ELECTRONIC CONFIGURATION

Introduction

The electron configuration of an element delineates the specific arrangement of electrons within its atomic orbitals. This configuration adheres to a standardized notation, wherein all atomic subshells containing electrons are organized in a sequential order, with the number of electrons in each subshell denoted as superscripts.

For instance, consider the electron configuration of sodium, which is represented as $1s^2 2s^2 2p^6 3s^1$. In this representation, the numbers and letters denote the principal energy level and the type of subshell (s, p), respectively, while the superscripts indicate the number of electrons within each subshell. Sodium, in this case, has electrons distributed across the 1s, 2s, 2p, and 3s atomic subshells according to this systematic notation.

The distribution of electrons into orbitals follows several fundamental rules, each playing a crucial role in determining the electronic configuration of an atom.

These rules are:

1. Aufbau's Principle

According to Aufbau's principle, electrons fill orbitals with the lowest energy levels before moving to orbitals with higher energy levels. This principle guides the sequential filling of atomic orbitals based on their increasing energy levels, contributing to the systematic arrangement of electrons within an atom.

2. Pauli's Exclusion Principle

Pauli's exclusion principle states that no two electrons within an atom can have the same set of quantum numbers. This means that each electron must have a unique combination of quantum numbers, which include the principal quantum number (n), azimuthal quantum number (l), magnetic quantum number (m), and spin quantum number (s). This principle ensures the distinctiveness of each electron's identity within the atomic structure.

3. Hund's Rule of Maximum Multiplicity

Hund's rule governs the way electrons are distributed among orbitals within a subshell. It dictates that electrons occupy degenerate orbitals (orbitals with the same energy level) in a way that maximizes the total spin, or in other words, the "multiplicity" of the system. Electrons tend to occupy separate degenerate orbitals with parallel spins before pairing up. This rule contributes to the stability and overall energy minimization of the electronic configuration.

In summary, Aufbau's principle establishes the order of orbital filling based on energy levels, Pauli's exclusion principle ensures the uniqueness of electron quantum states, and Hund's rule guides the distribution of electrons within degenerate orbitals to achieve maximum multiplicity. These rules collectively govern the systematic and organized arrangement of electrons in atomic orbitals.