**Chemistry 110**
**Spring 2011**
**Dr. Abrash**

# Experiment 6: Chemical Bonds, Molecular Models, and Molecular Shapes

## What is the purpose of this lab?

The purpose of this experiment is to understand some of the factors leading to the shapes and the bonding of some molecules that are either common in the atmosphere or are important in global warming.

## Why are the shapes and bonding of molecules important?

Molecular shapes and the nature of the chemical bonds holding the atoms together taken together are called molecular structure. Structure is important because it determines how stable a molecule is (i.e. how easy it is for it to fall apart), how reactive it is (how easily it reacts with another molecule) and how it absorbs light (important in global warming.)

## What is a chemical bond?

A chemical bond is the force that holds a pair of atoms together to make a molecule or ion. The force is usually generated by having one or more valence electrons shared by the two atoms which are held together. The strongest and stablest bonds come when one or more <u>pairs</u> of electrons are shared by the atoms being bonded.

## How can we tell which electrons are shared by two atoms?

The shared electrons are the ones that are found between the two atoms.

## What are valence electrons?

An atom's electrons are crudely arranged in shells. Each successive shell is a bit farther from the nucleus than the previous one. The outermost shell is called the valence shell. Valence electrons are electrons that are in the valence shell.

**Why is the valence shell important?**

The valence shell is important because the outermost electrons are the ones that are easiest for other atoms to reach, so they are the ones that are involved in bonding and in chemical reactions.

**Do all bonds have the same strength?**

Bonds have different strengths. The most important factor in determining the strength of a bond is the number of electron pairs shared by the bonding atoms. If there is a single pair of electrons between a pair of atoms, the bond is called a single bond, two pairs of electrons yield a double bond, and three pairs yield a triple bond. A triple bond is stronger than a double bond, which is stronger than a single bond. In the Chapman cycle, which you will learn about in lecture, absorption of UV light by ozone breaks the O-O single bond, not the stronger double bond.

**Do all bonds have the same tendency to react?**

No. Paradoxically, single bonds, which are the easiest to break, tend to be the least reactive, while multiple bonds tend to be the most reactive. One caution, though – you should not compare a single bond of one species with a multiple bond of another, because the atoms involved in a bond also affect its reactivity. I.e., saying that an O-O single bond is less reactive than a C≡C triple bond is not necessarily correct. The correct comparison would be between a C-C single bond and a C≡C triple bond.

**How can we predict whether a bond in a specific molecule is a single bond, a double bond or a triple bond?**

We create a diagram called a Lewis Dot Structure.

**What is a Lewis Dot Structure?**

A Lewis Dot Structure (Lewis structure for short) is a representation of the way that electrons are distributed around the atoms in a molecule. The atoms are represented by their chemical symbols, and the electrons by dots, most typically by pairs of dots. In a Lewis structure, there are two types of electrons, bonding electrons and electrons which are not involved in bonding and which are called unshared pairs or lone pairs.

A Lewis dot structure is extremely useful because it provides a convenient way to represent an important rule in chemistry called the octet rule.

**What is the octet rule?**

The octet rule says that to be stable, all atoms in a molecule must be surrounded by 8 electrons, either in lone pairs or bonds. Electrons in bonds are counted as part of the octet for both of the atoms in the bond.

Exceptions: Atoms in the first row (or period) of the periodic table need only two electrons instead of eight. In addition beginning with the third row many atoms can be surrounded by more than 8 electrons. This is called an expanded octet. However, this second exception is not necessary for any of the molecules in today's lab.

**How do you construct a Lewis dot structure? (Slightly modified from page 116 of your text)**

1) Determine the total number of outer electrons.
    a) Determine the number of outer electrons for each atom in the molecule.
        i) You do this by looking at the periodic table, and determining which column the atom is in. Each column in the periodic table is called a group. Atoms in the same group have the same number of outer electrons. For atoms in groups with labels ending in A (1A, 2A, 3A, etc), the number of outer electrons is equal to the column number.

           Examples: Hydrogen is in group 1A and has 1 outer electron. Nitrogen is in group 5A and has 5 outer eletrons.
    b) Add up the numbers of outer electrons for all atoms in the molecule.

2) Arrange the atoms around the central one (the unique one usually is the central atom, e.g. $\mathbf{C}H_4$).

3) Place one pair of electrons between the central atom and each of the atoms surrounding it. Use dots to represent the electrons. This is equivalent to creating single bonds between the central atom and the surrounding atoms. Subtract the electrons used in making these bonds from the total.

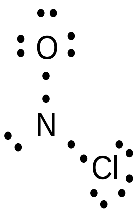
4) Distribute the remaining electrons around the OUTER atoms until each has a full octet. Remember that H, since it's in the first row, needs only two electrons, so the single bond fulfills its requirement for electrons.

5) Put any remaining electrons on the central atom

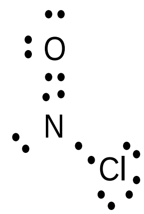
6) If the central atom does not have an octet, take a lone pair from one of the outer atoms, and place it between that atom and the central atom to make another bond. When two electron pairs are between two atoms it is called a double bond, and when three electron pairs are between two atoms it is called a triple bond.

Example

**ONCl** nitrosyl chloride



total outer electrons : 6 + 5 + 7 = 18
Use two electrons to make a bond between N and O
Use two more to make a bond between N and Cl
This leaves 14 electrons
Add three pairs to O to complete the octet
Add three pairs to Cl to complete the octet
This leaves two electrons. Put them on the N.



N doesn't have an octet. One of the atoms need to share a pair with N: which one???
If we make the double bond by sharing an extra pair between N and O
N gets its octet, and O ends up with the same number of electrons it started with (6)
So O is the proper choice for the double bond with N.
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Note that chemists will often draw a line to represent a pair of electrons.

**Once You Know the Lewis Dot Structure How Can You Figure Out the Shape?**

1) Make a model according to the directions in the lab manual

OR

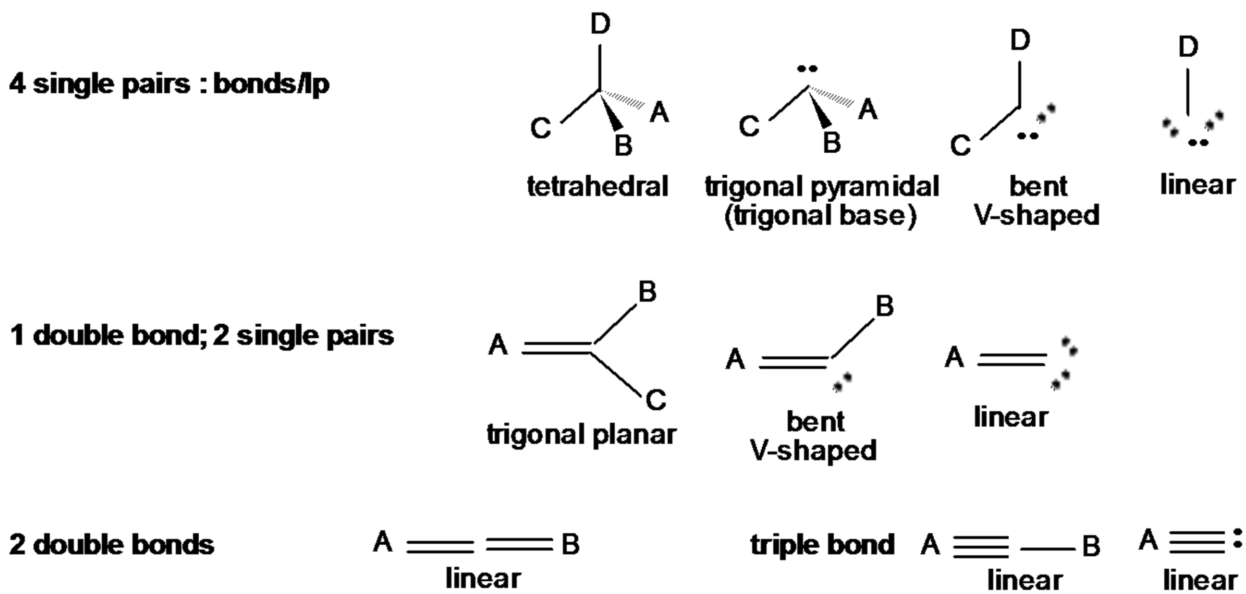
2) Figure out the shape from the Lewis Dot Structure. The principle we use is that pairs of electrons, whether they are in bonds or in lone pairs (or rarely, all by themselves) repel each other. (Remember that like charges repel.) Therefore the stablest structure will be the one that yields the most distance between electron pairs. As a result, we use the following rules to figure out the structure
    a. Create a Lewis structure
    b. Count the number of bonds and lone pairs surrounding the central atom (for this purpose a double or triple bond counts as one bond.
    c. assign the structure based on this number, which we'll call n
        i. if n=1 or n=2, the shape of the molecule is linear

ii. if n=3, the shape of the molecule is either trigonal planar, (if you have three bonds) or bent (if you have two bonds and a lone pair)

iii. if n=4, the shape of the molecule is called tetrahedral, if there are four bonds; pyramidal if there are three bonds and a lone pair; and bent if there are two bonds and two lone pairs.

Here is a list of all possible shapes, how they are drawn, and their names.

Possible geometries:




**What is the assignment for today's lab.**

For the molecules listed in your lab manual do the following:

1) Determine the number of outer electrons
2) With the rules above, determine the Lewis structure
3) By building the molecules with model kits, or using the rules above, determine the shapes
4) Fill in the table in your manual. Where it asks you for the structural formula in the last column, draw the molecule, as above, and give the proper name for the shape of the molecule.
5) Answer questions 1-7 at the end of the lab.