CH 151 Summer 2025:
“Chemical Models”
Lab - Instructions

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-6-3 through Ia-6-6 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Watch the **video introduction** for this lab **here:** **http://mhchem.org/t/6.htm**

The video introduction will help prepare you for the lab and assist you in completing the work before turning it in to the instructor.

Step Three:

Bring the printed copy of the lab with you on Friday, July 9. During lab in room AC 2509, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Four:

Complete the lab work and calculations on your own, then **turn it in (**pages Ia-6-5 through Ia-6-6 *only* to avoid a point penalty**) at 8 AM on Monday, July 14.** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor. Good luck on this assignment!

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Chemical Models Lab

A chemical bond is a force that holds groups of two or more atoms together and makes them function as a unit. Bonding involves only the valence (outer shell) electrons of atoms. In an **ionic bond**, valence electrons are transferred from a metal to a nonmetal. In a **covalent bond**, valence electrons are shared between atoms forming molecules. In this lab we will only consider covalent bonds.

The **Lewis structure** is a representation of how valence electrons are arranged among atoms in a molecule. It shows us how chemical bonds are used to hold together the atoms in a molecule. In drawing Lewis structures, we include only valence electrons, which are represented as dots. When a covalent bond forms, two electrons are shared, represented as a line. The rules for drawing Lewis structures are based on the **Octet Rule**, which states that the most important requirement for the formation of a stable compound is that the atoms achieve a noble gas electron configuration, or eight valence electrons.

Rules for drawing Lewis structures:

1. Determine the total number of valence electrons (outer shell electrons) for the compound. This corresponds to the "A" group the atom belongs on the periodic table, not the atomic number. For example, CO_2: carbon is in group IVA and has 4 valence electrons, oxygen is in group VIA and each has 6 valence electrons, giving a total of 16 valence electrons for the molecule.

2. Determine the central atom (the atom that is least electronegative).
   a. carbon or silicon are <u>always</u> in the middle.
   b. hydrogen is <u>never</u> in the middle.
   c. oxygen is usually not in the middle
   d. the halogens (F, Cl, Br, I) are usually not in the middle.

3. Draw the basic structure: the central atom connected to each of the surrounding atoms by a <u>single bond</u>. A single bond is a shared electron pair.

4. The octet rule for the outer atoms: Surround each outer atom with eight electron dots (four pairs of electrons). *Remember, the covalent bond counts as 2 electrons for each atom to which it is attached.
   *Remember, hydrogen can only have two valence electrons (no lone electrons for H)

5. The octet rule for the center atom:
   *Count the number of electrons; each line (covalent bond) counts as 2, each dot as 1.
   a) If there are any extra electrons from step 1, put these on the central atom.
   b) Does the center atom have an octet?
      i) Yes☺, this is your Lewis structure
      ii) No☹, you must make double or triple bonds. A double bond represents 4 shared electrons. When you draw a double bond, you must erase two electrons from the outer atom attached to the double bond and move it into the double bond. When you draw a triple bond, you must move 4 electrons from the outer atom.

**Hydrogen cannot make a double bond. Halogens do not make double bonds.*

Electron Pair Geometry and Bond Angle

Two systems are used to describe the three-dimensional geometry of molecules. The number of electron pairs surrounding the central atom determines its **electron pair geometry** (**EPG**) and **molecular geometry** (**MG**). All electrons, both shared and unshared, have a negative charge. Since like charges repel, electron pairs try to keep as far from each other as possible.

We will refer to electron pairs, both shared and unshared, as **electron clouds**. Each unshared pair of electrons surrounding the central atom occupies an electron cloud. In addition, each atom attached to the central atom (either by a single, double or triple bond) occupies an electron cloud.

In the EPG model, all electron clouds are positioned as far apart as possible to minimize electrostatic repulsion between negative charges. When there are two electron clouds surrounding the central atom, the furthest distance apart they can get is 180^o from each other that leads to a **linear** arrangement in space. When there are three electron clouds, the furthest distance apart is 120^o from each other, in a **triangular planar** arrangement in space. When there are four electron clouds, the furthest distance apart is about 109^o that leads to a **tetrahedral** arrangement of electron clouds.

Molecular Geometry

The molecular geometry describes the three dimensional **arrangement of atoms** around the central atom. We have just seen that based on the number of electron clouds surrounding the central atom we can determine the EPG geometry and bond angle. By further dividing the electron clouds (into the number of clouds that are unshared pairs and those that surround atoms) we can determine the molecular geometry (MG) of the compound. In determining EPG geometry, all electron clouds are considered. To determine the molecular geometry based on the EPG geometry, we consider only electron clouds that surround atoms.

Below is a chart summarizing electron geometry and molecular shape. Notice that if there are zero non-bonding/unshared electron pairs, the EPG geometry and molecular geometry are the same.

e- clouds (total)	**Atom clouds**	**LonePair Clouds**	**Electron Pair Geometry (EPG)**	**Bond Angle**	**Molecular Geometry (MG)**
2	2	0	Linear	180°	Linear
3	3	0	Trigonal Planar	120°	Trigonal Planar
3	2	1	Trigonal Planar	120°	Bent
4	4	0	Tetrahedral	109°	Tetrahedral
4	3	1	Tetrahedral	109°	Trigonal Pyramid
4	2	2	Tetrahedral	109°	Bent

Polarity: Molecules that have an imbalance or asymmetrical distribution of electrical charge are said to be **polar**. This imbalance of electrical charge is due to a combination of bond polarities (differences in electronegativities of the bonded atoms) and the shape of the molecule. If all bonds in the molecule are the same (i.e. the molecule is symmetrical), the molecule is **nonpolar (NP)**. For example, CCl_4 has polar C-Cl bonds, but is a non-polar molecule because the shape results in all the negative areas canceling each other. Thus, the molecule does not have a negative and positive end so it is a nonpolar molecule.

Chemical Bonding and Molecular Models

Name:
Lab Partner(s):

Procedure: Draw the Lewis structure and describe the Lewis structure for the following molecules. Complete the table below. *The table of EPG and MG terms on the previous page will be helpful while completing this worksheet.*

Part A: Overview of shapes

Molecule	# Val e^-	Draw the Lewis Structure	Electron Pair Geometry / Molecular Geometry	Bond Angle	Polar *or* NP
CH_4			*EPG:* *MG:*		
NF_3			*EPG:* *MG:*		
H_2O			*EPG:* *MG:*		
CO_2			*EPG:* *MG:*		
SO_3			*EPG:* *MG:*		
SO_2			*EPG:* *MG:*		

Part B: More Lewis Structures with one central atom

Molecule	# Val e^-	Draw the Lewis Structure	Electron Pair Geometry / Molecular Geometry	Bond Angle	Polar *or* NP
$TeCl_2$			*EPG:* *MG:*		
CF_2Cl_2			*EPG:* *MG:*		
SeO_2			*EPG:* *MG:*		
HCN			*EPG:* *MG:*		
PH_3			*EPG:* *MG:*		
CH_2O			*EPG:* *MG:*		
N_2O *N in middle*			*EPG:* *MG:*		
NO_2^{-1}			*EPG:* *MG:*		
CO_3^{2-}			*EPG:* *MG:*		