magnetic and electrical properties (like ferromagnetism)

- Impacts of Atomic Structure
  - Bonding: Determines whether a material is metallic, covalent, or ionic
  - Electron Configuration: Influences conductivity, magnetism, and optical properties
  - Semiconductors: Silicon (Si) covalent bonding and band gap make it ideal for transistors
  - Insulators: Aluminum Oxide (Al_2O_3) strong ionic bonds and high band gap prevent conductivity

- Periodic Table Organization
  - Structure:
    - Rows (Periods):
      - Indicate the principal quantum number (n) of the outermost electron shell
    - Columns (Groups):
      - Elements in the same group have similar valence electron configurations, leading to similar chemical properties