

Multi-Electron Atoms and the Periodic Table (mostly Chapter 1)

I. The Electronic Structure of Atoms in the First Three Periods

i.e. what orbitals do electrons occupy, and why?

- (1) Pauli Exclusion Principle: No more than 2 electrons per orbital
- (2) Aufbau (or building-up) Principle: Put electrons in the most stable (i.e. lowest energy) orbital available (if you are describing the ground state of an atom)
- (3) The energy of an orbital is determined largely by its principal quantum number n
 - n determines the average distance of an orbital from the nucleus (E_p !)
 - n determines the total number of nodes in an orbital (E_K !)
- (4) Paramagnetic atoms (at least one unpaired electron) are attracted to magnets; diamagnetic atoms (no unpaired electrons) are repelled by magnets.
- (5) Core electrons (occupy filled inner shells) play no direct role in chemical bonding; valence electrons (occupy the outermost shell(s)) are involved in chemical bonding.
- (6) For an isolated atom (e.g. in the gas phase), all of the orbitals within a subshell are degenerate (i.e. they have interchangeable orientations and are equal in energy)
e.g. $\{2p_x, 2p_y, 2p_z\}$ or $\{3d_{xy}, 3d_{yz}, 3d_{xz}, 3d_{x^2-y^2}, 3d_{z^2}\}$
- (7) Hund's Rule: Maximize the number of unpaired electrons—that is, avoid putting electrons in the same orbital, if possible

II. The Electronic Structure of Atoms in the Fourth Period and Below

- (8) The energy of an orbital is proportional to $n + l$
- (9) Groups 6 and 11 usually disobey the aufbau principle

III. The Electronic Structure of Ions

- (10) Main group: Add or subtract electrons from the neutral atom's configuration
- (11) Transition metal (Groups 3-12) cations: Always remove electrons from the ns subshell before the $(n - 1)d$ subshell