Lewis structure

A picture that shows the covalent bonds for a molecule.

Does not show accurately the location of all the electrons in the molecule.

Does not show accurately the shape or bonding angles of the molecule.

<u>Example</u>: H_2O

Drawing and understanding Lewis structures are starting points for Chapter 9 (CHEM 102) and for organic chemistry. You need to practice drawing and understanding them until they are easy for you. (Sections 8.2-8.4, 8.7-8.8)

General rules (see handout):

1. Count the valence electrons.
2. Arrange the atoms. Usually a central atom is surrounded by ligands (outer atoms or groups). The outer groups are the more electronegative atoms; the central atom is usually the least electronegative atom because it will be forced to share electrons with the outer groups. H is never a central atom.
3. Distribute bonding pairs between atoms (for each bond, 2 electrons are shared by the bonded atoms). *Note that C prefers 4 bonds, N prefers 3 bonds, O prefers 2 bonds, and halides (F, Cl, etc.) and H prefer 1 bond*.
4. Give as many of the leftover electrons as possible to the outer atoms, but remember the Octet Rule: For 2^{nd} row elements, only 4 pairs (8 electrons) can be touching each atom. Other elements are not required to follow the Octet Rule. *For 1^{st} row elements, only 1 pair*. Why do we give the leftover electrons to the outer atoms before giving them to the central atom?

5. Give all remaining electrons to the central atom. Note the possible exceptions to the Octet Rule.

6. Modify the structure using the Octet Rule and the formal charges to help.
   - Use the Octet Rule to make multiple bonds by sliding electron pairs from outer atoms to central atom.
   - Reduce formal charges to decide when to make multiple bonds.
   - Note the exceptions to the Octet Rule.

Examples:

CH_4

H_2O

<u>Example</u>:

CO_2

<u>Practice</u>: Draw Lewis structures for NH_3, H_3O^+, and SF_6.

Which bonds are polar?
Which molecules obey the Octet Rule?

Resonance

Resonance structures: multiple Lewis structures that collectively show the true structure, also called resonance hybrids, contributors, etc.

Example: CO_3^{2-}

<u>Practice</u>: Draw the Lewis structure for NO_2^-.

Calculating Formal Charge

Formal charge = starting e^- – current e^-
= valence e^- – Lewis structure e^-

- Can help in choosing the preferred structure
- Assumes bonding electrons are shared equally

<u>Example</u>: NCO^-

Practice: Draw the Lewis structure for carbon monoxide (CO), and calculate the formal charge for each atom. What is unusual about the formal charges in this molecule, and why may they be unrealistic? What does this say about the general rules for Lewis structures?