Experiment 5
Can You Model This?

OUTCOMES

After completing this experiment, the student should be able to:

- Differentiate between molecular compounds and ionic compounds.
- Construct Lewis-dot structures and three-dimensional models of molecular compounds.

DISCUSSION

A chemical bond is a force of attraction that holds atoms together in compounds. Bonds form to attain a more stable arrangement of valence electrons. Chemical bonds may be either ionic bonds or covalent bonds. An ionic bond results from the transfer of valence electrons from *metal* atoms to *nonmetal* atoms and the subsequent attraction of oppositely charged particles, while the covalent bond results from the sharing of valence electrons between two *nonmetal* atoms.

By the time you get to this experiment, you should have already discussed chemical bonding in general and ionic bonding. You should be familiar with the *octet rule*. In this experiment, we will only consider *molecular* compounds for making models and drawing Lewis structures. A molecule is group of atoms (usually only nonmetals) held together by covalent bonds. An atom is the smallest neutral particle of an element. An ion is a charged atom or group of atoms, formed by addition or removal of one or more valence electrons. If any of these terms are unfamiliar, please read about them in your Chemistry 1020 textbook before this experiment.

One of the disadvantages of constructing Lewis Structures (also known as electron-dot models) on paper is that they do not always convey information about the three-dimensional geometry of a molecule. The geometry of a molecule provides important information about the physical and chemical behavior of a molecule and the compound as a whole. Using the models, you will examine the three-dimensional structure of the carbon atoms in many of the molecules in this experiment.

The procedure for this experiment is rather informal and it has been assumed that your instructor will help lead you through the early stages. Your lecture instructor may elect to relate this experiment to geometry and polarity later in the course, so hold on to this experiment.

PROCEDURE

Obtain a molecular model kit and examine the pieces inside. The kit should contain different colored balls that have holes. The balls are color-coded to represent different elements. The color-coded scheme is as follows:

white or yellow ball	hydrogen	1 hole
black ball	carbon	4 holes
red ball	oxygen	2 holes
blue ball	nitrogen	3, 4, or 5 holes (only fill 3 in this experiment)
green ball	chlorine	1 hole
orange ball	bromine	1 hole
purple ball	iodine	1 hole

Each ball represents an atom, while each hole represents a missing valence electron needed to complete a stable octet. **The number of holes indicates the number of covalent bonds the atom will form in a molecule** because each covalent bond gives the atom access to one additional electron. When the model for a molecule is constructed, *all of the holes will be filled with connectors*. With these model kits, nitrogen can be an exception. It may be left with one or two empty holes, since nitrogen will only form three covalent bonds.

If two balls are joined by a single connector, the connector represents a single covalent bond comprised of two shared electrons. If two balls are joined by two connectors, the connectors represent a double bond and a total of four shared electrons. A triple bond and a total of six shared electrons is represented by two balls joined by three connectors.

To summarize, covalent bonds are represented in your models by the following connectors:

To represent:	Wood kit	Plastic kit	Represents
Single C-H bond	Short wood stick	Short plastic stick	one electron pair
All other single bonds	Long wood stick	Short plastic stick	one electron pair
Double bond	Two springs	Two long, flexible plastic connectors	two electron pairs
Triple bond	Three springs	Three long, flexible plastic connectors	three electron pairs

1. Construct models of the compounds whose formulas are given below and in the data table.

2. From the formula, add up the total number of valence electrons in the compound. Then, draw Lewis structures for each compound. Use lines to represent covalent bonds (shared electron pairs) and dots to represent unshared electron pairs.

3. For simple neutral molecules like those in this experiment, three simple checks will ensure that your Lewis structure is correct. For each diagram, place a check mark in the appropriate box in the rightmost columns once you've checked the following:

   - **The correct total number of valence electrons are shown in your completed Lewis structure.** Count the number of valence electrons shown in your diagram (remembering that each covalent bond is a pair of electrons). This should match your calculated total. (This rule applies to *all* electron-dot diagrams.)

   - **The octet rule or duet rule is followed in every atom in your Lewis structure.** You should be able to count 8 electrons (including bonds) around all atoms (except hydrogen, which should have 2). (Note, your instructor may introduce you to some molecules that do not obey the octet/duet rule, but the rule *does apply* to all molecules in *this experiment*.)

   - **Every atom in your Lewis structure has the correct number of bonds.** To make the best Lewis structure, atoms should have only the number of bonds necessary to fill their octet. For instance, chlorine has 7 of its own valence electrons, so it should only form 1 covalent bond. (Note, if you draw diagrams of polyatomic ions in lecture, this rule does not necessarily apply to them. This rule applies to simple *neutral* molecules like those in this experiment.)

4.1. Molecules of these substances have only **single bonds**.

(a) H_2
(b) H_2O
(c) HCl
(d) Br_2
(e) NH_3
(f) $HOCl$
(g) I_2
(h) IBr
(i) Cl_2
(j) CH_2Cl_2 *
(k) CH_4 *
(l) H_2O_2
(m) CCl_4 *
(n) CH_3NH_2 *

4.2. Molecules of these substances have **double bonds**. There may be more than one double bond present.

(o) O_2
(p) CH_2O *
(q) C_2H_4 *
(r) CO_2 *
(s) $HCOOH$ * (Be careful! The H's are bonded to different atoms.)

4.3. Molecules of these substances have **triple bonds**.

(t) N_2 (u) C_2H_2 * (v) HCN *

4.4. The following compounds have **isomers**. Isomers are compounds that have the same molecular formula, but different molecular structures. Since your experience with three-dimensional molecular models is limited, it is important that you build models of the following compounds to confirm your answers. Build as many isomers of each of the following compounds as possible.

(w) C_4H_{10} (2 isomers possible. Give both.)
(x) $C_2H_2Cl_2$ * (3 isomers possible. Give all three.)

5. *Optionally, your instructor may ask you to investigate the geometry of the carbon atoms in certain molecules in this experiment. For the starred (*) compounds, determine the geometry of carbon in the compound:

a. Carbon has a **tetrahedral geometry** when it is connected to four atoms, all by single bonds.

b. Carbon has a **trigonal planar geometry** when it is connected to three atoms, one by a double bond and two by single bonds (this is the only way carbon can be attached to three atoms but still contain a total of four covalent bonds).

c. Carbon has a **linear geometry** when it is connected to two atoms, either with two double bonds or one single and one triple bond.

As you are building the models, look at the position of the atoms connected to the carbon atoms in the starred compounds and be sure you understand the differences between these three geometries. Your instructor may ask you to draw your Lewis structures with the correct carbon geometry shown.

Name____ Lab Section________________

PRELAB QUESTIONS

1. Complete the following table:

Element:	H	C	N	O	F, Cl, Br, or I
Group number:					
Number of valence electrons:					
Normal number of covalent bonds:					

2. In your own words, what is the octet rule?

3. In your own words, what are some differences between atoms, ions, and molecules?

4. What is the difference between covalent and ionic bonds?

(more on back)

5. How many valence electrons does S have? How many valence electrons does S^{2-} have?

6. Give an example of an ionic compound.

7. Give an example of a molecular compound.

Name____ Lab Section______________

Partner's Name__________________________

Formula	Sum of valence e^-	Lewis Structures	All valence e^- shown?	All atoms follow octet or duet rule?	Correct # bonds per atom?
(a) H_2			❑	❑	❑
(b) H_2O			❑	❑	❑
(c) HCl			❑	❑	❑
(d) Br_2			❑	❑	❑
(e) NH_3			❑	❑	❑
(f) HOCl			❑	❑	❑
(g) I_2			❑	❑	❑
(h) IBr			❑	❑	❑

Name____ Lab Section______________

Formula	Sum of valence e^-	LewisStructures	All valence e^- shown?	All atoms follow octet or duet rule?	Correct # bonds per atom?
(i) Cl_2			❑	❑	❑
(j) CH_2Cl_2*			❑	❑	❑
(k) CH_4*			❑	❑	❑
(l) H_2O_2			❑	❑	❑
(m) CCl_4*			❑	❑	❑
(n) CH_3NH_2*			❑	❑	❑
(o) O_2			❑	❑	❑
(p) CH_2O*			❑	❑	❑

Name_______________________________________ Lab Section_____________

Formula	Sum of valence e^-	Lewis Structures	All valence e^- shown?	All atoms follow octet or duet rule?	Correct # bonds per atom?
(q) C_2H_4*			❑	❑	❑
(r) CO_2*			❑	❑	❑
(s) HCOOH*			❑	❑	❑
(t) N_2			❑	❑	❑
(u) C_2H_2*			❑	❑	❑
(v) HCN*			❑	❑	❑
(w) C_4H_{10}			❑	❑	❑
			❑	❑	❑

Name____ Lab Section_______________

Formula	Sum of valence e^-	Lewis Structures	All valence e^- shown?	All atoms follow octet or duet rule?	Correct # bonds per atom?
(x) $C_2H_2Cl_2$*			❑	❑	❑
			❑	❑	❑
			❑	❑	❑

POSTLAB QUESTIONS

1. Fill in the table below:

Atom or ion	Number of valence electrons
O	
O^{2-}	
N	
N^{3-}	
K	
K^+	

Name____ Lab Section_______________

2. Give an example when oxygen is covalently bonded to another atom. Give an example when oxygen is ionically bonded. Name both of your examples.

3. Why can't the sodium atom ever form a covalent bond with other metals?

(The following questions may be optional, *ask your instructor.*)

4. Which of the starred* compounds in this experiment has a **tetrahedral** carbon atom? Write the formula for each of the compounds. (*See page 5.4 for a description of carbon's geometries.*)

5. Which of the starred* compounds in this experiment has a **trigonal planar** carbon atom? Write the formula for each of the compounds.

6. Which of the starred* compounds in this experiment has a **linear** carbon atom? Write the formula for each of the compounds.