Mrs. Nielsen

Lewis Structure Rules for AP Chemistry

1. Determine the total number of valence electrons in the bonding atoms.
   *If the species is an ion, the charge of the ion must be considered. For anions, add electrons equal to the charge and for cations, subtract electrons equal to the charge.

2. Arrange the atoms to form a molecule. If C is present, it is the central atom. Otherwise, the least electronegative atom is (usually) central. H is never central.
   * When in doubt, make your molecule symmetrical unless otherwise noted in the formula

3. Connect the atoms with single bonds. *Remember each bond is a pair of electrons.*

4. Satisfy the octet rule for each atom by adding lone pairs.
   * Exceptions: H wants 2 electrons
     Be wants 4 electrons
     "Boron the Moron" wants 6 electrons
   * Note: Some atoms will fill with an expanded octet. This generally occurs with P, S and Cl and the atoms below these in the periodic table. Kr and Xe will also bond to form expanded octets. This WILL NOT be tested on the AP Exam.

5. Count the number of electrons drawn. The total number of electrons in your drawing must be equal to the total number of valence electrons determined in step # 1.
   - If there are too many electrons in your drawing: Remove lone pairs to form double or triple bonds as necessary.
   - If there are too few electrons in your drawing: Add lone pairs of electrons to the central atom to represent expanded octets (these will be hybridizations with d orbitals)

6. Based on the number of lone pairs and atoms bonded to the **central atom**, determine the geometric shape of your molecule. Draw the molecule in the correct geometric shape, including wedges and dashes to represent 3-D molecules.

7. Draw lines between bonded atoms to represent the number of bonds between them. Include lone pairs for all atoms in your final Lewis structure.
   * Ion structures should be placed in square brackets and the charge of the ion placed on the outside upper right of the bracket.

Now that your Lewis structure is complete, you should be able to:

- Indicate the bond polarities (non-polar covalent or polar covalent)
- Determine the molecule is polar or non-polar and draw the dipole moment for a polar molecule
- Determine the steric number and hybridization around each central atom
- Indicate the total number of sigma and pi bonds in the molecule
- Draw resonance structures, if they exist
- Determine which intermolecular forces are present (LDF, Dipole-dipole, Hydrogen bonds)

Helpful Hints: If possible, follow these guidelines

- Carbon always has 4 bonds, as do most elements in Group 14
- Group 15 elements usually have 3 bonds
- Group 16 elements usually have 2 bonds
- Halogens (Group 17) have 1 bond
- Hydrogen has 1 bond only

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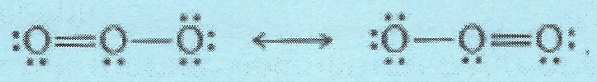
Resonance Structures and Formal Charge

Resonance Structures:

The concept of resonance is when two or more possible Lewis structures can be drawn for the same substance. The actual structure is an average of the possible structures. The different structures are drawn with a double-headed arrow between them to indicate this.



Formal Charge:

The formal charge of an atom in a Lewis structure is the charge it would have if all the bonding electrons were shared between the bonded atoms. Formal charge can be used to distinguish between different resonance structures and help determine which is the best model. To calculate formal charge of an atom, use the following equation:

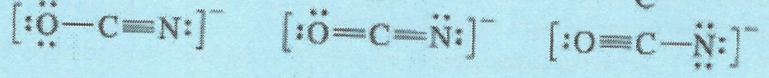
Formal Charge = # of valence electrons – (# of lone pairs + ½ # of bonding electrons)

Comparing the formal charges in competing structures, four rules apply:

1. The sum of all formal charges in a neutral molecule is zero
2. The sum of all formal charges in an ion must equal the charge of the ion
3. Small (or zero) formal charges on atoms are better than large ones
4. If formal charges cannot be avoided, the most negative charge would reside on the most electronegative atom.

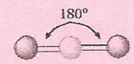
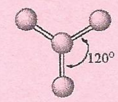
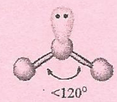
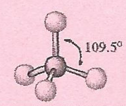
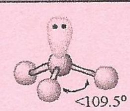
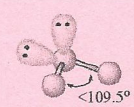
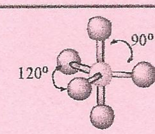
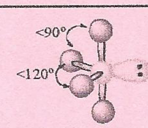
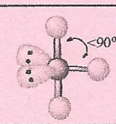
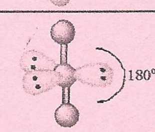
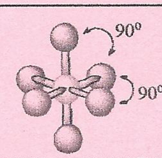
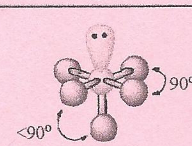
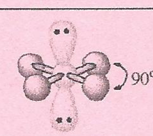
Example: $[OCN]^-$ has three resonance structures. The calculation of formal charge is shown below.

A B C



	A			B			C		
	$[:\ddot{O}-C\equiv N:]^-$			$[:\ddot{O}=C=\ddot{N}:]^-$			$[:O\equiv C-\ddot{N}:]^-$		
Number of valence e^-	6	4	5	6	4	5	6	4	5
−number of nonbonding e^-	−6	−0	−2	−4	−0	−4	−2	−0	−6
$-\frac{1}{2}$(number of bond e^-)	−1	−4	−3	−2	−4	−2	−3	−4	−1
Formal charge	**−1**	**0**	**0**	**0**	**0**	**−1**	**+1**	**0**	**−2**

Molecular Geometry Chart

# of Electron Groups	Number of Lone Pairs	Electron Pair Arrangement	Molecular Geometry	Approximate Bond Angles
2	0	linear		180°
3	0	trigonal planar		120°
	1	bent		<120°
4	0	tetrahedral		109.5°
	1	trigonal pyramid		<109.5° (~107°)
	2	bent		<109.5°(~105°)
5	0	trigonal bipyramidal		90°, 120°
	1	see-saw		<90°, <120°
	2	T-structure		<90°
	3	linear		180°
6	0	octahehral		90°, 90°
	1	square pyramidal		90°, <90°
	2	square planar		90°