V25.0109: General Chemistry I (Honors)

Problem set #7: due 11/6

Practice problems from Chapter 5: 11,12,15,17
Practice problems from Chapter 6: 13,16,21

Graded problems

1. A KCl molecule is struck by a UV pulse carrying 1000 kJ/mol of kinetic energy, which causes the molecule to dissociate into K and Cl. It is observed that each atom of the two atoms has approximately the same kinetic energy after dissociation. Assuming perfect ionic bonding, calculate the speed of each atom in cm/s. Take the equilibrium bond length of KCl to be 2.67 Å. **Hint**: You might find some of the graphs in chapter 4 helpful. Feel free to make your best estimates from these graphs.

2. The CaF bonds in the molecule CaF_2 are ionic with a $+2e$ charge on the calcium and $-e$ charge on each of the fluorines. If the first and second ionization energies of calcium are 590 kJ/mol and 1140 kJ/mol, respectively and the electron affinity of fluorine is 328 kJ/mol, calculate the formation energy of the molecule. Take the Ca-F bond lengths to be 2.56 Å, and the distance between the two fluorines to be 4.525 Å.

3. A form of pure molecular carbon, C_{60} (called buckminsterfullerene), was first discovered in 1985 by Kroto, Curl and Smalley. The molecular geometry is roughly that of a soccer ball, i.e., with the carbons located on the vertices of the hexagons and pentagons that compose it. Therefore, each carbon is bonded to three other carbon atoms. Using the rules for drawing Lewis dot structures, determine the total number of single and double bonds in the molecule and how many of each type of bond per carbon atom. You may assume that no triple bonds are formed. **Note**: Do *not* use symmetry considerations or plausibility arguments! *Derive* the result directly from the rules of Lewis structures. You do not actually need to draw the Lewis structure.

4. Hydrocarbons are a class of molecules having the general molecular formula C_nH_m, where each hydrogen is bonded *only* to a single carbon. (A carbon may be bonded to more than one hydrogen, however.)

   a. According to the rules of Lewis structures, how many of the valence electrons of a general hydrocarbon are bonding electrons and how many are in lone pairs? Express you answer in terms of n and m.

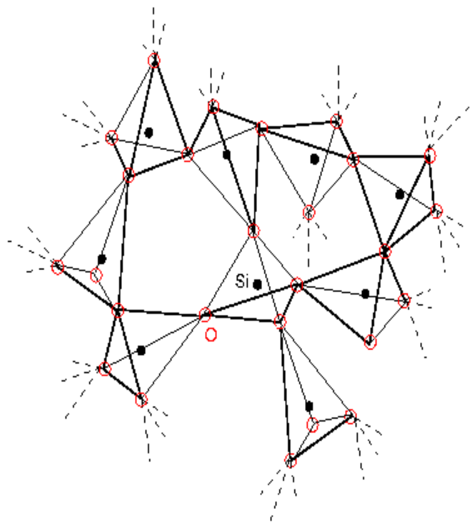
   b. Consider a class of hydrocarbons in which the carbons are bonded to each other in a *linear* arrangement. If we only allow for single bonds between carbons, then using only the information given, determine the number of bonds in such a molecule. Express your answer in terms of n and m. Show how you arrived at your answer.

   c. Based on the information in parts a and b, determine a relation between n and m.

   d. Consider a class of hydrocarbons in which the carbons are bonded to each other in a *cyclic* (ring) arrangement. If we, again, only allow for single bonds between carbons, then using only the information given, determine the number of bonds in such a molecule. Express you answer in terms of n and m. Show how you arrived at your answer.

   e. Based on the information in parts a and d, determine the relation between n and m for this class of hydrocarbons. What is the *empirical* formula for this class of hydrocarbons? How can the special case of $n = 2$ be incorporated?

5. a. Silica, whose empirical formula is SiO_2, is an example of a glassy material that forms an infinite network of well defined structural units. For silica, the basic unit is a tetrahedron with four oxygens at the vertices and a silicon atom at the center. The network is formed by virtue of the fact that any two tetrahedra share a common vertex, *i.e.* share an oxygen atom (see figure below): (Note that the lines in the figure do *not* indicate bonds!).



Lewis structures can be used to predict that such a network will form. Clearly, it is not possible to draw the Lewis structure for an infinite system, which means that the structure must be terminated at some point, resulting in the cutting of bonds. The cut bonds are known as "virtual bonds" and are of two types: i) "donor bonds" are bonds that still contain a shared electron pair but do not have a second atom next to the pair ahd ii) "acceptor bonds" that have neither a shared electron pair nor a second atom. Thus, in an ABA molecule, if one of the AB bonds is cut, donor and acceptor bonds might look as follows:

In any terminated Lewis structure, the number of donor bonds *must* equal the number of acceptor bonds.

On the basis of the information given, use Lewis structures to justify the networked structure in the above figure. Determine the formal charges on both silicon and on oxygen. You may use the next page (which is blank) to show your work.

Hint: Begin by drawing the smallest skeletal structure that includes one tetrahedral unit, but make sure your skeletal structure is consistent with the empirical formula. Determine the Lewis structure for your skeletal structure, making sure to indicate which bonds are virtual bonds.

b. Show that your Lewis structure gives the correct steric number for each tetrahedral unit.

c. Show that an Al can substitute for a Si in the network by acquiring a nonzero formal charge. What is the value of this formal charge? Show how your Lewis structure from part (a) changes when one Al substitutes for one Si?

d. Experiments show that Al can substitute for Si at various points in the network. When this happens, Al atoms are sometimes surrounded by 5 oxygens instead of 4. Explain why this is possible and show how your Lewis structure from part (c) changes. You might find it useful to note that according to the same experiments, Al-O-Al linkages are never observed.