Module 7:

1. Know how to do the Lewis dot structures. We will only do these for ionic compounds that contain a group A metal (not transition metals):

a. Binary Ionic Compounds-Those that contain a metal and a nonmetal. You need to know the formula for the compound based on the group that each element is in and the charged ion that it forms and knowing that the sum of the positive and negative ions is equal to 0. You then show the metal without dots but with the charge and the nonmetal with eight dots and with the charge. If more than one of either ion you show parenthesis around the symbol and charge and indicate how many of each ion outside the parenthesis.

b. Covalent Compounds or polyatomic ions-Those that contain only nonmetals and/or metalloids. You need to end up with 8 electrons around each atom (the electrons are shared through bonds). Exceptions are H (only 2 electrons), B (only 6 electrons), and Be (only 4 electrons). One way to come up with formula is by adding up valence electrons for each element plus addtl. electrons corresponding to the charge of a polyatomic ion then distributing them starting two for each bond, then around each atom except for central atom, finally on available spaces on central atom. If at this point the central atom and any other atom does not have 8 electrons (except the exceptions), you will take two electrons from one of the adjacent atoms to the central atom and share them as an additional bond to form a double bond. It may be necessary to do this again with another pair of electrons to form a second double bond (CO2), or a second time with the same atom to form a triple bond (N2, CN-).

c. Resonance structures result when more than one atom can participate in a double bond with the central atom. There can be 0, 2 (SO2) or 3 (NO3-) resonance structures.

2. Electron geometries that result from the number of electron pairs (nonbonded electron pair, single bond, double bond and triple bond each count as one electron pair) around the central atom of the Lewis dot structure:

a. 2 electron pairs-linear
b. 3 electron pairs-trigonal planar or planar triangle
c. 4 electron pairs-tertrahedral

3. Molecular geometries or shapes that result once the number of nonbonded electron pairs around the central atom is taken into account. If 0 nonbonded electron pairs then the moleculear geometry or shape is the same as the electron geometry. The bonding angles will be 180 degrees for linear, 120 degrees for trigonal planar or planar triangle and 109.5 degrees for tetrahedral.

a. Trigonal Planar electron geometry with one nonbonded electron pair is V-shape, angular or bent with bonding angles of less than 120 degrees.(SO2)

b. Tetrahedral electron geometry with one nonbonded electron pair is trigonal pyramid with less than 109.5 degrees bonding angles.(NH3)

c. Tetrahedral electron geometry with two nonbonded electron pairs is V-shape, bent or angular with less than 109.5 degrees bonding angles.(H2O)

4. Electronegativity-Attraction of atoms for electrons.

a. Know that it increases to the right and to the top of the periodic table. The most EN element is F.

b. You must know the following EN's:
F-4.0, O-3.5, N and Cl-3.0, C-2.5, H-2.1

c.The further from each other in the periodic table usually the higher the delta EN and the more polar or more ionic the bond.

d. You need to know the type of bond based on delta EN: less than or equal to .4 is nonpolar, between .5 and 1.9 is polar, more than or equal to 2.0 is ionic.

e. Remember that H is less EN than any halogen.