Practice Worksheet for Lewis Structures (Mahaffy Ch. 10.1 – 10.5)

1. Main concepts - Lewis Structures
    a. Connectivity
    b. Bonds & Lone pairs
    c. Electron Geometry & Molecular Shape
    d. Resonance Structures
    Formal Charge (**FC**) is a charge an atom in the molecule or ion would have if all bonding electrons were shared equally between the bonded atoms.

    $$\textbf{FC= \# of valence } e^{-} - \frac{1}{2}\textbf{\# of shared } e^{-} - \textbf{\#nonbonding } e^{-})$$

    e. Assigning Oxidation number using Lewis Structures
        i. Oxidation number is **0** for atoms in an element.
        ii. The sum of all oxidation numbers in a molecule or ion must add up to the total charge.
        iii. In compounds, alkalis (group 1) have oxidation number **+1**; alkaline earths (group 2) have oxidation number **+2**.
        iv. In compounds, fluorine (F) always has oxidation number **-1**. Other halogens (Cl ext.) have oxidation number **-1**, *except when bonded to elements that are more electronegative* such as fluorine or oxygen, where they can have positive oxidation numbers.
        v. In compounds, hydrogen has oxidation number **+1** (*when bonded to elements with higher electronegativity*). When bonded to metals hydrogen has oxidation number **-1** *(when bonded to elements with lower elecronegativity)*.
        vi. In compounds, oxygen has oxidation number **-2**, except when bonded to fluorine, where fluorine's rule takes precedence (in OF_2 oxygen has oxidation number **+2**), and if there are O-O (peroxide) bonds. (in H_2O_2 or Na_2O_2 oxygen has oxidation number **-1).**

2. **Useful Ideas to Consider When Drawing Lewis Structures**
    a. Arrange atoms so that the central atom has the lowest electronegativity.
    b. Determine the number of valence electrons.
    c. Place one electron pair between bonded atoms to form a single bond.
    d. Use remaining pairs as lone pairs around terminal atoms (except H) so each atom is surrounded by 8 electrons
    e. If the central atom has fewer than 8 electrons, move one or more of the lone pairs on a terminal atom into a position intermediate between the center and terminal position to form a multiple bond.
    f. The octet rule is a useful guideline when drawing Lewis structures
    g. Carbon forms four bonds (four single bonds; two single bonds and one double bond; or one single bond and one triple bond.) In uncharged species, nitrogen forms three bonds and oxygen two. Hydrogen typically forms only one bond to another atom.
    h. When multiple bonds are formed, both of the atoms involved are usually one of the following: C, N, O, and S. Oxygen has the ability to form multiple bonds with a variety of elements. Carbon forms many compounds having multiple bonds to another carbon, or to nitrogen or oxygen.
    i. Nonmetals may form single, double, and triple bonds but never quadruple bonds.
    j. Always account for single bonds and lone pairs before forming multiple bonds.
    k. Be alert for the possibility that the molecule or ion you are working on is isoelectronic with a species you have seen before.

<u>Example (for CO_2)</u>

V$_{\textbf{alence}}$ =4(C) +2*6(O) =16 e^-(total number of valence electrons ***you have***)

O$_{\textbf{ctet}}$ = 8(C) +2*8(O) =24 e^- (total number of electron ***needed*** for each atom to be satisfied by an octet rule (8 e^- each))

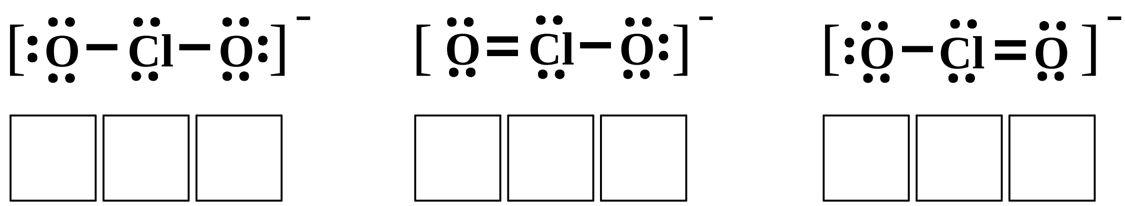
Exceptions: H only needs 2 e^-
Be sometimes only contribute 4 e^-
B usually only contribute 6 e^-

B$_{\textbf{onds}}$ (total number of bonds) = $\frac{\text{O(need)} - \text{V(have)}}{2} = \frac{24-16}{2}$ = 4 bonds

L$_{\textbf{one pairs}}$ **(**total number of nonbonding electron pairs) = $\frac{\text{V} - \text{B}\cdot 2}{2} = \frac{16-4\cdot 2}{2}$ = 4 lone pairs or nonbonding electron pairs (or 8 nonbonding electrons)

1. Choose most stable resonance structure based on Formal Charge for SCN^- and N_2O

2. Consider the following different Lewis structure for **$OClO^-$**



a. Use the boxes to indicate the **formal charge** for each atom.

b. Circle the least stable Lewis structure.

3. Assign **oxidation numbers** to each of the atoms in the following molecules:

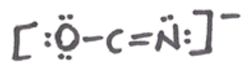
a. **PO_4^{3-}** **P_____** **O_____**

b. NaH **Na____** **H_____**

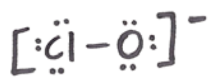
c. C_2H_6 **C_____** **H_____**

4. In which of the following species is it necessary to employ an expanded valence shell to represent the Lewis structure: PO_4^{3-}, PI_3, ICl_3, $OSCl_2$, SF_4, ClO_4^-? Explain your choices.

5. Only one of the following Lewis structures is correct. Select that one and indicate the errors in the others.

a) cyanate ion 

b) carbide ion 

c) hypochlorite ion 

d) nitrogen (II) oxide 

6. One of the following ions has a trigonal planar shape: SO_3^{2-}, PO_4^{3-}, PF_6^-, CO_3^{2-}. Which ion is it? Explain.

7. Use VSEPR theory to predict the geometric shapes of the following molecules and ions: (a) N_2 (b) HCN (c) NH_4^+ (d) NO_3^- (e) NSF.

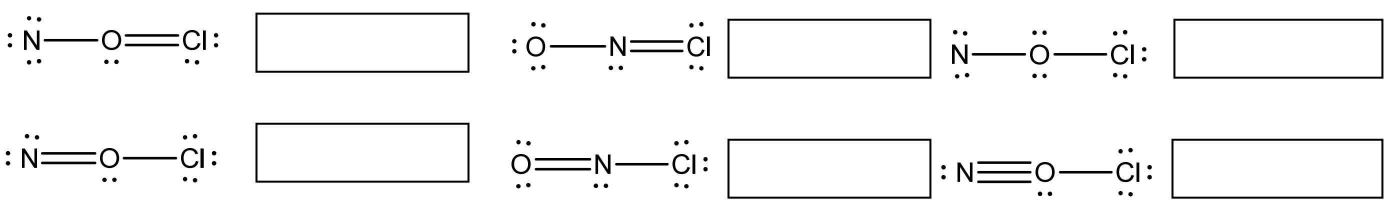
8. Give the Lewis dot structures, electron pair geometry, and molecular shape of the following molecules. For each atom on each molecule, indicate the formal charge and oxidation number.

a. $SeCl_4$
b. I_3^-
c. $PSCl_3$
d. IF_4^-
e. PH_2^-
f. TeF_4^{2-}
g. N_3^-

h. $SeOCl_4$
i. PH_4^+
j. NO^-
k. ICl_2^-
l. PCl_3
m. PF_5
n. ICl_3
o. IF_5
p. $SnCl_2$
q. NO_3^-
r. BrF_3
s. XeF_2
t. BF_3

9. Below are 6 possible Lewis structure. Not all the structures below are reasonable Lewis structures. In the box to the right of each Lewis structure, write the letter(s) corresponding to all the following that apply.

(a) Atom(s) with less than an octet (b) Non-zero formal charge(s) (c) Disallowed Expanded octet



molecule	$\#V_{e-}$	Lewis Structure (with Formal Charge on each atom)	Steric #	e^--geometry	Shape	Dipole?
SO_2 or NO_2^-						
CCl_4 CH_2Cl_2						
ClO_3^-						
NH_2^-						
CH_2O						

PF_5						
H_3PO_4						
C_2N_2						
N_3^-						
O_3						
NOCl						
NH_4^+						

Valence Electron Pairs ***Steric Number =*** ***#of atoms + # of lone pairs attached to the center atom***	**Electron Pair Geometry**	**Number of Bond Pairs** ***(Number of atoms bonded to the center atom)***	**Number of Lone Pairs**	**Molecular Shape**	**Examples**
2	Linear(<180°)	2	0	Linear	$BeCl_2$
3	Trigonal Planar (<120°)	3	0	Trigonal Planar	BF_3
3	Trigonal Planar	2	1	Bent	SO_2
4	Tetrahedral (<109.5°)	4	0	Tetrahedral	CH_4, BF_4^-
4	Tetrahedral	3	1	Trigonal Pyramidal	NH_3, PF_3
4	Tetrahedral	2	2	Bent	H_2O, SCl_2
5	Trigonal Bipyramidal	5	0	Trigonal Bipyramidal	PF_5
5	Trigonal Bipyramidal	4	1	Seesaw	SF_4
5	Trigonal Bipyramidal	3	2	T-shaped	ClF_3
5	Trigonal Bipyramidal	2	3	Linear	XeF_2
6	Octahedral	6	0	Octahedral	SF_6
6	Octahedral	5	1	Square Pyramidal	ClF_5
6	Octahedral	4	2	Square Planar	XeF_4