CHAPTER 1

Atomic Structure - electrons reside around the nucleus (protons and neutrons) in discrete energy levels (shells).

The shells contain orbitals - a mathematical representation of a probability that an electron will exist in that space.

Orbitals are described by quantum mechanics in distinct shapes - s spherical, p dumbell shaped. In this course we will not be dealing with d and f orbitals.

Electron configurations describe how an atom fills its shells and orbitals with electrons. Remember the principles from general chemistry:

Aufbau Principle - the lowest energy orbital fills up first.

Pauli Exclusion Principle - only two electrons can occupy an orbital and they must be of opposite spin.

Hund's Rule - if two or more orbitals of equal energy are available they are filled with one electron of parallel spin until all are half filled. Only then will electrons be paired up. This is important in filling the three p orbitals.

The electron configuration for Carbon is: 1s^2, 2s^2, 2p_x^1, 2p_y^1 Notice that there are four valence electrons and carbon makes four bonds with other atoms.

Bonding - The number of valence electrons determines the number of bonds an atom will make. The atom will try to satisfy the octet rule

eg. C - 4 bonds, H - 1 bond, F - 1 bond, O - 2 bonds, N - 3 bonds

Ionic Bonds - Atoms held together by electrostatic attraction (ions).

Covalent Bonds - Atoms share electrons between them - much stronger than ionic bonds.

Ionization Energy - The energy required to remove an electron from an isolated atom. Ionization energy generally increases (harder to remove electron) as you move the the right and up on the periodic table.

Electron Affinity - The tendency of an isolated atom to gain an electron. Electron affinity generally increases as you move the the right and up on the periodic table.

Na (1s^2 2s^2 2p^6 3s^1) will easily give up an electron to achieve a filled shell, Na^+ (1s^2 2s^2 2p^6). F (1s^2 2s^2 2p^5) will easily gain an electron to achieve a filled state, F^- (1s^2 2s^2 2p^6). NaF will be an ionic compound.

Carbon (1s^2 2s^2 2p^2) would have to gain or lose 4 electrons to achieve a filled shell and is therefore not easily ionized. Carbon forms covalent bonds rather than ionic bonds.

Molecule - A collection of atoms held together by covalent bonds.
Lewis Structures - represent the valence electrons around the atom by dots. Covalent bonds are shown with the two dots shared between the atoms.

Kekulé Structures - each covalent bond is represented by a single line (2 electrons). Unshared or lone pairs of electrons may or may not be shown.

Valence Bond Theory - Covalent bonds are formed by the overlap of atomic orbitals. Each atom maintains its atomic orbital, but shares the pair of electrons. The greater the overlap, the stronger the bond.