5. Structure, Geometry, and Polarity of Molecules

What you will accomplish in this experiment

This experiment will give you an opportunity to draw Lewis structures of covalent compounds, then use those structures to visualize the three-dimensional shapes of the molecules. You’ll also learn how to determine the polarity of the covalent bonds within the molecules.

You need to build these skills so that you’ll ultimately be able to make predictions about the behavior of molecules, specifically their ability to interact with neighboring molecules via intermolecular attractive forces. These interactions will dictate the physical properties of these covalent compounds (their physical state and solubility).

In summary, this experiment will help you to practice the following skills:

- Draw Lewis structures of simple molecules (containing one to three central atoms).
- Use the number of electron “groups” around each central atom in the structure to determine the three-dimensional geometry (spatial arrangement) of the bonding atoms and lone pairs.
- Assess the polarity of the covalent bonds in the structure by comparing the electronegativities of the two atoms sharing electrons.

Concepts you need to know to be prepared

You’ve learned that compounds are a chemical combination of elements, meaning that they’re created when two or more elements chemically react with one another. The driving force for that reaction is that the elements are trying to achieve the stability of a noble gas electron configuration (for most elements, this means an $s^2p^6$ “octet” of electrons).

In Lab 4, you saw that when the reacting elements are a metal and a nonmetal (elements with an electronegativity difference greater than 2.0), the metal atom achieves its octet when it loses one or more electrons (to become a positively-charged ion, a “cation”), and the nonmetal atom gains those electrons (to become a negatively-charged ion, an “anion”).

The chemical formula for the ionic compound that’s formed describes the fixed proportion in which vast numbers of the oppositely-charged ions must combine in order to appropriately balance out the positive and negative charges.
Covalent Bonds and Lewis Structures

When the reacting elements are nonmetals with a small electronegativity difference (less than 2.0), the atoms achieve their octet NOT by gaining or losing electrons, but by sharing electrons – by forming covalent bonds. For most nonmetal atoms, this means sharing enough electrons to achieve the “octet” (the $s^2p^6$ electron configuration of all noble gases except helium).

Lewis structures are illustrations of the covalent bond connections between atoms. When drawn correctly, these structures will help you to envision the three-dimensional geometry of each bonding atom. The polarity of each covalent bond, considered with the geometries of the bonding atoms, will enable you to make predictions about the behavior of the molecule, specifically its ability to interact with other molecules via intermolecular attractive forces.

Drawing Lewis Structures

The atoms that need the most electrons to achieve a noble gas electron configuration (and thus form the most covalent bonds) should be used as central atoms in a Lewis structure.

The atoms that need very few electrons to achieve a noble gas electron configuration (and thus form very few bonds) should be used as peripheral atoms in the structure.

Examples of central atoms:

- Carbon has 4 valence electrons. It needs 4 more electrons to achieve an octet. It gets those 4 electrons by sharing: by forming 4 covalent bonds. CARBON’S octet: 4 BONDS and NO LONE PAIR.

- Nitrogen has 5 valence electrons. It needs 3 more electrons to achieve an octet. It gets those 3 electrons by sharing: by forming 3 covalent bonds. NITROGEN’S octet: 3 BONDS and 1 LONE PAIR.

- Oxygen has 6 valence electrons. It needs 2 more electrons to achieve an octet. It gets those 2 electrons by sharing: by forming 2 covalent bonds. OXYGEN’S octet: 2 BONDS and 2 LONE PAIR.

Examples of peripheral atoms:

- Fluorine has 7 valence electrons. It needs 1 more electron to achieve an octet. It gets that 1 electron by sharing: by forming 1 covalent bond. FLUORINE’S octet (and that of the other halogens – chlorine, bromine, and iodine): 1 BOND and 3 LONE PAIR.

- Hydrogen has 1 valence electron. It needs 1 more electron to achieve a duet (like Helium). It gets that 1 electron by sharing: by forming 1 covalent bond. HYDROGEN’S duet: 1 BOND.

You should indicate the covalent bonds between the atoms with lines, and the lone pairs of electrons (unshared electrons) with dots. And you’ll need to arrange the bonds and lone pairs so that each atom achieves its noble gas electron configuration (an octet, or a duet for hydrogen atoms).

- A “single bond” is one pair of electrons shared between two atoms, and is represented in the Lewis structure by a single line between the bonding atoms.

- A “double bond” is two pairs of electrons shared between two atoms, and is represented by drawing two lines between the bonding atoms. Double bonds are most common between atoms of carbon, nitrogen, oxygen, and sulfur. (In organic and biological molecules, there are many examples of carbon-carbon, carbon-nitrogen, and carbon-oxygen double bonds.)

- A “triple bond” is three pairs of electrons shared between two atoms, and is represented by drawing three lines between the bonding atoms. Triple bonds are most common between carbon and nitrogen atoms.

Be sure to draw NO MORE and NO FEWER electrons than the TOTAL number of valence electrons that the atoms bring with them to the molecule.
Example Lewis Structure: Formaldehyde, CH₂O

The TOTAL number of valence electrons the atoms bring with them to the molecule:

1 C atom with 4 valence electrons 4
2 H atoms, each with 1 valence electron 2(1)
1 O atom with 6 valence electrons 6
Total available electrons 12

Consider the bonding patterns of our component atoms:

- Carbon is always a central atom because it forms 4 bonds.
- Oxygen tends to form 2 bonds and 2 lone pairs.
- Hydrogen atoms form just 1 bond.

Since carbon forms the most bonds, we’ll place it at the center and attach the O and H atoms to it. Since the C atom needs four bonds, and the O atoms need two bonds, we’ll place a double bond between them. 2 lone pairs of electrons are added to the oxygen to complete its octet. The total number of valence electrons the atoms brought into the molecule is 12. This Lewis structure accounts for ALL 12 of those electrons: 2 electrons in each of the 4 bonds (8 bonding electrons), plus 2 lone pairs (4 unshared/non-bonding electrons).

Molecular Geometry

The repulsive forces between bonding and non-bonding electrons determine the three-dimensional geometry (spatial arrangement) of the groups of electrons around a central atom.

Because the negative charges repel one another, the electron groups arrange themselves so that they are as far apart from one another as possible. This idea is known as the “Valence Shell Electron Pair Repulsion” (VSEPR) Theory.

Thus the geometry of the electrons around each central atom in a Lewis structure can be determined by:

1. Simply counting the “groups” of electrons that surround that atom, and then
2. Considering how these groups would arrange themselves so as to be as far apart as possible.

It’s essential to remember that a “group” of electrons is one of the following:

- A **single bond**, a **double bond**, a **triple bond**, or a **lone pair** of electrons

For example, the Lewis structure of formaldehyde indicates 3 groups of electrons around the central carbon atom: 2 single bonds (between the carbon and each of the two hydrogen atoms), and 1 double bond (between the carbon and oxygen atoms).

As shown in the list of potential “Electron Group Geometries,” in order for three groups of electrons to be as far apart as possible, they would all have to lie in the same plane and be directed toward the corners of a triangle. This electron group geometry is called “trigonal planar.” The bond angles in this geometry would be 120°.

- **TWO groups** of electrons around a central atom: LINEAR
- **THREE groups** of electrons around a central atom: TRIGONAL PLANAR
- **FOUR groups** of electrons around a central atom: TETRAHEDRAL

The table on the next page indicates the overall “shape” (or “Molecular Geometry”) of a central atom depending on whether the electron groups around it are covalent bonds to other atoms or simply lone pairs of electrons.
IF the electron groups are bonds, then atoms are present to “mark the corners” of each of the spatial arrangements: the two ends of the linear geometry, the three corners of the trigonal planar geometry, or the four corners of the tetrahedral geometry.

But – IF an electron group is a lone pair, then there is no atom visible to “mark that corner” of the geometry.

Thus the “Electron Group Geometry” is determined by the number of groups of electrons and the arrangement of those groups around the central atom.

But the “Molecular Geometry” depends on whether or not there is an atom to visibly mark each corner of the Electron Group Geometry.

Remember that you can only “see” atoms; you cannot “see” lone pairs of electrons.

<table>
<thead>
<tr>
<th># of Groups of Electrons</th>
<th>Electron Group Geometry</th>
<th>Number of Lone Pairs</th>
<th>Molecular Geometry</th>
<th>Approximate Bond Angles</th>
<th>Example Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear</td>
<td>0</td>
<td>Linear</td>
<td>180°</td>
<td>carbon dioxide, CO₂</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal Planar</td>
<td>0</td>
<td>Trigonal Planar</td>
<td>120°</td>
<td>formaldehyde, CH₂O</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral</td>
<td>0</td>
<td>Tetrahedral</td>
<td>109.5°</td>
<td>methane, CH₄</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>ammonia, NH₃</td>
</tr>
<tr>
<td></td>
<td></td>
<td>1</td>
<td>Trigonal Pyramid</td>
<td>107°</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>2</td>
<td>Angular (Bent)</td>
<td>104.5°</td>
<td>water, H₂O</td>
</tr>
</tbody>
</table>
The molecule used as the final example on the preceding table is water, H\textsubscript{2}O. The Lewis structure for water indicates that the oxygen atom is surrounded by four groups of electrons: two bonds and two lone pairs. These four groups will be arranged in the tetrahedral “Electron Group Geometry.” Hydrogen atoms mark the positions at two of the corners of this tetrahedron, but the lone pairs of electrons at the other two positions make those corners invisible. Thus the “shape” (“Molecular Geometry”) of the water molecule is actually “angular” or “bent.”

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

Polarity of Molecules

When the two atoms sharing electrons in a covalent bond have significantly different electronegativity values (a difference greater than 0.4, but less than 2.0), the electrons in the bond are NOT shared EQUALLY between the two atoms.

The element that has the higher electronegativity has greater “electron pulling power;” that is, the atom is able to attract the electrons in the bond toward itself. The result is that the covalent bond is “polarized.” Or to put it more simply, we say the bond is “polar.”

Since the bonding electrons spend more of their time in the region around the element with the higher electronegativity, that atom takes on a partial negative charge (represented by the symbol \(\delta^-\)). As the bonding electrons spend very little time in the vicinity of the less electronegative element, that atom takes on a partial positive charge (represented by the symbol \(\delta^+\)).

So, similar to a battery, a polar covalent bond has TWO “poles:” one partial positive (\(\delta^+\)), and the other partial negative (\(\delta^-\)). Thus we can say that it’s a “dipole,” and we can represent the polarized bonding electrons in an alternative way: by using a dipole moment vector. That is, we can draw a vector (simply an arrow) pointing in the direction that the electrons are being pulled: toward the atom with the higher electronegativity. To emphasize the fact that the atom at the other end of the bond has a partial positive charge, we make the “tail” of the arrow into a “plus” sign.

Thus the very polar hydrogen fluoride bond (electronegativities: \(\text{H} = 2.1, \text{F} = 4.0\)) could be represented either by labeling the two atoms with the partial charge symbols or by drawing a dipole moment vector:

\[
\begin{array}{c}
\delta^- \quad \delta^+ \\
\text{F} - \text{H} \\
\end{array}
\]

The geometry of a molecule is tremendously important when evaluating the polarity of each covalent bond around a central atom (in order to determine the type of intermolecular attractive forces that a molecule is capable of using). Quite often the three-dimensional geometry of the central atom’s bonding and nonbonding electron groups can enhance the overall polarity of a cluster of atoms. For example, consider the water molecule, H\textsubscript{2}O.
The hydrogen-oxygen single bonds in the angular H₂O molecule are polar (electronegativities: H = 2.1, O = 3.5). Thus the electrons in each bond are drawn toward the oxygen atom. This can be illustrated using either the partial charge symbols or dipole moment vectors.

But now consider the tetrahedral geometry of the four groups of electrons around the oxygen atom (2 bonds, 2 long pairs). This geometry greatly enhances the overall polarity of the angular (bent) water molecule. The dipole moment vectors are extremely useful in illustrating this additive effect of the two polar bonds. The partial negative charge of both dipoles is focused at the oxygen atom. Additionally, the oxygen’s two lone pairs of electrons contribute significantly to its overall partial negative charge. The result is that the water molecule is extremely polar, with the partial negative charge centered at the oxygen atom, and a partial positive charge at each of the two hydrogen atoms.

Occasionally, for very small molecules, the symmetry of the bonding electrons around the central atom can cause two or more dipole moment vectors to completely cancel one another, so that the molecule as a whole has no NET dipole. For example, consider the carbon dioxide (CO₂) molecule.

The carbon-oxygen double bonds in the linear CO₂ molecule are polar (electronegativities: C = 2.5, O = 3.5). The electrons in each of the double bonds are drawn toward the oxygens, so both oxygen atoms have a partial negative charge. This can be illustrated using either the partial charge symbols or dipole moment vectors.

The vectors are especially helpful in illustrating that, because of the linear geometry of the molecule, the bond dipoles are completely symmetric to one another. They are equal and opposite, so they cancel (the CENTERS of partial-negative and partial-positive charge are at the same place in the molecule). Thus the carbon dioxide molecule as a whole is nonpolar, even though it contains polar bonds.
Procedure that you will follow

You and your lab partner will be supplied with a molecular modeling kit. The colors of the plastic balls in the kit correspond to atoms of different elements as follows:

- carbon atoms (black)
- hydrogen atoms (white)
- oxygen atoms (red)
- nitrogen atoms (blue)
- chlorine atoms (green)
- bromine atoms (orange)

For each chemical formula specified on the report sheet, you will:

1) **Draw the Lewis structure** of the molecule. (Note that for one of the chemical formulas, **TWO different** Lewis structures are possible. These **different** structures that have the **same** chemical formula are said to be "structural isomers" of one another.)

2) Determine the **Electron Group** and **Molecular Geometry** of each central atom.

3) Verify that **geometry** by constructing a model of the molecule using the atoms (colored balls) and covalent bonds (plastic connectors) in the kit.

You’ll then:

4) Assess the **polarity** of the **covalent bonds** by comparing the **electronegativities** of the bonding atoms.

Electronegativity values of common Representative Elements are provided in the table below:

<table>
<thead>
<tr>
<th>Increasing electronegativity</th>
<th>Decreasing electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td></td>
</tr>
<tr>
<td>Li  1.0</td>
<td>B</td>
</tr>
<tr>
<td>Be  1.5</td>
<td>C</td>
</tr>
<tr>
<td>Na  0.9</td>
<td>N</td>
</tr>
<tr>
<td>Mg  1.2</td>
<td>O</td>
</tr>
<tr>
<td>Al  1.5</td>
<td>F</td>
</tr>
<tr>
<td>Si  1.8</td>
<td>P</td>
</tr>
<tr>
<td>S   1.7</td>
<td>Cl</td>
</tr>
<tr>
<td>Ge  1.8</td>
<td>As</td>
</tr>
<tr>
<td>Se  1.9</td>
<td>Br</td>
</tr>
<tr>
<td>In  1.9</td>
<td>Te</td>
</tr>
<tr>
<td>Sn  1.9</td>
<td>I</td>
</tr>
<tr>
<td>Sb  1.9</td>
<td>Cs</td>
</tr>
<tr>
<td>Te  1.9</td>
<td>Ba</td>
</tr>
<tr>
<td>I   2.5</td>
<td>Cs</td>
</tr>
<tr>
<td>0.7</td>
<td>0.9</td>
</tr>
</tbody>
</table>
5. Report Sheet: Structure, Geometry, and Polarity of Molecules

Student ____________________ Lab Partner ____________________ Date Lab Performed ____________

Section #_________ Lab Instructor ____________________ Date Report Received ____________

I. Hydrogen Cyanide, HCN

<table>
<thead>
<tr>
<th></th>
<th>H atom</th>
<th>C atom</th>
<th>N atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>How many valence electrons does this atom have?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many electrons does this atom need to get an octet (duet for H)?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>So how many covalent bonds will this atom form?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many lone pairs will this atom have?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>What is the electronegativity value of this atom?</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a) TOTAL number of valence electrons in the HCN molecule = _______________
b) Draw a Lewis structure for the HCN molecule, including all covalent bonds and all lone pairs of electrons. Your structure should have no more and no fewer electrons than the TOTAL number above.

c) How many groups of electrons are around your central atom? _____________
   (Remember that a “group” of electrons can be a: single bond, double bond, triple bond, or lone pair.)
d) What is the electron group geometry of the central atom? ____________________________

After noting whether the electron groups are covalent bonds or lone pairs:
e) What is the molecular geometry of the central atom? ____________________________

Now use your kit to construct a molecular model of the HCN molecule to verify the molecular geometry.
Compare the electronegativities of the two atoms participating in each covalent bond.
f) Are any of the bonds POLAR? _____________

If YES, clearly indicate the partial negative and partial positive atoms in your Lewis structure above.
g) Do the bond polarities reinforce one another (to make a net dipole), or cancel one another (to make no net dipole)? Explain.
II. Difluoromethane, $\text{CH}_2\text{F}_2$

<table>
<thead>
<tr>
<th>Each H atom</th>
<th>C atom</th>
<th>Each F atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>How many valence electrons does this atom have?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many electrons does this atom need to get an octet (duet for H)?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>So how many covalent bonds will this atom form?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many lone pairs will this atom have?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>What is the electronegativity value of this atom?</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a) TOTAL number of valence electrons in the $\text{CH}_2\text{F}_2$ molecule = ________________

b) Draw a Lewis structure for the $\text{CH}_2\text{F}_2$ molecule, including all covalent bonds and all lone pairs of electrons. Your structure should have no more and no fewer electrons than the TOTAL number above.

c) How many groups of electrons are around your central atom? ________________

(Remember that a “group” of electrons can be a: single bond, double bond, triple bond, or lone pair.)

d) What is the electron group geometry of the central atom? ____________________________

After noting whether the electron groups are covalent bonds or lone pairs:

e) What is the molecular geometry of the central atom? ____________________________

Now use your kit to construct a molecular model of the $\text{CH}_2\text{F}_2$ molecule to verify the molecular geometry.

Compare the electronegativities of the two atoms participating in each covalent bond.

g) Are any of the bonds POLAR? ________________

If YES, clearly indicate the partial negative and partial positive atoms in your Lewis structure above.

h) Do the bond polarities reinforce one another (to make a net dipole), or cancel one another (to make no net dipole)? Explain.
II. Dichloroethylene, \( \text{C}_2\text{H}_2\text{Cl}_2 \)

<table>
<thead>
<tr>
<th></th>
<th>Each H atom</th>
<th>( \text{C} ) atom</th>
<th>Each Cl atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>How many valence electrons does this atom have?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many electrons does this atom need to get an octet (duet for H)?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>So how many covalent bonds will this atom form?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many lone pairs will this atom have?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>What is the electronegativity value of this atom?</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a) \textbf{TOTAL} number of valence electrons in the \( \text{C}_2\text{H}_2\text{Cl}_2 \) molecule = ________________

b) \textbf{Draw a Lewis structure for the \( \text{C}_2\text{H}_2\text{Cl}_2 \) molecule, including all covalent bonds and all lone pairs.}
Recall that the video your lab instructor played for you showed two DIFFERENT Lewis structures for this chemical formula. \textbf{The Lewis structure you draw HERE should be DIFFERENT from those shown in the video.}
(And remember that your structure should have no more and no fewer electrons than the \textbf{TOTAL} number above.)

\textbf{In order to answer the questions below, number the two carbon atoms in your structure (1 and 2).}

c) How many groups of electrons are around carbon atom \#1? ________________

d) What is the electron group geometry of carbon atom \#1? ____________________________

e) What is the molecular geometry of carbon atom \#1? ____________________________

f) How many groups of electrons are around carbon atom \#2? ________________

g) What is the electron group geometry of carbon atom \#2? ____________________________

h) What is the molecular geometry of carbon atom \#2? ____________________________

\textbf{Now use your kit to construct a molecular model of the \( \text{C}_2\text{H}_2\text{Cl}_2 \) molecule to verify the molecular geometries.}

\textbf{Compare the electronegativities of the two atoms participating in each covalent bond.}

i) Are any of the bonds POLAR? ________________

\textbf{If YES, clearly indicate the partial negative and partial positive atoms in your Lewis structure above.}

j) Do the bond polarities reinforce one another (to make a net dipole), or cancel one another (to make no net dipole)? \textbf{Explain.}
**IV. Methyl amine, CH₃N**

<table>
<thead>
<tr>
<th></th>
<th>Each H atom</th>
<th>C atom</th>
<th>N atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>How many <em>valence electrons</em> does this atom have?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many electrons does this atom <em>need</em> to get an <em>octet</em> (duet for H)?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>So how many <em>covalent bonds</em> will this atom form?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many <em>lone pairs</em> will this atom have?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>What is the * electronegativity* value of this atom?</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a) *TOTAL* number of valence electrons in the CH₃N molecule = ________________

b) **Draw a Lewis structure for the CH₃N molecule, including all covalent bonds and all lone pairs of electrons.** Your structure should have no more and no fewer electrons than the *TOTAL* number above.

c) How many groups of electrons are around the *CARBON* atom? ________________
d) What is the *electron group geometry* of the *CARBON* atom? _______________________________
e) What is the *molecular geometry* of the *CARBON* atom? _______________________________
f) How many groups of electrons are around the *NITROGEN* atom? ________________
g) What is the *electron group geometry* of the *NITROGEN* atom? _______________________________
h) What is the *molecular geometry* of the *NITROGEN* atom? _______________________________

**Now use your kit to construct a molecular model of the CH₃N molecule to verify the molecular geometries.**

*Compare the electronegativities of the two atoms participating in each covalent bond.*

i) Are any of the bonds *POLAR*? ________________

*If YES, clearly indicate the partial negative and partial positive atoms in your Lewis structure above.*

j) Do the bond polarities *reinforce* one another (to make a net dipole), or *cancel* one another (to make no net dipole)? *Explain.*
V & VI. Ethanol and Dimethyl Ether, \( C_2H_6O \)

<table>
<thead>
<tr>
<th>Each H atom</th>
<th>Each C atom</th>
<th>O atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>How many valence electrons does this atom have?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many electrons does this atom need to get an octet (duet for H)?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>So how many covalent bonds will this atom form?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many lone pairs will this atom have?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>What is the electronegativity value of this atom?</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a) *TOTAL* number of valence electrons in the \( C_2H_6O \) molecule = ________________

b) *Draw a Lewis structure for the \( C_2H_6O \) molecule, including all covalent bonds and all lone pairs of electrons.* Your structure should have no more and no fewer electrons than the *TOTAL* number above.

---

In order to answer the questions below, number the two carbon atoms in your structure (1 and 2).

c) How many groups of electrons are around CARBON atom #1? ________________

d) What is the electron group geometry of CARBON atom #1? _________________________________

e) What is the molecular geometry of CARBON atom #1? _________________________________

f) How many groups of electrons are around CARBON atom #2? ________________

g) What is the electron group geometry of CARBON atom #2? _________________________________

h) What is the molecular geometry of CARBON atom #2? _________________________________

i) How many groups of electrons are around the OXYGEN atom? ________________

j) What is the electron group geometry of the OXYGEN atom? _________________________________

k) What is the molecular geometry of the OXYGEN atom? _________________________________

Now use your kit to construct a molecular model of the \( C_2H_6O \) molecule to verify the molecular geometries.

**Compare the electronegativities of the two atoms participating in each covalent bond.**

l) Are any of the bonds POLAR? ________________

If YES, clearly indicate the partial negative and partial positive atoms in your Lewis structure above.

m) Do the bond polarities reinforce one another (to make a net dipole), or *cancel* one another (to make no net dipole)? *Explain.*
### V & VI. Ethanol and Dimethyl Ether, $C_2H_6O$

<table>
<thead>
<tr>
<th>Each H atom</th>
<th>Each C atom</th>
<th>O atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>How many valence electrons does this atom have?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many electrons does this atom need to get an octet (duet for H)?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>So how many covalent bonds will this atom form?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>And how many lone pairs will this atom have?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>What is the electronegativity value of this atom?</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a) TOTAL number of valence electrons in the $C_2H_6O$ molecule = ________________
b) Draw a Lewis structure for the $C_2H_6O$ molecule, including all covalent bonds and all lone pairs of electrons. Your structure should have no more and no fewer electrons than the TOTAL number above.

In order to answer the questions below, number the two carbon atoms in your structure (1 and 2).

<table>
<thead>
<tr>
<th></th>
<th>1</th>
<th>2</th>
</tr>
</thead>
<tbody>
<tr>
<td>c) How many groups of electrons are around CARBON atom #1?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>d) What is the electron group geometry of CARBON atom #1?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>e) What is the molecular geometry of CARBON atom #1?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>f) How many groups of electrons are around CARBON atom #2?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>g) What is the electron group geometry of CARBON atom #2?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>h) What is the molecular geometry of CARBON atom #2?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>i) How many groups of electrons are around the OXYGEN atom?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>j) What is the electron group geometry of the OXYGEN atom?</td>
<td></td>
<td></td>
</tr>
<tr>
<td>k) What is the molecular geometry of the OXYGEN atom?</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Now use your kit to construct a molecular model of the $C_2H_6O$ molecule to verify the molecular geometries.

Compare the electronegativities of the two atoms participating in each covalent bond.

l) Are any of the bonds POLAR? ________________

If YES, clearly indicate the partial negative and partial positive atoms in your Lewis structure above.

m) Do the bond polarities reinforce one another (to make a net dipole), or cancel one another (to make no net dipole)? Explain.