1. (4 pts) Which molecule is the most polar?
   a) HBr $\rightarrow \Delta \chi = 0.7$
   b) CO $\rightarrow \Delta \chi = 1.0$
   c) CH$_3$CH$_2$CH$_2$CH$_3$ $\rightarrow \Delta \chi = 0.14$ in LONGOR of DIRECTION S
   d) Cl$_2$ $\rightarrow \Delta \chi = 0$
   e) These molecules are all equally polar

2. In this problem you will compare two triatomic molecules: H$_2$O and O$_3$
   a) (2 pts) Give the Lewis structure for H$_2$O.

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

   b) (4 pts) Give the Lewis structure for O$_3$.

   \[
   \begin{array}{c}
   \text{O} \\
   \text{O} \quad \text{O}
   \end{array}
   \]

   c) (2 pts) Describe the molecular shape of H$_2$O

   Bent

   d) (2 pts) Describe the molecular shape of O$_3$

   Bent

   e) (5 pts) Compare and contrast the shapes of the two molecules. Be as quantitative and descriptive as possible.

   Water has a tetrahedral electron group geometry with two lone pairs. Its bond angle is $\approx 109.5^\circ$.

   Ozone has trigonal planar electron group geometry with one lone pair. Its bond angle is $\approx 120^\circ$. 
3. (4 pts) Which of the following molecules have a nonzero dipole moment? Circle all that do.

(a) ClF₃  (b) (NH₂)₂CO  (c) CO₂  (d) XeF₄

4. a) (4 pts) Draw the Lewis structure for BrF₅.

\[ \text{Br} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} \]

7 + 5 + 7 = 42e⁻

b) (2 pts) Give the electron group geometry for BrF₅.

Octahedral

c) (2 pts) Give the molecular shape of BrF₅.

Square Pyramid

d) (2 pts) Does BrF₅ have a dipole moment?

Yes

e) (3 pts) Give the hybridization of the Br atom in BrF₅.

\[ sp^3d^2 \]

The following space is provided for your own free expression.
5. In this problem you will examine the relative bond strengths of NO\(^+\) and NO\(^-\).

a) (8 pts) First, sketch the valence molecular orbital energy diagrams for NO\(^+\) and NO\(^-\). For full credit, be sure to include all of the valence electrons and to label each molecular orbital with its bond type (\(\sigma, \sigma^*, \pi, \text{ or } \pi^*\)) and the atomic orbitals it was generated from (s or p). (HINT: The energy level ordering in NO is the same as in \(\text{O}_2\).)

\[
\begin{align*}
\text{NO}^+ & : 10e^- \\
\text{NO}^- & : 12e^-
\end{align*}
\]

b) (4 pts) Give the valence electron configuration for each molecule.
\(\text{e.g. } (\sigma_{2s})^2(\sigma^*_{2s})^2\ldots\)

\[
\begin{align*}
\text{NO}^+: & \quad (\jmath_{2s})^2(\jmath^*_{2s})^2(\jmath_{2p})^2(\jmath^*_{2p})^4 \\
\text{NO}^-: & \quad (\jmath_{2s})^2(\jmath^*_{2s})^2(\jmath_{2p})^2(\jmath^*_{2p})^4(\jmath^*_{2p})^2
\end{align*}
\]

c) (4 pts) Calculate the bond order for each molecule.

\[
\begin{align*}
\text{NO}^+: & \quad \frac{1}{2}(6-0) = 3 \\
\text{NO}^-: & \quad \frac{1}{2}(6-2) = 2
\end{align*}
\]

d) (2 pts) Which molecule has the stronger bond (or are they the same)?

\[\text{NO}^+\]
e) (3 pts) Sketch the highest energy molecular orbital that is occupied by electrons in NO⁺

\[ \pi^* \]

f) (3 pts) Sketch the highest energy molecular orbital that is occupied by electrons in NO⁻

\[ \pi^* \]

6. Consider the following two molecules—two different forms of C₄H₄O:

![Molecules A and B](image)

a) (4 pts) Indicate the hybridization of all of the carbon atoms in both structures.

b) (2 pts) Choose the best statement:
   i) A is more stable than B
   ii) B is more stable than A
   iii) Their stabilities are essentially the same

c) (4 pts) Explain your choice.

A includes conjugation over 3 C and the O, which stabilizes the molecule. In B, there are 2 single bonds between the double bonds, so there is no conjugation.
7. Consider the following 2 sets of possible Lewis structures for SO₂.

I) \[
\begin{align*}
\text{O} & \equiv \text{S} - \text{O} \\
\text{O} & \quad +1 \\
\text{O} & \quad -1
\end{align*}
\]

\[
\text{O} \quad \text{S} \quad \text{O}
\]

II) \[
\begin{align*}
\text{O} & \equiv \text{S} = \text{O} \\
\text{O} & \quad 0 \\
\text{O} & \quad 0
\end{align*}
\]

a) (6 pts) Indicate the formal charge on each atom in the above structures (all 9 of them).

b) (24 pts) Complete the following table. Note that a typical S-O single bond length is 157 pm. A typical S=O double bond length is 142 pm.

<table>
<thead>
<tr>
<th>Structures</th>
<th>I</th>
<th>II</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron Group Geometry (about the sulfur)</td>
<td>TRIGONAL PLANAR</td>
<td>TRIGONAL PLANAR</td>
</tr>
<tr>
<td>Molecular Shape (about the sulfur)</td>
<td>BENT</td>
<td>BENT</td>
</tr>
<tr>
<td>O-S-O Angle</td>
<td>&lt; 120°</td>
<td>&lt; 120°</td>
</tr>
<tr>
<td>The two O-S bond lengths are: (same/different)</td>
<td>SAME</td>
<td>SAME</td>
</tr>
<tr>
<td>Estimate of the length of the ‘left’ O-S bond</td>
<td>150 pm</td>
<td>142 pm</td>
</tr>
<tr>
<td>Estimate of the length of the ‘right’ O-S bond</td>
<td>150 pm</td>
<td>142 pm</td>
</tr>
</tbody>
</table>