Chemical Bonding and Molecular Models

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 (often 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 surrounded by the remaining atoms.

4. Between the central atom and each surrounding atom, place two electrons or a line to designate a single covalent bond (i.e. C-H) composed of a shared electron pair.

5. The octet rule: Surround each atom with eight electron dots (four pairs of electrons).
   *Remember, the covalent bond counts as two electrons for each atom to which it is attached.
   *Remember, hydrogen can only have two valence electrons.

6. Count the number of electrons; each line (covalent bond) counts as two, each dot as one.
   a. If the number of electrons equals the valence electrons in step 1, this is the Lewis structure.
   b. If you have more electrons than in step one, you must replace one or more single bonds with multiple bonds. When you draw a double bond, you must erase two electrons from each atom attached to the double bond (erase four electrons total) to satisfy the octet rule.

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

Valence Shell Electron Pair Repulsion (VSEPR) Geometry and Bond Angle

Several models are used to describe the three-dimensional geometry of molecules. The number of electron pairs surrounding the central atom determines its electron geometry and molecular shape. 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 occupy 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 this lab, we will consider molecules in which the central atom is surrounded by two, three, or four electron clouds.

In the VSEPR 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 which 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** arrangement in space. When there are four electron clouds, the furthest distance apart is about 109^o. This leads to a **tetrahedral** (a four sided three dimensional shape) arrangement of electron clouds. [Note that the tetrahedron is more spacious than a square with angles of only 90^o.]

Molecular Shape

The molecular shape 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 VSEPR 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 shape of the compound. In determining VSEPR geometry, all electron clouds are considered. To determine the molecular shape based on the VSEPR 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 VSEPR geometry and molecular shape are the same.

e-clouds (total)	Atom clouds	LonePair Clouds	VSEPR Geometry	Bond Angle	Molecular Shape
2	2	0	Linear	180^o	Linear
3	3	0	Planar Triangle	120^o	Planar Triangle
3	2	1	Planar Triangle	120^o	Bent (angular)
4	4	0	Tetrahedral	109^o	Tetrahedral
4	3	1	Tetrahedral	109^o	Pyramidal
4	2	2	Tetrahedral	109^o	Bent (angular)

Polarity: Molecules may have a negative end and a positive end. If this is the case they are *polar* molecules. 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 non-polar, the molecule is always non-polar. If there are polar bonds present in the molecule, usually the molecule is polar. There are important exceptions to this generalization, however. For example CCl_4 has polar C-Cl bonds, but is a non-polar molecule because its shape results in all the negative areas canceling each other. Thus the molecule doesn't have a negative and positive end so it is a non-polar molecule.

Chemical Bonding and Molecular Models Name:

Construct three-dimensional models of various molecular compounds using a model kit. The balls represent atoms and the sticks represent covalent bonds or pairs of electrons. Two springs = a double bond. The different colored balls represent different types of atoms. (See front of box for details.)

PART A:

Molecule	Val e-`	Lewis Structure	VSEPRGeo/MolecShape	Angle	Polar?
CO_2					
SO_3					
SO_2					
CH_4					
NH_3					
H_2O					

Part B: More Lewis Structures with one central atom

Molecule	Val e-`	Lewis Structure	VSEPRGeo/MolecShape	Angle	Polar?
CF_2Cl_2					
HCN					
PBr_3					
CH_2O					
N_2O (N is the central atom.)					

PART C: Negative charged molecules: add one valence electron for every negative charge.
(Indicate geometry, shape, and angle.)

NO_3^-

CO_3^{2-}

SO_3^{2-}

PART D: More than one central atom. (Indicate geometry, shape, angle and polarity of the entire molecule.)

C_2H_6

C_2H_4

C_2H_2

CH_3OH (Indicate geometry and shape around carbon and also around oxygen)

$C_2H_4Cl_2$ (There are two possible structures, called **ISOMERS.** Draw and describe both structures)

Postlab Question:

In some chemistry courses, the instructor requires students to purchase a model kit. Suppose you were in such a course and decided to economize by making your kit out of different colored gumdrops and toothpicks. What problems might you have in using your kit?