Honors Chemistry Hour\_\_\_\_\_ Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_  
Dr. Wexler  
Lewis Structures of Covalent Compounds Worksheet 1 (23pts)  
Date:

**Background:**   
Drawing Lewis Structures will enable you to determine the structure of an atom both in terms of the types of covalent bonds (single, double, triple, resonant), and the 3D shape of the molecule.

One goal in drawing these structures is to minimize the formal charge on each atom in the molecule. This will be zero for a neutral atom, but non-zero charges often exist for polyatomic ions.   
Also, when different elements are in a molecule the one that is single or the one (except for hydrogen) that is least electronegative will be at the center of the structure.

A single covalent bond contains two shared electrons.   
A double covalent bond contains four shared electrons.  
A triple covalent bond contains six shared electrons.  
In a resonant structure, a double bond alternates with a single bond (alternate structures).

**General Procedure:**Examples:  
A. Single bond: Br2  
Step one Add up the valence electrons. In this case, 7 + 7 = 14  
Step two Draw the skeleton using single bonds Br:Br (same as Br-Br)  
Step three Distribute the rest of the valence electrons using the octet rule (14 – 2 = 12)

Step four Count the formal charges on each atom:   
 valence number – # unpaired electrons – ½(#shared electrons) = 7 – 6 – ½(2) = 0

Since the formal charges are 0, and the octet rule is followed, then this is the best structure.

B. One or more double bonds: SiO2

Step one Add up the valence electrons: 4 + 6 + 6 = 16  
Step two Draw the skeleton using single bonds  


Step three Distribute the unused valence electrons (16 – 2 – 2 = 12)

  
You will note that the silicon atom does not have an octet.   
Also the formal charges on all atoms are non-zero (for O it is 6 – 6 – ½(2) = -1, and for C it is 4 – 0 – ½(4) = +2)  
  
Step four Both problems can be solved by allowing each oxygen to share two of its unbonded electrons with carbon, creating two double bonds:  


The formal charge on O is now 6 – 4 – ½(4) = 0, and the formal charge on C is now 4 – 0 – ½(8) = 0

**Questions:**

1. What is the duet rule? Which atom(s) does this apply to?

2. What is the octet rule?

3. Draw the electron dot (Lewis) structures for each covalent molecule.

F2 O2

H2S CO2

HCl NH3

H2 N2

PCl3 SiH4

BeCl2 BH3

Which molecules above are exceptions to the octet rule?

4. Draw the Lewis Dot Structures for these polyatomic ions. Be sure to adjust the total number of valence electrons according to the charge on the ion. Enclose your final structure in brackets with the charge indicated in the upper right hand corner outside the bracket. If the structure is resonant, indicate that as well.

OH- H3O+

NH4+ CN-

NO2-

ClO4-

PO43-