Chem 115 Lecture 23 [me]
4/28/09

A few notes about class:
-There will be no assignment #13, instead continue working on assignment #12, discussion classes will review for exam 3
-Exam 3 will be next Tuesday (May 5)

Review of physical properties whose trend we have covered:

Atomic radius (*Distance between nucleus at center of atom (protons) and most loosely bound valence electron)*
□ Increases down group (because more shells)
□ Decreases across (L→R) period (because Zeff increases)
□ He is smallest atom

Ionization energy (*Energy (endothermic) required to remove the most loosely bound electron)*
□ Decreases/become less endothermic down group (easier to remove e^- because farther away from + charged nucleus)
□ Increases across period (stronger force holding e^- to core because Zeff increases)
□ Cs has smallest ionization energy, i.e., "wants" to get rid of electron the most (Fr would have smaller IE but is too radioactive to be useful)

Electron affinity *(Energy change resulting from adding one e^-… in most cases it is exothermic (releasing heat energy), but sometimes endothermic (must input energy to force atom to accept the e^-)*
□ Generally becomes less exothermic (smaller negative value) down group (more shells so new e^- not attracted as strongly)
□ Generally becomes more exothermic (larger negative value) across period (because Zeff increases)
□ F has most exothermic electron affinity, i.e., "wants" electron the most

Electron affinity and ionization energy are opposites of each other. Ionization adds and electron, while electron affinity removes an electron

Importance of physical properties:
- Covalent bonding in molecules involves valence electrons of atoms
-Electrons in the bonds are shared (covalent bonding theories in chapter 9)
- By the two atoms that the bond is holding together (electrons localized in *valence bonds*)
- Across the entire molecule (electrons delocalized in *molecular orbitals*)
- Electrons in the bonds are not necessarily shared equally
- The more unequally the electrons are shared, the more ionic character the bond takes on (*bond polarity*)
-Bonds with more ionic character are stronger, due to additional attraction between δ+ pole and δ- pole
- You can predict qualitative comparisons of ionic character of bonds

- The more exothermic an atom's EA is (closer to F), the more easily it attracts the electrons in the bond and becomes stronger δ- pole
- The smaller an atom's IE is (closer to Cs), the more easily it allows electrons in the bond to be attracted away from it and becomes stronger δ+ pole

Stability predictions that can come from Lewis dot structures:

1. Resonance
2. Formal charge is important when resonance structures are not equal
3. Bond order (useful for strengths)
4. Bond Length comparison (comparison between molecules)

Resonance Structures: Lewis structures allow us to see a blending of molecules. Allows comparing two structures, and seeing what they have in common, as well as what is different.

Bond order = # of bonds/ # of locations. To calculate, Find valence number by adding electrons of all elements. Next see which has the least electron infinity (in a row they are on the metal side, on a column they are on the bottom)

Bond Order/Bond length/strength:

- Bond order

 Single bond is bond order 1
 Double bond is bond order 2
 Triple bond is bond order 3
 Fractional bond orders occur when there are resonance structures

- Bond strength

 The greater the bond order, the stronger the bond (the more energy required to break the bond). The strongest is a triple bond.

-Bond length

 The greater the bond order, the shorter the bond length because of less attraction.

Formal Charges:

Used to explain why some are more stable than others. To find a formal charge, find the valence number and subtract from what it would be if it were democratically assigned.

Formal charge when resonance structures aren't equivalent:

 If more than one Lewis structure exists, the most stable structure is the one in which the formal charges make most sense
 Negative formal charges on atoms with large electron affinity
 Positive formal charges on atoms with small ionization energies (small electron affinity)
 The most stable structure of the 3 structures below is the rightmost
because O is more stable with negative formal charge on it because it has largest EA of the 3 elements (O, C and N)

Range of types of bonds:

-Nonpolar covalent (or perfect)- perfectly covalently shared electrons

-Polar covalent- Bonding electrons exhibit greater electron density on one atom in the bond than on the other.

-Ionic- Ions with full charges (unshared electrons) strong attraction.

Electronegativity Model: A scale using numbers 1-4 to assign elements by how electronegative they are. We don't need to memorize this scale, however know that F is the highest, and Cs is the lowest. H is an exception. The larger altitude difference of height on the "electronegativity mountain" (thinking of electronegativity as a 3-D bar graph like in the text book) to climb from one atom to the other, the more polar the bond is (stronger dipole moment)

Predicting Bond Polarity:

-Electronegativity is a measure of the ability of an atom to attract electrons to itself
- Bond polarity depends on the **difference** between the electronegativities of the two elements that are in a bond