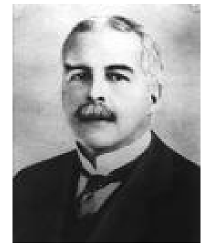
I'M TOO SEXY FOR THIS LAB: A MOLECULAR (SUPER)MODEL!



QUESTIONS

What do molecules look like? Why does it matter what they look like?

BACKGROUND

As you know, molecules—and the atoms that make them up—are much too small for us to see. Therefore, in order to study them, chemists have developed models that show the basic structure of atoms. From these models they are able to determine how atoms combine to form molecules. Understanding not only how molecules form but also their structure is important to understanding why substances have the particular properties they do.

One atomic model you have studied is the Bohr Model, which has the protons and neutrons in the nucleus and the electrons orbiting the nucleus in different energy levels. You are also familiar with the quantum model of the atom, where the protons and neutrons are still in the nucleus but the electrons are not orbiting the nucleus, but rather found in orbitals of varying energies outside of the nucleus. In this lab you will be studying a model for representing valance electrons, called a **Lewis structure** after its inventor, the American chemist Gilbert Newton Lewis. This model, sometimes also called a dot structure because it uses dots to represent electrons, focuses on an atom's valence electrons only and is particularly useful for studying the basic structure of molecules.

BACKGROUND QUESTIONS

1. What are valence electrons?

2. What is the importance of an octet? To which element does the octet rule not apply?

3. Why do you suppose Lewis structures only show valence electrons?

4. There are two ways atoms can bond with each other. Describe both of them.

5. Which of these two types of bonds do atoms use to make a molecule?

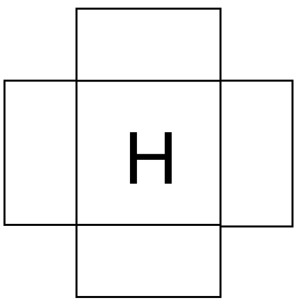
6. When studying the quantum model of an atom, you were given a rule called Hund's Rule. Explain what this is. (You may use your book if you need to!)

PROCEDURE

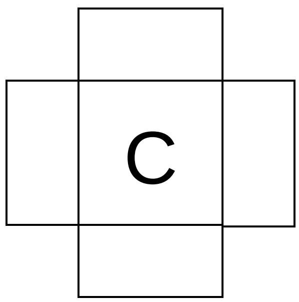
Part 1: Building Lewis Dot Models of Atoms

1. By each of the symbols written below, determine the number of valence electrons an atom of that element would have. Write the number in the blanks provided.
2. Using the Cheerios® as electrons, place the correct amount of valence electrons around each element's symbol. Use the boxes to help you with proper electron placement. The boxes represent the *s* and *p* orbitals of a valence energy level. Place the atom's first valence electron on the right side of the symbol, the second below the symbol, the third to the left, and so on moving clockwise around the symbol until all electrons have been placed. Some of your boxes might have 2 electrons in them.

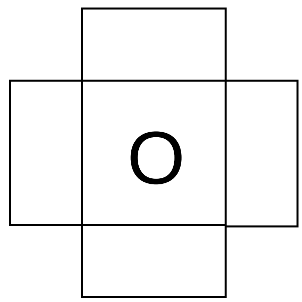
Number of valence electrons: _______



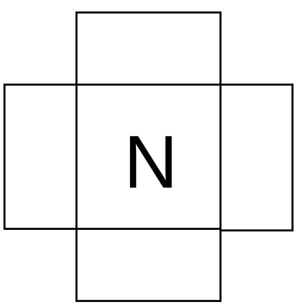
Number of valence electrons: _______



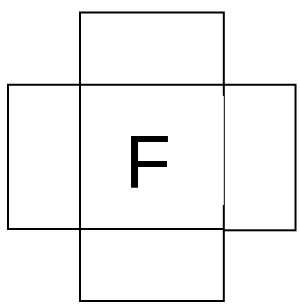
Number of valence electrons: _______



Number of valence electrons: _______



Number of valence electrons: _______



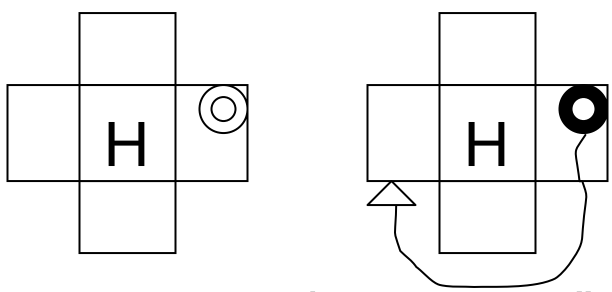
3. Check with your teacher before proceeding. __________
4. When your Lewis structures have been checked by your teacher, glue the Cheerios® in place.

Part 2: Building Lewis Structures for Molecules

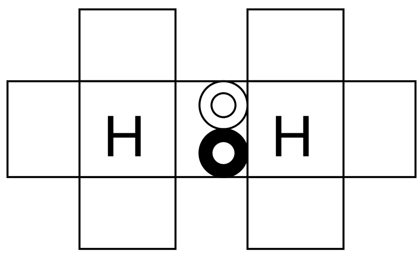
1. Your Lewis structures of atoms can be used to build molecules. The electrons that are in pairs are called **lone pairs**, or **nonbonding electrons**. In other words, these electrons are not usually used to bond because the orbital they are in already has its full complement of electrons. The unpaired electrons in your Lewis structures are known as **bonding electrons**. Each bonding electrons is capable of making one **single covalent bond**. If an atom has more than one bonding electron, it is capable of making multiple covalent bonds. To make a **double covalent bond**, an atom needs to have 2 bonding electrons. To make a **triple covalent bond**, an atom needs three bonding electrons. Fill in the table below:

Element	**Number of Lone Pairs**	**Number of Bonding Electrons**	**Number of Single Covalent Bonds the atom can form**	**Number of Double Covalent Bonds the atom can form**	**Number of Triple Covalent Bonds the atom can form**
H					
C					
O					
N					
F					

2. Remember that valence electrons are more stable when they are paired up. In order for bonding electrons to form a pair they must pair up with electrons from another atom. When this pairing of electrons from different atoms occurs, the atoms are said to be sharing their valence electrons and they have formed a covalent bond. Look at the diagrams below:
   a. Hydrogen gas is a molecule composed of 2 hydrogen atoms. According to the table you filled in, each H has one bonding electron:

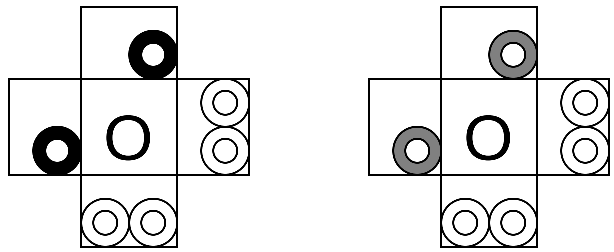


   b. Remember, atoms are not stationary but can rotate, so really the one electron can be in any of the 4 orbitals! If the H atoms get close enough so that the orbitals containing bonding valence electrons can overlap, the valence electrons can pair up:

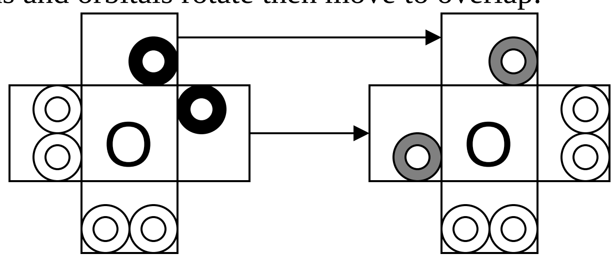


A single covalent bond has now been formed.

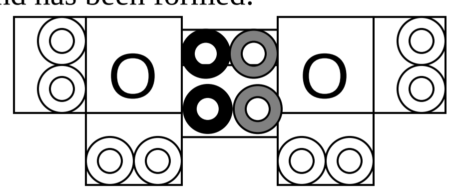
3. How might a molecule of oxygen gas, O_2, be made?
    a. Look at the table above; each O has 6 valence electrons, 2 of which are bonding electrons:



    b. The electrons and orbitals rotate then move to overlap:

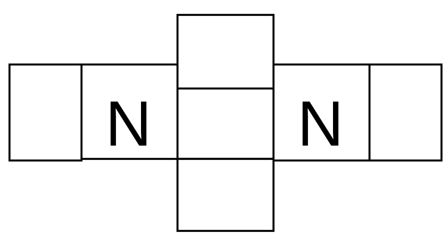


    c. As each O atom shares two of its bonding electrons with the other, a double covalent bond has been formed:



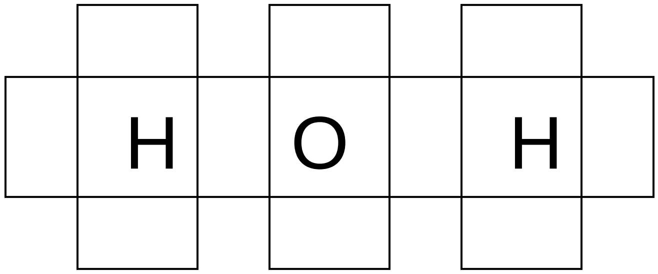
If you count the 4 shared electrons as belonging to BOTH atoms, each O atom now has an **octet**!

4. According to the table, each N has ___ valance electrons and ___ bonding electrons. Therefore it is capable of making a _______ bond. This would mean a molecule of nitrogen gas, N_2, would have a total of 10 electrons. Take 10 Cheerios® and build a molecule of nitrogen:



Does each N atom have an octet?

5. Now try building a more complex molecule: H_2O.
    a. Refer back to the table. Your Lewis structure should contain a total of ______ electrons: ______ from each H and ______ from O.
    b. H forms only __________ bonds. Oxygen is able to form a total of _______ single bonds.
    c. Using 8 Cheerios®, arrange them to show the two single covalent bonds in the water molecule. Use the boxes (next page) to help you.



6. When your Lewis structure for water has been checked by your teacher, attach the electrons to the paper. ________

7. Lewis models can be used to determine formulas of compounds you may not know. For example, how will carbon and hydrogen bond? For each pair of elements below, determine what molecule they will make by using Cheerios® to build the Lewis structure. Glue the Lewis structures to the paper and then write the molecular formula of each molecule. (Hints: Refer to your table! How many total electrons does each atom have? How many bonding electrons? How many single bonds can each make? Does each atom at the end (except H) have an octet? Are all electrons in pairs at the end?)

    a. C and H

    b. H and F

    c. H and N

    d. C and O (Hint: don’t forget about double bonds!)

ANALYSIS QUESTIONS

1. How can Lewis structures help us understand and predict bonding?

2. Will hydrogen ever make a double bond? Explain.

3. Explain why nitrogen doesn't make 5 covalent bonds, even though it has 5 valence electrons.

4. Using dots instead of stick-ums, draw the Lewis structures for the following elements:
   a. Si
   b. P
   c. S
   d. Cl

5. How many single covalent bonds can Si make? P? S? Cl?

6. Look back at the Lewis models you made earlier. Explain how you could use the periodic table to predict how many covalent bonds an element can make.

7. Lewis structures are used to determine **structural formulas** of molecules. Look at the structural formula of water below. What does each line in the drawing represent? The lone pairs are missing! Draw them in.

H-O-H

8. Look at the structural formula of carbon dioxide. What do the two lines between the C and O represent? The lone pairs are missing! Draw them in.

O=C=O

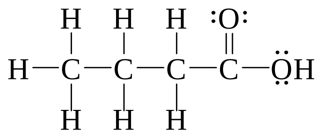
9. Predict the structural formula of silicon dioxide.

10. Predict the molecular formula of the compound that phosphorus and chlorine will make.

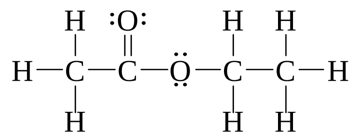
Name ____

GOING FURTHER

1. The Lewis structures for two compounds are shown below. How are the molecules below similar? How are they different?



Molecule #1
Butyric acid



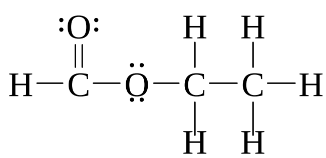
Molecule #2
Ethyl acetate

2. Molecule #1 smells like nasty dirty feet. Molecule #2 has a sweet, fruity flavor. What could account for the difference?

3. Explain how Lewis structures help a person determine the molecular formula of a compound.

4. Explain why knowing the Lewis structure of a molecule is more helpful than just knowing the molecular formula.

5. Look at the Lewis structure below. Would you expect it to smell smell sweet and fruity or nasty? Why?



6. Smell is just one property that depends upon what a molecule looks like. Another property is reactivity. Based on what you have learned about Lewis dot models and structures, give an explanation as to why ozone is more reactive than oxygen.

Oxygen Ozone