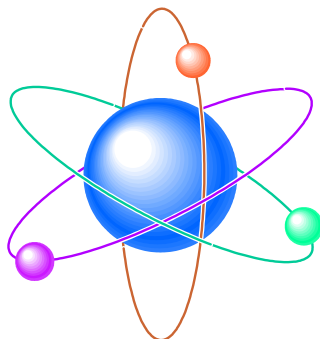


# BUILDING MOLECULES



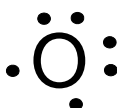
## INTRODUCTION

Atoms, the smallest (and invisible) free particles of an element, combine to form molecules. For example,  $O_2$ ,  $CO_2$ ,  $CH_4$  are all molecules. Atoms use electrons located in the outer shell in this process. These electrons are called valence electrons.

For "A" family elements, the number of valence electrons equals the Group number. For example, aluminum is found in group 3A, and therefore it has three valence electrons and its electron dot symbol is:



This procedure can be done for any "A" family element. There is a complicating feature, however. After four electrons have been added to the valence shell, the next electron(s) must be paired with one of the previous four electrons. For example, oxygen would look like:



Notice that oxygen now has two lone pairs (nonbonding pairs) of electrons and two electrons remain unpaired. These unpaired electrons remain available to form bonds (bonding pairs) with other atoms.

The elements found in column 8A have eight electrons in their valence shell and are particularly stable and unreactive toward other atoms. Atoms having fewer than eight electrons in the outer shell tend to react with other similar atoms to achieve the stability exhibited by the elements in column 8A (also known as the noble gases). The tendency for atoms to form bonds to acquire eight electrons is known as the "octet rule." The octet rule is only an approximation to aid in understanding the bonding and molecular shape of molecules.

The octet rule can be illustrated with hydrochloric acid. Atoms, such as hydrogen and chlorine (each having one unpaired electron), combine their unpaired electrons to form a single bonding pair. The chlorine atom appears to have eight electrons (3 lone pairs, 1 bonding pair), while the hydrogen appears to have only two (1 bonding pair).

**NOTE: Hydrogen and helium are exceptions to the octet rule; they can only have two electrons in their valence shell.**

Oxygen, having six electrons in its valence shell, can form two single bonds or one double bond. As you work through the lab, you will see examples of this. Nitrogen has the potential to form three single bonds, one double bond and one single bond, or one triple bond.

It is possible for an atom with an odd number of electrons to form stable molecules with atom(s) containing an even number of electrons. Such molecules violate the octet rule. Examples of this phenomenon are in Section V. Another apparent violation of the octet rule occurs when the shape of a molecule in space can't be represented by a single figure on paper. Several configurations of electrons can be drawn that satisfy the octet rule. To properly represent the molecule, all of these must be drawn. This is called **resonance** and each drawn representation is called a **resonance structure**. An example of this is the ozone molecule (O<sub>3</sub>).

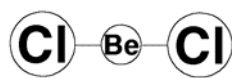
Bonding pairs and lone pairs influence the shape and properties of molecules. The shape of the molecule is described by the position of the atoms in space. Atoms that are able to form the most bonds are generally located in the center of the molecule, or form the "backbone" of the molecule. Lone pairs must be considered when determining the shape of molecules in space. (They seem to take up more space than bonding pairs). Table I describes the three-dimensional arrangement of several molecules. This table will be used in assigning three-dimensional shapes to some of the molecules that you will be working with today.

Table I. Bonding and the Shape of Molecules

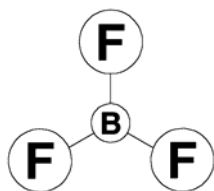
Number of Bonded Atoms	Number of NBP*	Sum of Bonded Atoms and NBP's	Molecular Shape	Examples
2	0	2	Linear	BeCl <sub>2</sub> HgCl <sub>2</sub> CO <sub>2</sub> HCN
3	0	3	Triangular	BF <sub>3</sub> AlBr <sub>3</sub> CH <sub>2</sub> O
4	0	4	Tetrahedral	CH <sub>4</sub> CBr <sub>4</sub> SiCl <sub>4</sub>
3	1	4	Pyramidal	NH <sub>3</sub> PCl <sub>3</sub>
2	2	4	Bent	H <sub>2</sub> O H <sub>2</sub> S SCl <sub>2</sub>
2	1	3	Bent	SO <sub>2</sub> O <sub>3</sub>

NBP\* = nonbonding pair of electrons.

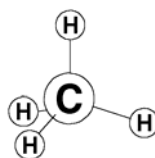
Chemistry for Changing Times - John W. Hill, 6th Edition.



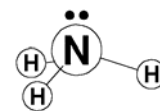
Linear ( $\text{BeCl}_2$ )



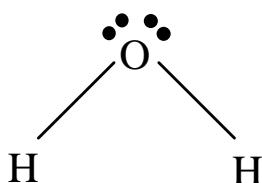
Triangular ( $\text{BF}_3$ )



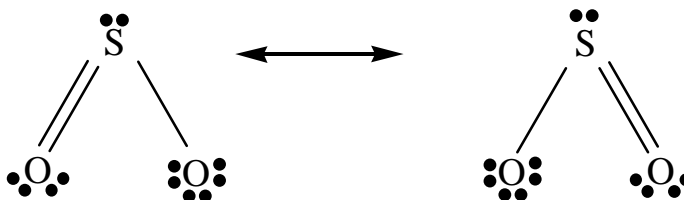
Tetrahedral ( $\text{CH}_4$ )



Pyramidal ( $\text{NH}_3$ )



Bent ( $\text{H}_2\text{O}$ )



Bent ( $\text{SO}_2$ )

## EXPERIMENTAL PROCEDURE

### I. Valence Electrons and Electron Dot Symbols for the Elements

Give the electron dot symbol for each of the following elements and specify how many valence electrons each contains.

Hydrogen (H)

Helium (He)

Lithium (Li)

Beryllium (Be)

Boron (B)

Carbon (C)

Nitrogen (N)

Oxygen (O)

Fluorine (F)

Neon (Ne)

Sodium (Na)

## II. Single Bonds

Draw Lewis dot structures and (if possible) indicate the molecular shape of the following molecules. (Table I will help you with this). After you have completed these, construct the following molecules using the Prentice-Hall Organic Model Kit.

Key to colors in the Organic Molecular Model Kit from Prentice Hall

carbon - black	fluorine or iodine - purple	hydrogen - white
chlorine - green	oxygen - red	sulfur - blue
nitrogen - blue	bromine - orange	

Use the short, gray connectors for single bonds. Use the long, gray connectors to construct double and triple bonds. Use the short, white connectors for lone pairs.

### A. Compounds of hydrogen

1. Methane (CH<sub>4</sub>)
2. Water (H<sub>2</sub>O)
3. Ammonia (NH<sub>3</sub>)
4. Hydrogen sulfide (H<sub>2</sub>S)
5. Hydrogen fluoride (HF)

B. Compounds of carbon:

1. Tetrachloromethane (Carbon tetrachloride)  $\text{CCl}_4$

2. Fluorotrichloromethane ( $\text{CFCl}_3$ )

3. Dichlorodifluoromethane ( $\text{CF}_2\text{Cl}_2$ )

4. HCFC-22 ( $\text{CHF}_2\text{Cl}$ )

5. HFC 134a ( $\text{CF}_3\text{CH}_2\text{F}$ )

6. Propane ( $\text{C}_3\text{H}_8$ )

III. Double Bonds

Draw Lewis dot structures and (if possible) indicate the molecular shape of the following molecules (Table I will help you with this.). After you have completed these, construct the molecule.

A. Carbon Dioxide ( $\text{CO}_2$ )

B. Oxygen ( $\text{O}_2$ )

C. Teflon ( $C_2F_4$ )

$$\begin{array}{c} O \\ // \\ \text{D. Formaldehyde (HCH)} \end{array}$$

#### IV. Triple Bonds

Draw Lewis dot structures and (if possible) indicate the molecular shape of the following molecules (Table I will help you with this). After you have completed these, construct the molecule.

A. Nitrogen ( $N_2$ )

B. Hydrogen cyanide (HCN)

C. Carbon Monoxide ( $CO$ )\*

D. Acetylene ( $C_2H_2$ )

\*For this model ONLY, use BLUE to represent oxygen.

#### V. Violation of the Octet Rule and Resonance

Draw Lewis dot structures and (if possible) indicate the molecular shape of the following molecules (Table I will help you with this). After you have completed these, construct the molecule.

A. Hydrogen ( $H_2$ )

B. Ozone ( $O_3$ )

C. Nitrogen oxide ( $NO$ )

D. Nitrogen dioxide ( $NO_2$ )

E. Sulfur dioxide ( $SO_2$ )

## REVIEW QUESTIONS

### BUILDING MOLECULES

1. When an atom is bonded, how many electrons should it have in its valence shell? What elements prove to be exceptions to this general rule? Why?
2. How do nonbonding pairs of electrons influence the shape of the molecule?
3. How does a double or triple bond affect the shape of a molecule?
4. Draw Lewis dot structures for the following atoms:
  - a. carbon
  - b. magnesium
  - c. sulfur
5. Draw the Lewis structure for a nitrogen-containing molecule. Try  $\text{NO}_2$ . What problem arises? Can resonance solve it?
6. Explore the concept of resonance by making the following molecules and anions obey the octet rule
  - a. sulfur dioxide ( $\text{SO}_2$ )
  - b. sulfur trioxide ( $\text{SO}_3$ )
  - c. carbonate ( $\text{CO}_3^{2-}$ )
  - d. nitrate ( $\text{NO}_3^-$ )