

## Work Session 10: Lewis Structures

Consult the textbook for the rules used to draw Lewis structures.

Here is a brief summary:

- Count the total number of valence electrons in the structure. If the structure is an ion, take the charge on the ion into account.  
(Example:  $\text{NO}_3^-$  has 24 electrons, five from N, six from each O, and one from the negative charge.)
- Draw a skeletal structure for the molecule with all single bonds. The first atom listed, usually the less electronegative of the atoms, often is the central atom in the structure. H and F atoms are *never* central atoms.
- Counting each bond as two electrons, subtract the number of electrons used from the total number available. Place the remaining electrons on the atoms in the structure as non-bonding pairs, so that each atom has a total of 4 pairs of electrons around it. H atoms only require one pair.
- If there are too many electrons, place the extra electrons in pairs around the central atom. Put them in diagonal positions from the bonds already there.
- If there are too few electrons, then for each two electrons the molecule is short, place a non-bonding pair into a bonding position to form a multiple bond.
- If there is more than one position possible for the multiple bond, or if one triple bond could be used in place of two double bonds, then the molecule has resonance structures. Draw each one.
- An atom in a molecule has a *formal charge* if the number of electrons around the atom in the molecule differs from the number of electrons around the atom as an isolated atom.

**Formal Charge** = (# of Valence electrons) - ( $\frac{1}{2}$  # of bonding electrons) - (# of non-bonding electrons).

The smaller the formal charge distribution, the better the structure. A more negative formal charge should reside on the more electronegative atom. In polyatomic ions, the formal charges of the atoms will add up to give the charge on the ion. Note: formal charges are different than oxidation numbers.

**A simple method:** each bonding pair belongs half to each atom bonded. Each non-bonding pair belongs only to the atom it's on. Add up the electrons the atom possesses in the molecule. Compare this number to the number of valence electrons in the unbonded atom. Each extra electron in the molecule gives a negative charge. Each electron less in the molecule gives a positive charge.

### Exceptions to the Octet Rule:

- Hydrogen. Only one pair of electrons.
- Beryllium: Only two pairs of electrons.
- Boron: only three pairs of electrons.
- In rule number 4 above, the atoms that have extra pairs of electrons around them must be in the third row or lower on the periodic table. The possibility of d orbitals in the electronic configurations for these atoms gives them the capability to hold more than eight electrons.
- If there is an odd number of electrons in the structure, there will be unpaired electrons. Such molecules are called *free radicals*.

### EXAMPLES

Molecule	# of Valence Electrons	Skeleton	# of Remaining Electrons	Lewis structure (show any formal charges)	Resonance Structures
$\text{H}_2\text{O}$	8	H—O—H	4		no
$\text{PF}_3$	26		20		no
$\text{SF}_4$	34		26		no
$\text{NO}_3^-$	24		18		

Molecule	# of Valence Electrons	Skeleton	# of Remaining Electrons	Lewis structure (show any formal charges)	Resonance Structures
CF <sub>4</sub>					
NH <sub>3</sub>					
Cl <sub>2</sub> O					
PO <sub>4</sub> <sup>3-</sup>					
NH <sub>4</sub> <sup>+</sup>					
C <sub>2</sub> H <sub>6</sub>					
C <sub>2</sub> H <sub>4</sub>					
C <sub>2</sub> H <sub>2</sub>					
CO					

Molecule	# of Valence Electrons	Skeleton	# of Remaining Electrons	Lewis structure (show any formal charges)	Resonance Structures
XeF <sub>4</sub>					
BrO <sub>3</sub> <sup>-</sup>					
ClO <sub>4</sub> <sup>-</sup>					
I <sub>3</sub> <sup>-</sup>					
NO <sub>2</sub>					
CO <sub>3</sub> <sup>2-</sup>					
H <sub>2</sub> O <sub>2</sub>					
CH <sub>3</sub> OH					
CH <sub>3</sub> NH <sub>2</sub>					

**QUESTIONS**

- $F_2$  has a bond length of 128 pm, and a bond energy of  $154 \text{ kJ}\cdot\text{mol}^{-1}$ .  
 $O_2$  has a bond length of 121 pm, and a bond energy of  $498 \text{ kJ}\cdot\text{mol}^{-1}$ .  
 $N_2$  has a bond length of 110 pm, and a bond energy of  $945 \text{ kJ}\cdot\text{mol}^{-1}$ .

Draw the Lewis structures for each of these molecules. Bond order is the number of bonds between two atoms. Do the bond orders of these three molecules correspond to the bond length and bond energy? Explain.

- $PF_5$  and  $AsF_5$  exist, but  $NF_5$  does not exist. Explain why.

- The molecule  $N_2O$  has six possible Lewis structures. Three have the N-N-O bond skeleton, and three have the N-O-N bond skeleton.

- Draw the six structures.

- The best Lewis structures have the smallest formal charge distributions. Redraw these below.

- Oxygen is more electronegative than nitrogen, so it is more likely to have a negative formal charge than nitrogen. Redraw the best Lewis structure from your answer in part b), given this guideline.

- Resonance structures have atoms in place, with electrons shifting positions. Isomers have atoms shifting positions. Use circles and arrows or other appropriate markings on your answer in part a) to categorize the six structures you have drawn into isomer and resonance structure categories