1. How many valence e- does each of the following have? How many bonds does each of the following tend to form in compounds? How many lone pairs does each of the following have?

<table>
<thead>
<tr>
<th></th>
<th># valence e-</th>
<th># bonds formed (a.k.a. valence)</th>
<th># lone pairs</th>
<th>Sum valence e- and # bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>H</td>
<td>1</td>
<td>0</td>
<td>2</td>
</tr>
<tr>
<td>b.</td>
<td>O</td>
<td>6</td>
<td>2</td>
<td>8</td>
</tr>
<tr>
<td>c.</td>
<td>N</td>
<td>5</td>
<td>3</td>
<td>8</td>
</tr>
<tr>
<td>d.</td>
<td>C</td>
<td>4</td>
<td>4</td>
<td>8</td>
</tr>
</tbody>
</table>

2. Draw the Lewis structures for each of the following
   a. \( \text{N}_2\text{H}_4 \)
      \[
      \begin{array}{c}
      \text{H}_3\text{N} \quad \text{NH}_2 \\
      \end{array}
      \]
      hydrazine
   b. \( \text{CO}_2 \)
      \[
      \begin{array}{c}
      \text{O} \quad \text{C} \\
      \end{array}
      \]
      carbon dioxide
   c. \( \text{HNO}_3 \)
      \[
      \begin{array}{c}
      \text{O} \quad \text{N} \\
      \end{array}
      \]
      nitric acid
   d. \( \text{CH}_3\text{N}_2^+ \)
      \[
      \begin{array}{c}
      \text{N} \quad \text{N} \\
      \end{array}
      \]
      Methanediazonium
   e. \( \text{H}_2\text{NO}^- \)
      \[
      \begin{array}{c}
      \text{H} \quad \text{N} \quad \text{O} \\
      \end{array}
      \]
      or hydroxylamine ion

3. Calculate the formal charge on each atom in each of the structures in #5.
   a. \( \text{N}_2\text{H}_4 \)
      All 0
   b. \( \text{CO}_2 \)
      All 0
   c. \( \text{HNO}_3 \)
      0 +1 2 w/0 and 1 w/-1
   d. \( \text{CH}_3\text{N}_2^+ \)
      0 0 1 w/0 and 1 w/+1
4. Give the hybridization of the central atom in each of the following species as well as the bond arrangement/geometry (linear, trigonal planar, tetrahedral, bent, etc).

   a. \( \text{HNO}_3 \)  
      \[ \begin{array}{c}
         \text{O} \\
         \text{N} \\
         \text{O}
      \end{array} \]
      nitric acid, \( sp^2 \), trigonal planar

   b. \( ^+\text{CH}_3 \)  
      \[ \begin{array}{c}
         \text{H} \\
         \text{H} \\
         \text{H}
      \end{array} \]
      \( sp^2 \), trigonal planar (tetrahedral)

   c. \( \text{H}_2\text{NO}^- \)  
      \[ \begin{array}{c}
         \text{N} \\
         \text{O} \\
         \text{O}
      \end{array} \]
      \( sp^2 \), trigonal planar

   d. \( ^-\text{CH}_3 \)  
      \[ \begin{array}{c}
         \text{H} \\
         \text{H} \\
         \text{C}
      \end{array} \]
      \( sp^3 \), trigonal pyramidal (tetrahedral)

   e. \( \text{CO}_2 \)  
      \[ \begin{array}{c}
         \text{O} \\
         \text{C} \\
         \text{O}
      \end{array} \]
      carbon dioxide, \( sp \), linear

5. Assign hybridization to the indicated atoms, give the molecular geometry and predict the bond angles.

   \[ \begin{array}{c}
      \text{C} \\
      \text{N} \\
      \text{C} \\
      \text{C} \\
      \text{O}
   \end{array} \]
   \[ \begin{array}{c}
      \text{H} \\
      \text{H} \\
      \text{H} \\
      \text{H}
   \end{array} \]
   \[ \begin{array}{c}
      \text{H} \\
      \text{H} \\
      \text{H} \\
      \text{H}
   \end{array} \]

   \[ \begin{array}{c}
      109.5 \quad 107 \quad 120 \quad 180 \quad 104.5
   \end{array} \]

6. Indicate whether each of the following molecules is polar or nonpolar.
   a. \( \text{SO}_3 \)
Lewis Structures, Geometry and Hybridization - KEY

b. \( \text{SO}_2 \) or sulfur dioxide, trigonal planar, so \text{nonpolar}.

c. \( \text{H}_2\text{CO} \) or formaldehyde, trigonal planar, so \text{polar}.

d. \( \text{BCl}_3 \) or boron trichloride, trigonal planar, so \text{nonpolar}.

e. \( \text{CH}_2\text{Cl}_2 \) or dichloromethane, tetrahedral, so \text{polar}.

Additional information:

**Steps for drawing a Lewis structure of compounds:** (+ pg 17 text)

1. Sum the valence electrons from all the atoms
   *If charged…
   a. add 1 electron for every negative charge
   b. subtract 1 electron for every positive charge
2. Optional: Divide the total number of electrons by two to give the number of electron pairs.
3. # of bonds made by each atom usually = 8 - # valence e- (See question #1)
4. If there are three or more atoms, choose a central atom
   a. central atoms are usually written first in the chemical formula
      i. **Carbon** – if there is C, put it in the middle
      ii. **Nitrogen or Phosphorous** – if there no C, put N or P in the middle
      iii. **Oxygen or Sulfur** – if there is no N, P or C, put O or S in the middle (if both are present, S will be the central atom)
      iv. **Chlorine, Bromine, Iodine** – if there is no N, P, C or S, put Cl, Br or I in the center (O can be present)
5. Draw out the atoms, connecting each with a single bond (single line). [Avoid O-O bonds, except in peroxides and peroxyacids, as they are very weak.]
6. Use double (=) (even triple (≡)) bonds to satisfy the octet of the central atom
7. Place remaining electrons as lone pairs (two dots) so as to fill up the octets of the outer atoms first.
   a. start with the more electronegative atoms

EX. Draw the Lewis structure for water

Chemical formula = H₂O
Sum of valence electrons = 2(1) + 6 = 8 valence e⁻/2 = 4 e⁻ prs
Central atom = O (there is only one of them and no N, P or C)

EX. Draw Lewis structures for the following molecules:

1. SiH₄ 4(1) + 4 = 8/2 = 4 e⁻ prs

2. PO₄³⁻ 5 + 4(6) + 3 = 32/2 = 16 e⁻ prs phosphate

3. H₂S 2(1) + 6 = 8/2 = 4 e⁻ prs

4. ClO₃⁻ 7 + 3(6) + 1 = 26/2 = 13 e⁻ prs chlorate ion

5. NH₄⁺ 5 + 4(1) - 1 = 8/2 = 4 e⁻ prs

6. BH₃ 3 + 3(1) = 6/2 = 3 e⁻ prs

7. C₂H₄ (each C) 2(4) + 4(1) = 12/2 = 6 e⁻ prs ethene

Which is the “best” Lewis Structure?
   The one with the lowest formal charge(s) (and negative formal charge on more EN atoms). Careful though b/c having more bonds adds stability.

Assigning formal charge
FC = # valence e⁻ - (# lone pair e⁻ + ½ # bonding e⁻)
There are also some general patterns which would be beneficial to learn to recognize:

<table>
<thead>
<tr>
<th>“central atom”</th>
<th>C</th>
<th>N</th>
<th>O</th>
<th>F</th>
</tr>
</thead>
<tbody>
<tr>
<td>FC on “central atom” ↓</td>
<td>C</td>
<td>H</td>
<td>N</td>
<td>O</td>
</tr>
<tr>
<td>0</td>
<td>C making 4 bonds w/ NO lone pairs</td>
<td>N making 3 bonds w/ 1 lone pair</td>
<td>O making 2 bonds w/ 2 lone pairs</td>
<td>F making 1 bond w/ 3 lone pairs</td>
</tr>
<tr>
<td>+</td>
<td>C making 5 bonds w/NO lone pairs</td>
<td>N making 4 bonds w/ NO lone pair</td>
<td>O making 3 bonds w/ 1 lone pair</td>
<td>F making 2 bonds w/ 2 lone pairs</td>
</tr>
<tr>
<td>-</td>
<td>C making 3 bonds w/ 1 lone pair</td>
<td>N making 2 bonds w/ 2 lone pairs</td>
<td>O making 1 bond w/ 3 lone pairs</td>
<td>F making 0 bonds w/ 4 lone pairs</td>
</tr>
<tr>
<td>+</td>
<td>C making 3 bonds w/NO lone pairs</td>
<td>N making 2 bonds w/ 1 lone pair</td>
<td>O making 2 bonds w/ 1 lone pair</td>
<td>F making 0 bonds w/ 3 lone pairs</td>
</tr>
<tr>
<td>0</td>
<td>C making 2 bonds w/ 1 lone pair</td>
<td>N making 1 bond w/ 2 lone pairs</td>
<td>O making 0 bonds w/ 3 lone pairs</td>
<td><img src="https://example.com/image.png" alt="Image" /></td>
</tr>
</tbody>
</table>

**CHEM 109A**

**Lewis Structures, Geometry and Hybridization - KEY**

- CH₄, methane
- NH₃, ammonia
- H₂O, water
- C₂H₄O, formaldehyde
- C₂H₂, acetylene
- N₂H₂, diazene or diimine
- FHF, hydrofluoric acid
- C₃H₆O₂, 2-propanone
- CH₅+, N⁺, NH₄⁺, O⁺, F⁺, H⁺, H₂O⁺, [F⁺]
- CH₃-, N⁻, NH₂⁻, O⁻, F⁻, C⁻, CH₂⁻, [N⁻]

Page 5 of 8
Determining Molecular Shape/Geometry:

Steps:
1. Draw the Lewis structure for the molecule
2. Count the electron groups around the atom of interest
   *each of the following counts as one electron group
   
   Lone pair, single bond, double bond, triple bond
3. Find the geometry in the Molecular Geometry Table

Molecular Geometry Table:

<table>
<thead>
<tr>
<th># e-groups</th>
<th># bonding groups</th>
<th># lone pairs</th>
<th>Hybridization</th>
<th>e- geometry</th>
<th>Molecular geometry</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1</td>
<td>0</td>
<td>N/A</td>
<td>Linear</td>
<td>Linear</td>
<td>H₂</td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>0</td>
<td>sp</td>
<td>Linear</td>
<td>Linear</td>
<td>CO₂</td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>0</td>
<td>sp²</td>
<td>Trigonal Planar</td>
<td>SO₃</td>
<td>SO₃</td>
</tr>
<tr>
<td>2</td>
<td>1</td>
<td>1</td>
<td>sp²</td>
<td>Bent</td>
<td>NO₂⁻</td>
<td>NO₂⁻</td>
</tr>
<tr>
<td>4</td>
<td>0</td>
<td>0</td>
<td>sp³</td>
<td>Tetrahedral</td>
<td>CH₄</td>
<td>CH₄</td>
</tr>
<tr>
<td>3</td>
<td>1</td>
<td>0</td>
<td>sp³</td>
<td>Trigonal Pyramidal</td>
<td>NH₃</td>
<td>NH₃</td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>0</td>
<td>sp³</td>
<td>Bent/angular</td>
<td>H₂O</td>
<td>H₂O</td>
</tr>
<tr>
<td>1</td>
<td>3</td>
<td>0</td>
<td>sp³</td>
<td>Linear</td>
<td>HCl</td>
<td>HCl</td>
</tr>
</tbody>
</table>

EX. Find the e- and molecular geometries for the following molecules (we already drew the Lewis structures)
1. SiH₄  tetrahedral/tetrahedral
2. PO₄³⁻  tetrahedral/tetrahedral
3. $\text{H}_2\text{S}$ tetrahedral/bent
4. $\text{ClO}_3^-$ tetrahedral/trigonal pyramidal
5. $\text{NH}_4^+$ tetrahedral/tetrahedral
6. $\text{BH}_3$ trigonal planar/trigonal planar
7. $\text{C}_2\text{H}_4$ geometry around each C is trigonal planar/trigonal planar - ea C is $sp^2$ hybridized

**Summary of Hybridization**

<table>
<thead>
<tr>
<th>Type of bond</th>
<th>bond length</th>
<th>bond strength</th>
<th>% hybridization</th>
<th>s-character</th>
<th>bond angle*</th>
</tr>
</thead>
<tbody>
<tr>
<td>single</td>
<td></td>
<td></td>
<td>$sp^3$</td>
<td>25%</td>
<td>109.5</td>
</tr>
<tr>
<td>double</td>
<td>$\uparrow$</td>
<td>$\downarrow$</td>
<td>$sp^2$</td>
<td>33.3%</td>
<td>120</td>
</tr>
<tr>
<td>triple</td>
<td></td>
<td></td>
<td>$sp$</td>
<td>50%</td>
<td>180</td>
</tr>
</tbody>
</table>

*Exceptions:
- ammonia – one lone pair causes bond angles to be less than 109.5 (107.3 actually)
- water - two lone pairs cause bond angle to be less than 109.5 (104.5 actually)

**Polar vs. nonpolar**

Need to look at bond polarity/nonpolarity *as well as* molecular geometry.

See also Table 1.3 in text

**Difference in EN:**

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>EN difference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Covalent</td>
<td>0.0-0.4</td>
</tr>
<tr>
<td>Polar Covalent</td>
<td>0.5-1.7</td>
</tr>
<tr>
<td>Ionic</td>
<td>1.8 +</td>
</tr>
</tbody>
</table>

**Electronegativities (EN)**

<table>
<thead>
<tr>
<th>H 2.1</th>
<th>Electronegativities (EN)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>1.0</td>
</tr>
<tr>
<td>Be</td>
<td>1.5</td>
</tr>
<tr>
<td>B</td>
<td>2.0</td>
</tr>
<tr>
<td>C</td>
<td>2.5</td>
</tr>
<tr>
<td>N</td>
<td>3.0</td>
</tr>
<tr>
<td>O</td>
<td>3.5</td>
</tr>
<tr>
<td>F</td>
<td>4.0</td>
</tr>
<tr>
<td>Na</td>
<td>0.9</td>
</tr>
<tr>
<td>Mg</td>
<td>1.2</td>
</tr>
<tr>
<td>Al</td>
<td>1.5</td>
</tr>
<tr>
<td>Si</td>
<td>1.8</td>
</tr>
<tr>
<td>P</td>
<td>2.1</td>
</tr>
<tr>
<td>S</td>
<td>2.5</td>
</tr>
<tr>
<td>Cl</td>
<td>3.0</td>
</tr>
<tr>
<td>K</td>
<td>0.8</td>
</tr>
<tr>
<td>Ca</td>
<td>1.0</td>
</tr>
<tr>
<td>Ga</td>
<td>1.6</td>
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<tr>
<td>Ge</td>
<td>1.8</td>
</tr>
<tr>
<td>As</td>
<td>2.0</td>
</tr>
<tr>
<td>Se</td>
<td>2.4</td>
</tr>
<tr>
<td>Br</td>
<td>2.8</td>
</tr>
<tr>
<td>Rb</td>
<td>0.8</td>
</tr>
<tr>
<td>Sr</td>
<td>1.0</td>
</tr>
<tr>
<td>In</td>
<td>1.7</td>
</tr>
<tr>
<td>Sn</td>
<td>1.8</td>
</tr>
<tr>
<td>Sb</td>
<td>1.9</td>
</tr>
<tr>
<td>Te</td>
<td>1.9</td>
</tr>
<tr>
<td>I</td>
<td>1.5</td>
</tr>
<tr>
<td>Cs</td>
<td>0.7</td>
</tr>
<tr>
<td>Ba</td>
<td>0.9</td>
</tr>
<tr>
<td>Ti</td>
<td>1.8</td>
</tr>
<tr>
<td>Pb</td>
<td>1.9</td>
</tr>
<tr>
<td>Bi</td>
<td>2.0</td>
</tr>
<tr>
<td>Po</td>
<td>2.1</td>
</tr>
<tr>
<td>At</td>
<td>2.1</td>
</tr>
</tbody>
</table>

**Flow Chart**

Does the molecule have polar covalent bonds?  
**Yes**  
Is the Lewis structure symmetrical?  
**Yes**  
**Nonpolar**  
(Ex. $\text{CCl}_4 \ \varepsilon = 2.2$)  
**No**  
**Polar**  
(Ex. $\text{CH}_2\text{Cl}_2 \ \varepsilon = 9.1, \ \text{CH}_3\text{OH} \ \varepsilon = 30-33$)

EX. Show the direction of polarity of each of the indicated bonds (using $\delta$s or an arrow).

1. HO-H  
2. H$_3$C-NH$_2$  
3. H$_3$C-MgBr
EX. Indicate whether each of the molecules is polar or nonpolar.

1. HO-H \textit{polar (bent geometry)}
2. H$_3$C-NH$_2$ \textit{polar (C-N bond)}
3. H$_3$C-MgBr, polar (bent geometry and Mg-C diff than Mg-Br)
4. I-Cl \textit{polar (I-Cl bond)}
5. H$_2$N-OH, polar (bent geometry)