## Reading Assignment – Chapter 1

Most of the information was the same as we learned in class, so I understood it pretty well. We did not talk about electrostatic potential maps – I liked the way the colors made it more visual. We also did not use the EN numbers in class, but I prefer the way we did it, just going by the direction in the Periodic Table.

In class we talked about charged ions, but didn't call them "formal charges." We also didn't talk about the way to calculate the charge by looking at the valence electrons minus the number of bonds minus the number of nonbonding electrons. I like the way it was done in the notes better, where we just looked at where the atom had lost or gained electrons. Looking again at the formula I guess it comes out pretty much the same. The sample problems were very helpful, and there was a table that will also be helpful to refer back to.

I didn't understand resonance structures very well, so I read this section extra carefully. I noticed that it followed the learning guide mostly, but there were a few differences. I liked the analogy of the horse and mule making a donkey. Using the word delocalized confused me – I preferred to talk about the electrons "spreading out."

I am pretty good at reading, but still find it a chore. I tried reviewing the learning guide before reading, and that helped. I also tried reading out loud, and that seemed to help me focus. I tried not not go on until I understood what I had read. Next time maybe I will try to find a better time of day to read – late afternoon I am too sleepy!

Notes:

1-6 Electronegativity and Bond Polarity (1 page)

nonpolar bond – electrons shared equally (same element) polar bond – electrons attracted more toward one nucleus (different elements)

C-Cl – bonding electrons attracted to Cl C has partial positive charge Cl has an equal amount of negative charge

polar bond symbolized by an arrow with going towards the more EN atom, plus on other end

bond polarity measured by dipole moment – amount of charge separation multiplied by bond length

lower case delta + or - means a small amount of positive or negative charge

partial charges can be shown on a electrostatic potential map – red areas are electron-rich, blue areas electron-poor

electronegativities are a guide to predict amount and direction of bond polarity organic chemists use Pauling scale atom with higher EN is the negative end

EN increases left to right N, O, halogens more EN than C Na, Li, Mg less EN than C C and H are similar, C-H bonds are considered nonpolar

Problem 1-5 Predict direction of dipole using electronegativities. a) C-Cl b) C-O c) C-N d) C-S e) C-Br  $\rightarrow$   $\rightarrow$   $\rightarrow$   $\rightarrow$ f) N-Cl g) N-O h) N-S i) N-B j) B-Cl  $\rightarrow$   $\rightarrow$   $\leftarrow$   $\leftarrow$   $\rightarrow$ 

1-7 Formal Charges (1 page)

partial charges in polar bonds are real; formal charges keep track of electrons, but may not correspond to real charges – helps figure out where most of the charge is, what the overall charge is on a molecule

to calculate formal charge: compare valence e- of neutral atom to number of bonds and nonbonding electrons

 $CH_4$  – carbon has 4 valence e-; 4 bonds; charge is 0  $H_3O^+$  – oxygen has 6 valence e-; 3 bonds, 2 nonbonding e-; charge is +1  $H_3NBH_3$  – nitrogen has 5 valence e-; four bonds; charge is +1 boron has 3 valence e-; four bonds; charge is -1 [H<sub>2</sub>CNH<sub>2</sub>]<sup>+</sup> – carbon has 4 valence e-; 4 bonds, charge is 0 nitrogen has 5 valence e-; 4 bonds; charge is +1 or carbon has 4 valence e-; 3 bonds; charge is +1 nitrogen has 5 valence e-; 3 bonds, 2 nonbonding e-; charge is 0

in o-chem, there are only a few common ions

Table on Page 13 shows positive, neutral, and negative ions for B, C, N, O, and Cl \*\*\*very useful\*\*\*

1-9 Resonance (3 pages)

some compounds can't be adequately represented by a single Lewis structure

resonance structures are different ways of drawing the same compound, not different compounds; actual molecule is a hybrid of the resonance forms

electrons are delocalized - spread out into more than one position

spreading charge out makes an ion more stable - the ion is said to be resonance-stabilized

resonance is often important in compounds with double bonds

acetic acid is more acidic because its conjugate base is resonance stabilized each O has  $\frac{1}{2}$  a negative charge, each C-O bond has  $\frac{1}{2}$  bonds

resonance structures are separated by double-headed arrows and enclosed in brackets an equilbrium is represented by two arrows in opposite directions

nitromethane has a full positive charge on the N,  $^{1\!/}_{2}$  negative charge on each O,  $1^{1\!/}_{2}$  bonds between N and O

molecules with resonance structures do not change back and forth between the structures; they always have some characteristics of both -a donkey is a hybrid of a horse and a mule, but it always looks like a donkey, not sometimes like a mule and sometimes like a horse

resonance structures can have different energies - if so, the more stable ones are closer representations of the real molecule than less stable ones

the more stable form is the major contributor, less stable the minor contributor

a second resonance structure could be written for formaldehyde in which there is a positive C and negative O - this helps explain the polarity of the bond

the best resonance structures have the lowest energy; try to make the most possible octets and bonds

all resonance structures must be valid Lewis structures

nuclei (atoms) can't be in different places, only electrons

major resonance contributor has the lowest energy; lower energy structures have octets on all atoms, negative charges on electronegative atoms, no charge separation