**Chemical Bonds, Molecular Models, and Molecular Shapes**

**Prelab Assignment**
Write an objective, and answer the following questions in your laboratory notebook before coming to lab. Read the entire laboratory write up before answering the prelab questions.

1. Briefly explain how VSEPR theory explains electron distribution within a molecule and molecular shape.
2. What do a solid triangle and a striped triangle indicate when drawing a molecular shape in two dimensions?
3. Do ball and stick molecular models illustrate the exact shape of a molecule?
4. In this lab you will build models of molecules having the five common molecular shapes linear, bent, trigonal planar, trigonal pyramidal, and tetrahedral. Sketch each of these shapes using the solid and striped triangle method described in this lab.
5. In this lab you will build models of molecules having the five common molecular shapes linear, bent, trigonal planar, trigonal pyramidal, and tetrahedral. What other molecular shapes are described in your textbook?
6. What is the difference between a molecule or ion’s electron domain (also called electron pair or orbital shape) and molecular shape?
7. Briefly describe how you would determine whether a molecule is polar or non-polar. Note that there are two different characteristics you should evaluate. You may find the answer to this question either in your book or in the lecture notes.

**Background Information**
The properties of chemical compounds are directly related to the ways in which atoms are bonded together to make molecules. In this exercise you will use simple “ball-and-stick” models to construct representations of some common molecules. These “ball-and-stick” models serve as a three-dimensional representation of an abstract idea. Although it is possible to represent molecular structures with reasonable accuracy using relatively simple models, modern molecular representations are a good deal more abstract and often mathematical in nature.

The chemical bonds that hold atoms together in molecules consist of pairs of electrons (or two electrons) shared or transferred between atoms. Atoms tend to share valence, or outer, electrons in such a way that each non-metal atom has a share in an octet, or eight, valence electrons. The Group number of a main Group (Group A) element can tell you the number of valence electrons each element has. A single bond consists of two shared or transferred electrons. A double bond consists of four electrons, and a triple bond consists of six electrons. Recall that two electrons may be found in each orbital. Therefore, in molecules electrons are always found in pairs. **Lone electron pairs** are pairs of electrons that are not part of a bond.
Valence Shell Electron-Pair Repulsion (VSEPR) theory states that electron pairs repulse each other and therefore wish to be as far apart as possible. Thus molecular shapes result from bond and lone electron pairs spacing themselves as far apart as possible. The electron domain (also called the electron pair or orbital) shape of an atom includes the lone electron pairs as well as the bond pairs. The molecular shape is the shape of the atoms alone. Lone pairs actually repulse other electron pairs more strongly than do bond pairs. Therefore lone pairs strongly influence molecular shapes. Large molecules have more than one central atom and therefore more than one shape. The common shapes that we will examine in this lab are linear, bent, trigonal planar, trigonal pyramidal, and tetrahedral.

The three-dimensional shape of molecules is represented in two-dimensions by using solid lines to represent bonds in the plane of the paper. Bonds coming toward you out of the paper are represented by solid triangles. Bonds going backward into or behind the paper are represented by dashed triangles.
See the previous illustration and the illustration below.

Molecules may be either symmetric or asymmetric. **Symmetric molecules** are identical (or mirror images) on either side of planes through the center of the molecule in all three dimensions. **Asymmetric molecules** will be different on either side of the centered plane in at least one dimension.

**PROCEDURE**
Your kit contains colored spheres with spikes on them and hollow white tubes. Use the white tubes to form bonds connecting atoms represented by the colored spheres with spikes. Use the black spheres with spikes for carbon atoms or halogens when they are central atoms. Use the red spheres with spikes for nitrogen. Use blue spheres with spikes for oxygen. Use yellow spheres with spikes for sulfur. Use green spheres for chlorine or fluorine when they are not the central atom. Use the short white tubes with a single spike for hydrogen.

For each molecule in the table below:
1. Determine which element will be the central atom
2. Calculate the total number of valence electrons,
3. Draw the Lewis structure (unless it is already given),
4. Determine the number of lone pairs and bonds around each central atom,
5. Build the model,
6. Name the **electron domain (also called the electron pair or orbital) shape** around the central atom(s) of the molecule,
7. Name the **molecular shape(s)** of the molecule, and
8. Sketch the molecular shape(s) in your lab notebook using the conventions described and illustrated previously under the background section of this lab.
9. Indicate whether each shape is symmetric or asymmetric.

**Look for trends in molecular shape versus the electron pair shape of the molecule and the number of lone pairs in the molecule.**

**Make the models first. Then you may check your work.** You may find the following websites useful: [https://www.youtube.com/watch?v=Q9-JiyAEqnU](https://www.youtube.com/watch?v=Q9-JiyAEqnU) and [https://phet.colorado.edu/en/simulation/molecule-shapes](https://phet.colorado.edu/en/simulation/molecule-shapes)
<table>
<thead>
<tr>
<th>Molecule</th>
<th>Total Valence Electrons</th>
<th>Around the Central Atom(s)</th>
<th>Lewis Structure</th>
<th>Sketch of Molecular Shape (sketched as illustrated below)</th>
<th>e- Domain and Molecular Shape Name(s)</th>
<th>Symmetric or Asymmetric</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH$_4$</td>
<td>8</td>
<td>4</td>
<td>0</td>
<td><img src="image" alt="Lewis Structure" /></td>
<td>tetrahedral</td>
<td>symmetric</td>
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<tr>
<td>Methane</td>
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<tr>
<td>NH$_3$</td>
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<tr>
<td>Ammonia</td>
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<tr>
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<td>Cyanide</td>
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<td>Chemical Formula</td>
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<tr>
<td>HNO₃</td>
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<tr>
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<tr>
<td>CH₃CH₂OH</td>
<td>Ethanol</td>
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</tbody>
</table>
| **NaHCO₃**  
Sodium bicarbonate (baking soda) |   |   |   |   |
|-----------------------------------|---|---|---|---|
| **CH(CH₃)₃**  
Isobutane |   |   |   |   |
| ![Isobutane structure](image) |   |   |   |   |
| **C₉H₁₈**  
Octane |   |   |   |   |
| ![Octane structure](image) |   |   |   |   |
| **C₉H₈O**  
Cinnamaldehyde  
gives cinnamon its smell |   |   |   |   |
| ![Cinnamaldehyde structure](image) |   |   |   |   |
Linoleic acid (polyunsaturated fat) | NA | ![Molecular Structure](image) | Build this just for fun if you have time.

### QUESTIONS

Answer all questions in your laboratory notebook

1. What impact do lone pairs have on molecular shape and why do they have this impact?

2. Can a single molecule be described by more than one shape and if this is possible under what situations might this occur?

3. Based on the figure included in the background section of this lab, figures in your textbook, and the trends you observed between the orbital shape, the number of lone pairs, and the molecular shape, how might you determine a molecule’s molecular shape from its Lewis structure without using a model kit?

4. For each molecule you made in lab list whether or not the molecule will be polar or non-polar. If the molecule is polar indicate the direction of polarity.