

# Oakland Schools Chemistry Resource Unit

## **Bonding**

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## **Bonding**

### **Content Statements:**

#### *C3.2x Enthalpy*

*Chemical reactions involve breaking bonds in reactants (endothermic) and forming new bonds in the products (exothermic). The enthalpy change for a chemical reaction will depend on the relative strengths of the bonds in the reactants and products.*

#### *C4.4x Molecular Polarity*

*The forces between molecules depend on the net polarity of the molecule as determined by shape of the molecule and the polarity of the bonds.*

#### *C5.5x Chemical Bonds - Trends*

*An atom's electron configuration, particularly of the outermost electrons, determines how the atom can interact with other atoms. The interactions between atoms that hold them together in molecules or between oppositely charged ions are called chemical bonds.*

#### *C5.5x Chemical Bonds*

*Chemical bonds can be classified as ionic, covalent, and metallic. The properties of a compound depend on the types of bonds holding the atoms together.*

#### *C4.3x Solids*

*Solids can be classified as metallic, ionic, covalent, or network covalent. These different types of solids have different properties that depend on the particles and forces found in the solid.*

### **Content Expectations:**

C3.2b - Describe the relative strength of single, double, and triple covalent bonds between nitrogen atoms.

C4.4b - Identify if a molecule is polar or nonpolar given a structural formula for the compound.

C5.5A - Predict if the bonding between two atoms of different elements will be primarily ionic or covalent.

C5.5c - Draw Lewis structures for simple compounds.

C5.5d - Compare the relative melting point, electrical and thermal conductivity, and hardness for ionic, metallic, and covalent compounds.

C5.5e - Relate the melting point, hardness, and electrical and thermal conductivity of a substance to its structure.

C4.3e - Predict whether the forces of attraction in a solid are primarily metallic, covalent, network covalent, or ionic based upon the elements' location on the periodic table.

## Background Information:

An easy way to determine whether a compound will be ionic or covalent is the electronegativity of its substituent elements. Electronegativity is a measure of how much an element wants to pull electrons away from an element that it has bonded to. Chemical compounds in which the two elements have very different electronegativities are ionic, because the element with a higher electronegativity pulls valence electrons completely off the less electronegative element. Chemical compounds in which the two elements have similar electronegativities are covalent, because neither atom gives up its electrons.

**COVALENT BONDING:** Valence electrons are equally shared between the bonded atoms of nonmetallic elements.

**POLAR COVALENT BONDING:** Electrons are shared but NOT equally between atoms of nonmetallic elements. Many compounds have the characteristics of BOTH ionic and covalent bonding. Electronegativity differences determine the balance of character. Example: water, carbon dioxide

In **NONPOLAR COVALENT BONDING:** Electrons are shared equally between two atoms that are exactly the same from nonmetallic elements. Example: hydrogen gas

**IONIC BONDING:** Valence electrons are completely transferred from one atom to the other atom of metallic elements. Ionic bonds occur between metals and nonmetals when there is a large difference in electronegativity. Example: salts

**METALLIC BONDING:** Valence electrons are shared among all of the atoms of metallic elements. Metallic bonding occurs when metals bond to either themselves or mixed with other metals in alloys. Example: iron, magnesium

Another way to determine the type of bonding between elements is by the location of the elements on the periodic table. Looking at the formula of the substance, if the elements are a mix of metals and nonmetals it is an ionic compound. A compound containing only nonmetals is covalent, and only metals is metallic.

The reason why atoms bond can be explained by the octet rule, which states that atoms will acquire eight valence electrons (the outermost electrons) to become stable. When atoms become more stable, they lower in energy.

## Properties of Ionic, Covalent, and Metallic Compounds

Ionic Compounds	Covalent Compounds	Metallic Compounds
-Formed from a combination of metals and nonmetals. -Electron transfer from the cation to the anion. -Opposite charged ions attract each other.	-Formed from a combination of nonmetals. -Electron sharing between atoms.	-Formed from a combination of metals -"sea of electrons"; electrons can move among atoms
Solids at room temperature	Can be solid, liquid, or gas at room temperature.	Solids at room temperature
High melting points	Low melting points	Various melting points
Dissolve well in water	Do not dissolve in water	Do not dissolve in water. (Sugar is an exception)
Conduct electricity only when dissolved in water; electrolytes	Do not conduct electricity; nonelectrolytes	Conduct electricity in solid form.
Brittle, hard	Soft	Metallic compounds range in hardness. Group 1 and 2 metals are soft; transition metals are hard. Metals are malleable, ductile, and have luster.

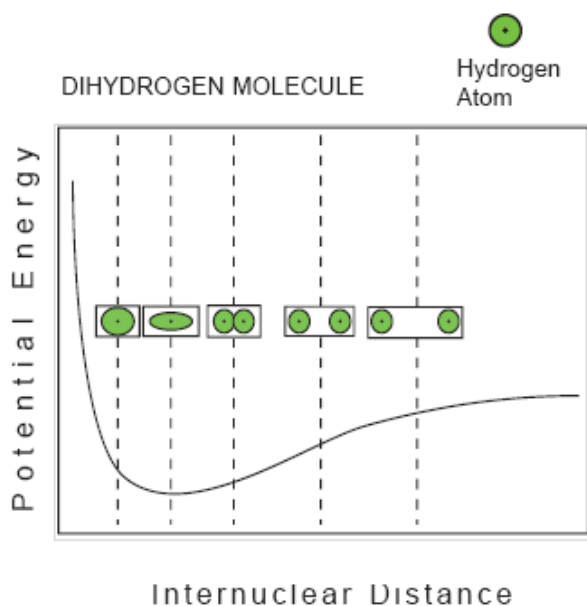
\*MOST compounds are a mixture between ionic and covalent

## Drawing Lewis Structures for Covalent Compounds

1. Determine the number of valence electrons available in the atoms to be combined. (Number of atoms of each element • number of valence electrons)
2. Arrange the atoms to form a skeleton structure for the molecule. If carbon is present, it is the central atom. Otherwise, the least electronegative atom is central (except for hydrogen, which is never central). Connect the atoms by bonds. Remember that a dash represents two electrons.
3. Add remaining electrons as unshared pairs of electrons to each nonmetal atom (except hydrogen) so that each is surrounded by eight electrons.
4. Count the electrons in the Lewis structure to be sure that the number of valence electrons drawn in the structure equals the number of available valence electrons.
  - a. Check to see that each atom has eight valence electrons. If not, move lone pairs in order to make double or triple bonds between non-hydrogen atoms until the outer shells of all atoms are completely filled.

## Changes in Potential Energy in the Bonding Process

### COVALENT H<sub>2</sub>



Confining electrons makes them 'dance'. This is part of quantum theory. The tighter electrons are squeezed the harder they dance. Dancing electrons have kinetic energy. But electrons will slow down if they can. When they have more room they can slow down, which means they have less kinetic energy.

In a hydrogen molecule the electrons can move through the space of two atoms instead of one, which means that they have more room and thus can dance slower and have less kinetic energy. (Picture 1)

There are also electric attractions and repulsions between the particles in the molecule. Picture 2 shows the repulsions of like charges as colored arrows and the attractions of opposite charges as black arrows.

The additive combination of the electric and kinetic energy effects gives the covalent bond for hydrogen.

Suppose you have two well separated hydrogen atoms and begin moving them closer together. From the picture above you can see that the energy of the system will decline as they are being moved together until at some distance the system will have a minimum energy.

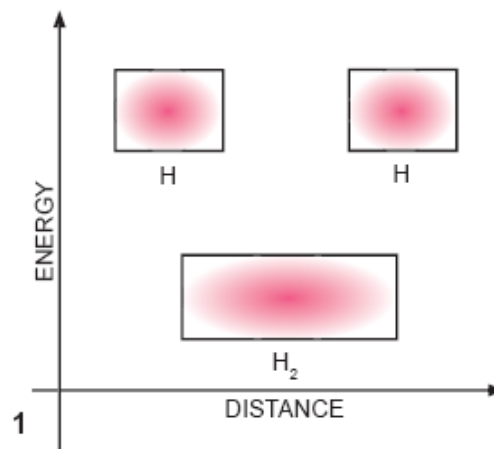
What causes energy to rise as the atoms are moved closer than the minimum energy?

A dihydrogen molecule consists of two hydrogen nuclei (protons) held a fixed distance apart and surrounded by a probability density cloud of two electrons.

As you can see from the picture above, the separation is that at which the system is in the state of lowest energy. But what are the factors which cause this to be a low energy state?

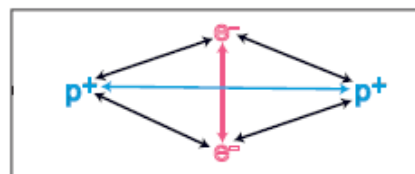
There are primarily two factors. They are quantum and electrostatic effects.

Quantum theory produces two effects, lowered energy and discrete energy levels.



1

2

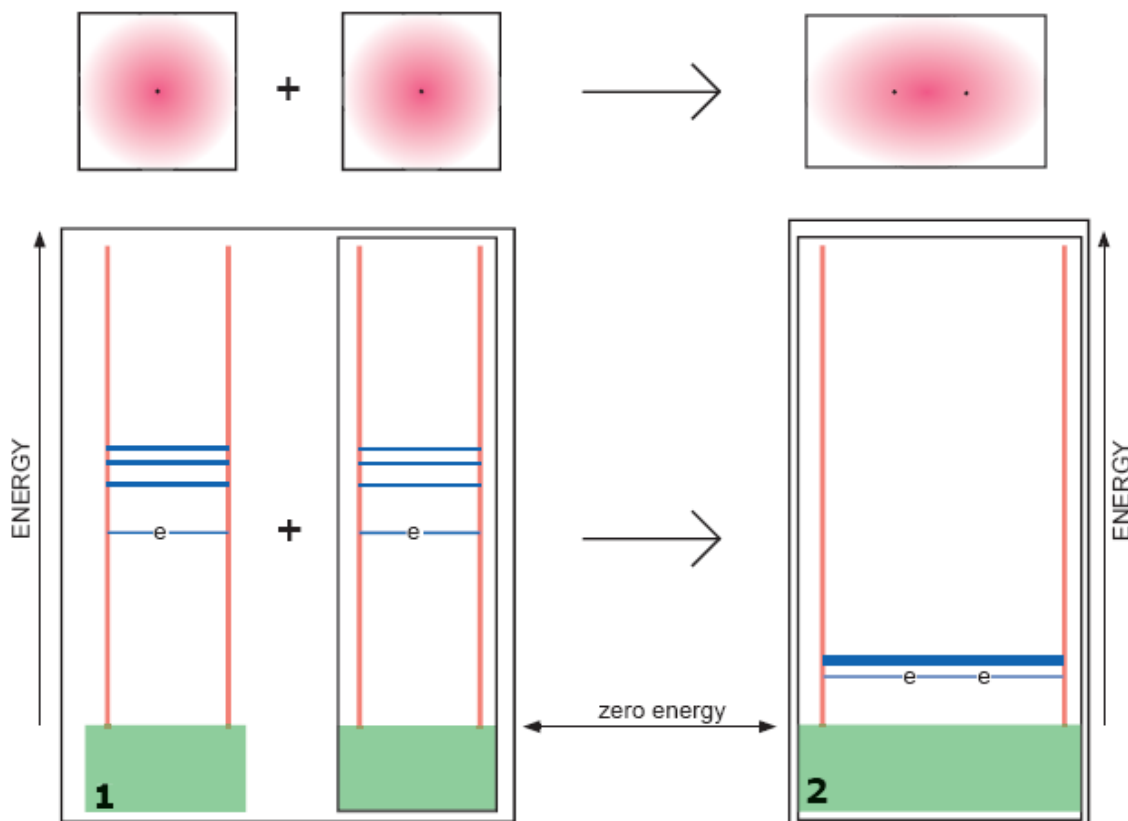


## Changes in Potential Energy in the Bonding Process - Continued

### QUANTIZATION

Just below we show two hydrogen atoms and their combination as H<sub>2</sub> on the right.

The red electron cloud represents the probable location of the electrons. Notice that the space for electrons is larger in the H<sub>2</sub> molecule than it is in the separated hydrogen atoms.



One of the basic principles of quantum mechanics is that whenever anything is confined in a finite space, it can only occupy one of a discrete set of energy levels. It is also the case that when the space is made larger the energy states are lower.

In picture 1 the blue lines represent the energy states available for a particle confined between the orange walls.

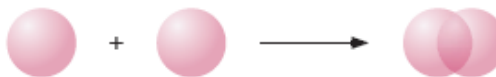
In picture 2 the blue lines show how the energy states are lower when the particles are given more space.

Which picture, number 1 or number 2, has the lowest total energy?

If picture 1 represents the energy states of two separate hydrogen atoms, then what could picture 2 represent?

These pictures do not necessarily show that if you move two hydrogen atoms close together they will bond to form a hydrogen molecule but they do show that the hydrogen molecule will be at a lower energy state than the combined energies of the separate atoms and that you would need to add energy to the molecule to get the atoms separated and that therefore the molecule will hold together until you add that energy.

Now that we know why covalent bonding occurs we will use simplified pictures known as overlapping orbitals to describe more complicated molecules. Just to the right we show this model for hydrogen.



## Terms and Concepts

Chemical Bond

Metallic bond

Electron sharing

Polarity

Potential energy

Dipole

Molecular Geometry

Covalent Bond

Bond Length

Electron transfer

Nonpolar bond

Intramolecular Force

Bond Energy

Valence Shell Electron Pair  
Repulsion

Ionic Bond

Double and Triple Bond

Ion

Polar bond

Electronegativity

Hybridization

## Websites and Resources for Bonding

Chemical Bonding Web quest

<http://nths.newtrier.k12.il.us/academics/faculty/gressel/mendelewebquest/>

Flash Animation for Ionic vs. Covalent Bonding

[http://www.mhhe.com/physsci/chemistry/animations/chang\\_7e\\_esp/bom1s2\\_11.swf](http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/bom1s2_11.swf)

VSPER Practice Problems

<http://www.chem.purdue.edu/gchelp/vsepr/cmp2.html>

Practice "flashcards" for naming/formulating ionic compounds

<http://jchemed.chem.wisc.edu/JCEDLib/WebWare/collection/open/JCE2005WWOR002/nom/nom1.html>

VSEPR 3-D Models of Molecular Shapes

[http://www2.wnorton.com/college/chemistry/gilbert/tutorials/chapter\\_07/vsepr/index.html](http://www2.wnorton.com/college/chemistry/gilbert/tutorials/chapter_07/vsepr/index.html)

IM Forces and Phase Diagram – Flash Tutorial

<http://www.wnorton.com/college/chemistry/gilbert/tutorials/ch9.htm>

H-Bonding (water) - Flash

<http://www.northland.cc.mn.us/biology/Biology1111/animations/hydrogenbonds.html>

Tutorials for Bonding Concepts (see details below)

<http://www2.wnorton.com/college/chemistry/gilbert/tutorials/ch7.htm>

Section 7.3 Expanded Valence Shells (p. 343-347)

Explore the exceptions to the octet rule, and learn to identify the conditions under which an element will expand its outer electron shell to hold more than 8 electrons. Includes practice exercises.

Section 7.6 VSEPR Model (p. 354-363)

This unit presents interactive three-dimensional representations of all the molecular geometries, as well as chemical examples of each. . Simply click and drag a molecule to rotate it in space.

Section 7.7 Hybridization (p. 363-371)

This tutorial animates the formation of hybrid orbitals from individual s and p orbitals, shows examples of their geometry, and describes how they can produce single, double, and triple bonds.. Includes practice exercises.

# Bonding Unit

## Activity #1: We're Rich! Changing Pennies into Gold

### Questions:

How do metals bond together?  
What is an alloy?

### Objectives:

To investigate metallic bonding.

### Teacher Background Information

#### Physical Properties of Metals

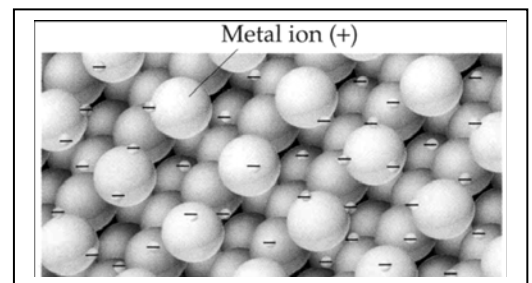
Metals, although distinct from one another, share certain similarities that enable us to classify them as metallic. A fresh metal surface has a characteristic luster. In addition, metals that we can handle with bare hands have a characteristic cold feeling related to their high heat conductivity. Metals also have high electrical conductivities; electrical current flows easily through them. The heat conductivity of a metal usually parallels its electrical conductivity. For example, silver and copper, which possess the highest electrical conductivities, also possess the highest heat conductivities. This observation suggests that the two types of conductivity have the same origin in metals, which we will soon discuss.

Most metals are malleable, which means that they can be hammered into thin sheets, and ductile, which means that they can be drawn into wires. These properties indicate that the atoms are capable of slipping with respect to one another. Ionic solids or crystals of most covalent compounds do not exhibit such behavior. These types of solids are typically brittle and fracture easily. Consider, for example, the difference between dropping an ice cube and a block of aluminum metal onto a concrete floor.

#### Electron-Sea Model for Metallic Bonding

Most metals form solid structures in which the atoms are arranged as closely packed spheres. For example, each copper atom is in contact with 12 other copper atoms. The number of valence electrons available for bond formation is insufficient for a copper atom to form an electron-pair bond to each of its neighbors. If each atom is to share its bonding electrons with all its neighbors, these electrons must be able to move from one bonding region to another.

The electron-sea model for metallic bonding cannot adequately explain all of the properties of metals. It is, however, a very simple model that can account for most of the important characteristics of metals. In this model



the metal is pictured as an array of metal cations in a "sea" of valence electrons, as illustrated in the figure. The electrons are attracted to the metal by electrostatic attractions to the cations, and are uniformly distributed throughout the structure.

However, the electrons are mobile, and no individual electron is confined to any particular metal ion. When a metal wire is connected to the terminals of a battery, electrons flow through the metal toward the positive terminal and into the metal from the battery at the negative terminal.

The high heat conductivity of metals is also accounted for by the mobility of the electrons, which permits ready transfer of kinetic energy throughout the solid. The ability of metals to deform (their malleability and ductility) can be explained by the fact that metal atoms form bonds to many neighbors. Changes in the positions of the atoms brought about in reshaping the metal are partly accommodated by a re-distribution of electrons.

### Transition Metals

Many of the most important metals of modern society are transition metals. Transition metals, which occupy the d block of the periodic table, include such familiar elements as chromium, iron, nickel, copper, silver and gold. Transition metals are commonly used for the fabrication of items and construction of buildings. Additionally, many of them are required in trace amounts for proper biological function. These include iron, manganese, zinc, cobalt and copper. Transition metals owe their location in the periodic table to the filling of the d subshells. The existence of these d electrons is partially responsible for several characteristics of transition metals:

- They often exhibit more than one stable oxidation state.
- Many of the transition metal compounds and solutions containing transition metal ions are colored. For example, solutions containing  $\text{Cu}^{+2}$  are blue,  $\text{Fe}^{+3}$  are yellow and  $\text{Ni}^{+2}$  are green.
- Transition metals and their compounds exhibit interesting magnetic properties.

### Electron Configurations and Oxidation States of the Transition Metals

When transition metals are oxidized (become positively charged ions) they lose their outer s electrons before they lose electrons from the d subshell. For example, the electron configuration of Fe is  $[\text{Ar}]3d^64s^2$ , whereas that of  $\text{Fe}^{+2}$  is  $[\text{Ar}]3d^6$ . Formation of  $\text{Fe}^{+3}$  requires the loss of one 3d electron, giving  $[\text{Ar}]3d^5$ . Most transition-metal ions contain partially occupied d subshells. The +2 oxidation state, which commonly occurs for nearly all of these metals, is due to the loss of their two outer 4s electrons. This oxidation state is commonly found for all these elements except Sc, where the +3 ion with an [Ar] configuration is particularly stable. Oxidation states above +2 are due to successive losses of 3d electrons. The maximum oxidation states of the transition metals range from +2 to +8. In general, the maximum oxidation states are found only when the metals are combined with the most electronegative elements.

## Color

The color of transition metals is due to the presence of partially filled d orbitals. In an isolated atom, all five of the d orbitals are equivalent in energy. When the atom is surrounded by solvent molecules or anions, the varying shape of the d orbitals (recall that two of the d orbitals have a shape different than the other three), cause them to have slightly different energies. The difference in energy between the levels corresponds to the energy of visible light. The metal complex is therefore able to absorb photons of visible light. When it does so, an electron jumps from a lower energy d orbital to a higher energy d orbital and the solution will appear to be the complementary color of the light absorbed. For example, a copper ion in aqueous solution is surrounded by six water molecules. This results in two sets of d orbitals that are different in energy by about  $3.2 \times 10^{-19}$  joules. A photon of light that has this energy can be absorbed causing an electron to be excited from one of the lower energy levels to one of the higher energy levels. Since the absorbed photon has a wavelength of 620 nm, orange light is being absorbed, and the solution appears to be orange's complementary color which is blue.

## Magnetism

The strongest form of magnetism is called ferromagnetism. Ferromagnetism arises when the unpaired electrons of the atoms or ions in a solid are influenced by the orientations of the electrons of their neighbors. The most stable (lowest energy) arrangement results when the spins of electrons on adjacent atoms or ions are aligned in the same direction. When a ferro magnetic solid is placed in a magnetic field, the electrons tend to align strongly along the magnetic field. When the external magnetic field is removed, the interactions between the electrons cause the solid as a whole to maintain a magnetic moment. We then refer to it as a permanent magnet. The most common examples of ferromagnetic solids are the elements Fe, Co, and Ni. Many alloys exhibit greater ferromagnetism than do the pure metals themselves. Some metal oxides (for example  $\text{CrO}_2$  and  $\text{Fe}_3\text{O}_4$ ) are also ferromagnetic. Several ferromagnetic oxides are used in magnetic recording tape and computer disks.

## Alloys

An alloy is a material that contains more than one element and has the characteristic properties of metals. Alloying is important because it is one of the primary ways of modifying the properties of pure metallic elements. As an example, pure gold is too soft to be used in jewelry, whereas alloys of gold and copper are quite hard. Pure gold is termed 24 karat; the common alloy used in jewelry is 14 karat, meaning that it is 58 percent gold ( $14/24 \times 100\%$ ). A gold alloy of this composition has suitable hardness to be used in jewelry. Other examples are given below. In this lab, heating the penny fuses the zinc and copper to form an alloy called brass. Brass is a homogeneous metal that varies from 60-82% Cu and from 18-40% Zn. Brass has a relatively low melting point, so the coating can be destroyed by heating the penny for too long.

Name of Alloy	Composition by Mass	Examples
Yellow Brass	67% Cu, 33% Zn	Hardware Items
Steel	97% Fe, 3% C	Steel chain, tools
Stainless steel	80.6% Fe, 0.4% C, 18% Cr, 1% Ni	Tableware
Plumber's solder	67% Pb, 33% Sn	Soldering Joints
Sterling silver	92.5% Ag, 7.5% Cu	Tableware
Dental amalgam	70% Ag, 18% Sn, 10% Cu, 2% Hg	Dental fillings

### Copper: A Coinage Metal

Copper is sometimes found in metallic form in nature and evidence suggests that such naturally occurring copper was used during the Stone Age. As early as 10,000 years ago, (8000 B.C.), the metal was hammered into items such as coins, jewelry, tools, and weapons. About 3800 B.C. the discovery was made in the Middle East that bronze formed when tin combined with copper, thus starting the Bronze Age. The discovery was likely accidental, occurring during the reduction of copper ore in a wood fire while some tin ore was present. The advantage of bronze is that it is sufficiently hard to keep a cutting edge, something that copper cannot do. Because of its properties, all early civilizations that produced bronze used it to create weapons as well as exquisite works of art. Brass is an alloy made of varying proportions of copper and zinc (20 to 45% Zn), which becomes harder as the percentage of zinc increases. Because it is easy to forge, cast, and stamp, brass is widely used for pipes, valves, and fittings.

Copper is used in all U.S. coins. By 1982, the price of copper rose to where it was costing the U.S. Treasury Department more than 1 cent to make a penny. Since then, to conserve copper and reduce costs, the penny has been made of 97.5% zinc and 2.5% copper, with the zinc core sandwiched between two thin layers of copper. The silver-colored coins actually contain no silver. The nickel is made of a copper (75%) and nickel (25%) alloy of uniform composition throughout the coin. The other silver-colored coins are a "sandwich" made of a pure copper core covered with a thin layer of a copper and nickel alloy (91.67% copper).

### **Sources:**

Procedure taken from: Szczepankiewicz, S. H., Bieron, J.F. and Kozik, M. The "Golden Penny"

Demonstration: An Explanation of the Old Experiment and the Rational Design of the New and Simpler Demonstration. *J. Chem. Educ.* 1995 72 386.

The background material was modified from:

Moore, J.W., C.L. Stanitski and P.C. Jurs Chemistry: The Molecular Science Brooks/Cole, Australia, 2002.

Brown, T.L., H.E. LeMay and B.E. Bursten Chemistry: The Central Science Prentice Hall, 8th edition, New Jersey, 2000.

## Lab: Preparation of "Gold" Pennies

### **Materials:**

Granular zinc  
3 M NaOH  
Evaporating dish  
Tongs  
Hot plate

### **Safety Concerns:**

Use granular zinc, not powdered zinc or zinc dust. Sodium hydroxide is caustic – Use care when heating and handling this solution. Safety glasses and a lab apron **MUST** be worn. Avoid contact of all chemicals with eyes and skin. The granular zinc can be rinsed with water and used again for this lab. Consult MSDS for any additional safety, handling, and disposal information.

### Pre-lab Procedure Check Questions

- In order to make the silver colored penny; the penny is placed in an evaporating dish containing what solution and what solid material?
- What will you use to heat the silver colored penny?
- How should you heat the silver penny (strongly or gently)?

### **Procedure**

Pair up with someone. While each person should prepare their own pennies, the two people can both use the same evaporating dish and zinc.

- Place ~2 g of granular zinc in an evaporating dish on a hot plate set to medium high.
- Add ~ 20 ml of 3 M sodium hydroxide.
- Gently heat the zinc and sodium hydroxide, being careful **NOT** to boil the solution.
- Add a penny and simmer until the penny becomes silvery. Flip it occasionally with tongs.
- Use tongs to remove the penny from the evaporating dish and rinse it with water.
- Holding one edge of the penny with tongs, **GENTLY** heat the penny in a Bunsen burner until it turns gold. Flip the penny back and forth while heating. You can ruin your penny by overheating it.
- Drop the penny in a beaker of water. Remove the penny from the beaker and dry.

### Follow-Up Questions

1. What metal made the penny appear silvery?
2. What made the penny appear gold in color?
3. If the penny is heated too strongly it changes back to a copper color. Suggest a reason for this.
4. If you had actually changed your penny into gold, how much would it be worth? You may use any means you wish to find the selling price of gold.
5. What is the name given to an alloy made from the two metals found in a penny?
6. Pennies minted prior to 1982 are made from an alloy composed of approximately 98% copper and 2% zinc.
  - a. Why might a penny be made of an alloy instead of pure copper?
  - b. If one of these pennies were examined microscopically, it would appear uniform throughout. Would you predict the penny to be heterogeneous or homogeneous?

# Bonding Unit

## Activity #2: Build – a – Molecule

### Questions

What do molecules look like?

How can we predict the molecular shapes of molecules?

### Objectives

1. Draw Lewis structures for molecules.
2. Construct models to represent the geometrical shapes of molecules.
3. Determine if a molecule is polar or nonpolar.

### Introduction

The shape of a molecule greatly affects its properties. In this activity you will use clay and toothpicks to determine the shapes of the small molecules represented by the formulas below. You will use toothpicks to represent both bonding pairs and lone pairs of electrons.

HCN	H <sub>2</sub> CO	CCl <sub>2</sub> H <sub>2</sub>	CO <sub>2</sub>
O <sub>3</sub>	H <sub>2</sub> S	PH <sub>3</sub>	SF <sub>6</sub>

### Prelaboratory Assignment

1. How do you determine which atom to use as the central atom?
2. How do you find the number of valence electrons for an element?
3. How do you adjust your calculation of electrons for ions with a negative charge?  
A positive charge?
4. How many electrons do most atoms desire to reach stability? How many electrons does hydrogen want to reach stability?
5. Draw Lewis structures for each of the molecules you will build.

### Materials

- Ruler
- Protractor
- Toothpicks
- Modeling clay

## Real-World Connections

Ideas behind molecular shape can be connected to:

Smells and scents

Miscibility of liquids – “Like dissolves like” refers to the polarity of a molecule

Physical properties of substances such as boiling and melting point

Chromatography

### Source:

Adapted from World of Chemistry Lab Generator CD. McDougal Littell

### Procedure

1. Use toothpicks and modeling clay to make models for each of the molecules given in the Prelaboratory Assignment. Make sure the toothpicks stick out an equal distance from the center clay ball (make sure this distance is the same for all molecules; in this way we can compare measurements for different molecules). Use plain toothpicks to represent bonding pairs of electrons between atoms and colored toothpicks for lone pairs of electrons.
2. Position the atoms and lone pairs of electrons as far apart from each other as possible.
3. Measure and record the bond angles between the toothpicks.
4. Record the electron pair geometry and shape of each of the molecules.

### Cleaning Up

1. Take apart your models. Return all materials as instructed by your teacher.
2. Wash your hands thoroughly before leaving the laboratory.

### Analysis and Conclusions

1. Which shape has the longest distance between the ends of the toothpicks? Which shape has the shortest distance between the ends of the toothpicks?
2. Group together molecules that have the same number of atoms coming off the central atom. Do these molecules have the same geometry? The same shape?
3. What do the plain toothpicks represent? The colored toothpicks?
4. Why do we position the atoms (clay balls) at the ends of the toothpicks as far apart as possible?

5. Just by looking at the molecular formula, are you able to predict the shape of a molecule? Explain.

6. By looking at the geometry of a molecule, how can you determine if it is polar or nonpolar?

### Summary Table

Complete the table below:

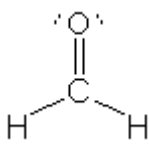
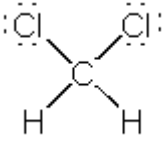

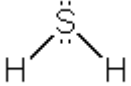
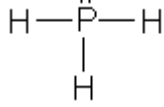
Molecule	Total number of atoms and lone pairs of electrons around central atom	Number of atoms around central atom	Number of lone pairs of electrons around the central atom	Molecular Shape	Polarity
HCN					
O <sub>3</sub>					
CCl <sub>2</sub> H <sub>2</sub>					
H <sub>2</sub> CO					
H <sub>2</sub> S					
PH <sub>3</sub>					
CO <sub>2</sub>					
SF <sub>6</sub>					

## Teacher Notes: Build - A - Molecule Activity

### Post lab Discussion

Have the students compare their results. The distances between the ends of the toothpicks are directly related to the bond angles. The actual distances between the ends of the toothpicks will differ. However, the HCN model will have the longest distance since it is linear ( $180^\circ$  bond angle).

Make sure the students understand that lone pairs affect the actual bond angles. For example,  $\text{H}_2\text{S}$ , a three atom molecule, has tetrahedral electron pair geometry but a bent shape with an angle of approximately  $109.5^\circ$ . It is not linear as students might expect since there are only two atoms bound to the central atom.

HCN	$\text{H}_2\text{CO}$	$\text{CCl}_2\text{H}_2$
$\text{H}-\text{C}\equiv\text{N}:$		
$\text{O}_3$	$\text{H}_2\text{S}$	$\text{PH}_3$
		

### Analysis and Conclusions: Answer Key

1. Which shape has the longest distance between the ends of the toothpicks? Which shape has the shortest distance between the ends of the toothpicks?

*The linear shape has the longest distance between the ends of the toothpicks. The tetrahedral shape has the shortest distance between the ends of the toothpicks.*

2. Group together molecules that have the same number of atoms coming off the central atom. Do these molecules have the same geometry? The same shape?

*2 atoms off the central atom: HCN,  $\text{O}_3$ ,  $\text{H}_2\text{S}$ ; 3 atoms:  $\text{H}_2\text{CO}$ ,  $\text{PH}_3$ ; 4 atoms:  $\text{CCl}_2\text{H}_2$ . The molecules in a given group do not necessarily have the same electron pair geometry, or the same shape.*

3. What do the plain toothpicks represent? The colored toothpicks?

*The plain toothpicks represent bonding electron pairs. Colored toothpicks represent lone pairs.*

4. Why do we position the ends of the toothpicks as far apart as possible?

*The toothpicks represent electrons, and electrons are negatively charged and therefore repel each other. To minimize electron pair interactions, we position the toothpicks as far apart as possible.*

5. Can we use the molecular formula to predict the shape of a molecule? Explain.

*We cannot use a molecular formula to predict the shape of small molecules. We need to also know how many lone pairs of electrons are present (from the Lewis structure) to determine the shape.*

6. By looking at the geometry of a molecule, how can you determine if it is polar or nonpolar?

*Polarity is determined by the symmetry of the molecule. To determine symmetry, look at the atoms and the lone pairs on the central atom. If there are any lone pairs of electrons on the central atom, the molecule is polar. If the molecule has the same atoms around the central atom, then it is nonpolar.*

See the next page for the answer key to the chart students fill out during the lab activity.

Molecule	Total number of atoms and lone pairs of electrons around central atom	Number of atoms around central atom	Number of lone pairs of electrons around central atom	Molecular Shapes	Polarity
HCN	2	2	0	Linear	Polar
O <sub>3</sub>	3	2	1	Bent	Polar
CCl <sub>2</sub> H <sub>2</sub>	4	4	0	Tetrahedral	Polar
H <sub>2</sub> CO	3	3	0	Trigonal planar	Polar
H <sub>2</sub> S	4	2	2	Bent	Polar
PH <sub>3</sub>	4	3	1	Trigonal pyramid	Polar
CO <sub>2</sub>	2	2	0	Linear	Nonpolar
SF <sub>6</sub>	6	6	0	Octahedral	Nonpolar

# Bonding Unit

## Activity #3: Ionic, Covalent, and Metallic Bonding Lab

### Question

What physical properties do ionic, covalent, and metallic compounds exhibit?

### Materials

NaCl sodium chloride (ionic)  
C<sub>6</sub>H<sub>12</sub>O<sub>22</sub> sugar (polar covalent)  
C<sub>18</sub>H<sub>36</sub>O<sub>6</sub> stearic acid (nonpolar covalent)  
Aluminum granules (metallic)  
Test tubes  
Balance  
Bunsen burner  
Test tube holder  
Small beakers  
Stirring rods  
Conductivity tester  
Mortars and Pestles

### Safety Concerns

Wear goggles and lab apron. Avoid contact of all chemicals with eyes and skin. Tell students to use small amounts of the chemicals for each test. Each lab group should only need about one gram of each substance, which should be enough to do all of the tests. When heating the substance, students should hold the mouth of the test tube away from others. All chemicals in this lab can be disposed of according to Flinn Suggested Disposal Methods #26a and b. Consult MSDS for any additional safety, handling, and disposal information.

### Real World Connections

Minerals are ionic compounds that are composed of unit cells of specific geometric shapes.

The understanding of physical properties of specific compounds is the foundation for development of new materials by scientists and engineers.

### Source:

Lab adapted from M. Wiseman, IB Chemistry Teacher  
(Website: [www.mwiseman.com/courses/chem\\_ib/](http://www.mwiseman.com/courses/chem_ib/)) and also Flinn Scientific "Chemtopic Labs: Chemical Bonding Volume 5"

## Procedure

You will conduct tests on each substance to observe the following physical properties:

- Solubility of the compound in water.
- Conductivity of the solid *and* in aqueous solution. (2 separate solubility tests!)
- Melting point
- Brittleness

## Procedure Help

You can write the procedure to the lab either as separate tests that are organized by the physical property you are testing, or as one long list of steps. The procedure should include enough detail that another group could repeat exactly what you did and get similar results.

Here are some questions to think about as you plan your procedure:

- Does it matter how much of each substance you use?
- What will be the controls in your tests? The variables?
- Does the temperature of the flame matter? How can you control it?
- Does it matter if you used tap water or distilled water?

## Data Collection

As the summary of your observations, complete a data table that organizes the properties to each *general* type of compound. You may copy the one below.

**Summary of Results: Properties of Compounds**

Physical Property	Ionic	Covalent	Metallic
Solubility in Water			
Conductivity as a Solid			
Conductivity of Aqueous Solution			
Brittleness			
Melting Point			

\*Remember to look at the grading checklist to ensure that you have included all criteria as you type your lab report!

## Ionic, Covalent, and Metallic Compounds Formal Lab Grading Checklist

Name(s): \_\_\_\_\_ Date: \_\_\_\_\_ Hr: \_\_\_\_\_  
\_\_\_\_\_

### Purpose:

- The reason for performing the lab is clear \_\_\_\_
- The purpose is no more than one or two sentences long \_\_\_\_

### Materials:

- The materials list is complete \_\_\_\_
- The list includes the size and number of materials used, if applicable \_\_\_\_

### Procedure:

- Every step of the experiment is included \_\_\_\_
- The lab could be repeated using only the instructions given \_\_\_\_
- The procedure is written as numbered steps without any pronouns (I, you, we) \_\_\_\_

### Results: (Data Table)

- Quantitative data was recorded with units and significant figures \_\_\_\_
- All tables of information are properly labeled \_\_\_\_
- All tables of information are neatly written \_\_\_\_

### Summary of Results: (Data Table)

- Describes the properties of ionic and covalent compounds \_\_\_\_

### Real-World Connections:

- Describes two practical connections between ionic, covalent, and/or metallic compounds to our society \_\_\_\_

### Overall Appearance:

- The lab is grammatically correct with no spelling mistakes \_\_\_\_
- The lab is organized by the order listed on this grading sheet \_\_\_\_
- The lab is typed \_\_\_\_

### Behavior in Lab:

- Volume of voices were low \_\_\_\_
- Focused on completing lab and helping lab partners \_\_\_\_
- Did not bother other groups or engage in disruptive behavior \_\_\_\_

# Bonding Unit

## Activity #4: Molecular Modeling Lab

### Questions

What do molecules look like?

What shapes can molecules form?

How does atomic radius, bond length, and bond strength relate to each other?

### Objectives

Since molecules are too small to see, we will be using an online modeling software program to visualize what these molecules look like and to see the effect of lone pairs, atomic radius, electronegativity, and multiple bonds on molecular shape and polarity.

### Teacher Notes

A single covalent bond is formed when two atoms share a pair of electrons. Each atom provides one of the electrons of the pair. If the two atoms are alike in electronegativity, the bond is said to be nonpolar covalent. If the atoms differ in electronegativity, the more electronegative element exerts a greater attractive force on the electrons, and the bond is polar covalent. More than one pair of electrons may be shared. This results in a double or triple bond.

A group of atoms held together by covalent bonds is called a molecule. Molecules can be either polar or nonpolar. If all bonds are nonpolar, the molecule is nonpolar. If the bonds are polar, molecules can still be nonpolar if the charge distribution throughout the molecule is symmetrical. A molecule's symmetry depends on its shape, that is, the positions in space of the atoms making up the molecule. Some possible shapes are linear, bent, trigonal pyramid, tetrahedral, and trigonal planar. The geometries of a small molecule can be predicted by examining the central atom and identifying the number of atoms bonded to it and the number of unshared electrons pairs surrounding it. The shape of the molecule can be predicted using the VSEPR rule, which states that electron pairs around a central atom will position themselves to allow for the maximum amount of space between them.

The polarity of a molecule is probably one of the most important properties for its interaction in the natural world. A dipole is a polar molecule, a molecule where there is a separation of positive and negative charge. Dipoles attract each other and exert forces on each other. This is why water is a liquid at room temperature, but carbon dioxide is a gas.

## **Real-World Connections**

Ideas behind molecular shape can be connected to:

Smells and scents

Miscibility of liquids – “Like dissolves like” refers to the polarity of a molecule

Properties of substances such as boiling and melting point

Chromatography

## **Source**

Jamie Benigna: High School Chemistry Teacher, Roper High School, Birmingham, MI

## Procedure

1. Before entering the lab, complete as much of the data table as possible.
  - a. Name each compound
  - b. Draw the Lewis dot structure
  - c. Predict the geometry
  - d. Predict the polarity
    - i. If there is no dipole, write "no dipole" in the box
    - ii. If there is a dipole, draw an arrow to indicate the direction
2. At your computer, access the website  
<http://cheminfo.chem.ou.edu/~mra/jmol/jmol.php>
3. For each of the molecules listed, there is a file under the pull-down menu in the upper right corner of the screen. The numerical file number is provided under the formula in your data table. Any molecule with an "x" at the beginning is at the bottom of the list.
4. Do not worry about recording data in any of the grayed-out boxes.
5. Using the program, investigate each molecule:
  - a. You can label each element using the "element labels" check box
  - b. You can spin the molecule by clicking on one of the atoms and dragging it around
  - c. You can check the polarity of the molecule using the "molecular dipole" check box
    - i. If a dipole is present, a large dipole arrow appears, pointing towards the direction in which electrons are pulled in the molecule
    - ii. If no dipole is present, then no arrow is displayed
  - d. Be sure to check the "multiple bonds" box to check for double and triple bonds on each molecule
  - e. To identify bond length:
    - i. Double-click on one atom
    - ii. Move the cursor to the atom on the other side of the bond and double-click again
    - iii. The bond length will be displayed in nanometers (nm)
  - f. To identify bond angle:
    - i. Double-click on one of the outside atoms
    - ii. Move the cursor and click once on the central atom
    - iii. Move the cursor and double-click on a different outside atom
    - iv. The bond angle will be displayed on the model
    - v. If the molecule is not symmetrical, try checking all of the bond angles – they might not all be equal!

Name	Formula	Lewis Dot Structure	Molecular Shape	Molecular Polarity	Bond Lengths	Bond Angle
Beryllium Chloride	BeCl <sub>2</sub> I.A.1.BeCl2					
	O <sub>2</sub> VI.A.2.o2				O-O	
	N <sub>2</sub> xN2				N-N	
	CH <sub>4</sub> VI.B.3.ch4					
Chloromethane	CH <sub>3</sub> Cl VI.B.4.ch3cl				C-H C-Cl	H-C-H Cl-C-H
Formaldehyde	H <sub>2</sub> CO XH2CO				C-O C-H	H-C-H O-C-H
	CO <sub>2</sub> VI.A.4.co2					
	H <sub>2</sub> O VI.B.1.h2o					
	BF <sub>3</sub> VI.C.bf3					
	BF <sub>2</sub> Cl VI.C.4.bf2cl				B-F B-Cl	F-B-F Cl-B-F
	NH <sub>3</sub> VI.C.2.nh3					
	C <sub>2</sub> H <sub>6</sub> II.A.1.ch3ch3				C-C	
Ethene	C <sub>2</sub> H <sub>4</sub> II.A.2.ch2ch2				C=C	
Ethyne	C <sub>2</sub> H <sub>2</sub> II.A.3.chch					

## Analysis and Conclusion Questions

1. Can a molecule with nonpolar bonds ever be polar? Why or why not?
2. Can a molecule with polar bonds ever be nonpolar? Why or why not?
3. What is a dipole?
4. List the five different molecular geometries we will encounter in this lab and indicate the bond angles for each.
5. Does there appear to be a correlation between molecular geometry and polarity? Do you notice a pattern between symmetry of a molecule and polarity?
6. Did the bond angles from the software differ from those you had predicted? (Look at  $\text{BF}_2\text{Cl}$ ,  $\text{H}_2\text{CO}$ , and  $\text{CH}_3\text{Cl}$ ) Why might this occur?
7. Looking at the last three molecules, does your data support the idea that triple bonds are shorter than double bonds, which are shorter than single bonds? Why or why not?
8. Looking at the first two molecules, explain why  $\text{N}_2$  has a shorter bond length than  $\text{O}_2$ , even though oxygen has a smaller atomic radius.
9. The polarity of a substance can have a great effect on its reactivity and solubility. A rough rule for solubility is "like dissolves like." Knowing the general rule, what can you predict about the polarity of alcohol if you know that alcohol dissolves in water? Why do you think that dry cleaners do not use water or alcohol to dissolve greasy stains?

# **Bonding Unit**

## **Activity #5: Polarity Stations**

### **Question**

How can we explain interactions of matter by the concept of polarity?

### **Objective**

Explain interactions of matter by using the concepts of polarity, as well as density, bonding, and molecular geometry.

### **Materials**

Station 1: Bending Water – balloons, water

Station 2: Floating Chads – chads from an index card with one side blackened with a pencil, large Erlenmeyer flask, about 1 -2 inches depth of hexane and water

Station 3: Mini Lava Lamp – Pre-form 2L bottle, pop bottle screw cap, mineral oil, water, food coloring, glitter (optional)

Station 4: Liquid Art – Petri dish, milk, food coloring, dish soap, beaker, toothpick

Station 5: Magic Salt – balloon, salt

Station 6: Magic Sand – Regular and Magic sand, water, 2 beakers, spoon, tub for “used” sand

Station 7: Stain remover – turpentine, kerosene, alcohol, water, used motor oil, strips of cloth,

Station 8: Penny Stacker – pennies, water, alcohol

### **Safety Concerns**

Hexane is flammable! Keep away from flame and heat. Store volatile liquids in the fume hood to evaporate. Liquids can be used again for the lab the following year. Refer to MSDS and Flinn Suggested Disposal Methods for additional information on safety, handling, and disposal methods

### **Source**

Tracy Veresh, Ferndale High School, Ferndale, MI and  
Janet Lucius, Rochester High School, Rochester, MI

## Station Directions

### **Station 1: Bending Water**

**USES BALLOONS—stay away if you have a latex allergy!**

1. Run a small stream of water under the faucet.
2. Place the “unrubbed” balloon near the stream. Do you notice anything occurring?
3. Rub the FACE side of the balloon vigorously against your clothes or in your hair and bring it close to, but not touching, the stream of water. In order to charge the balloon, the hair it is rubbed on must be clean (not oily), and product-free.

### **Station 2: Floating Chads**

1. Observe the contents of the flask.
2. After making CERTAIN that the flask is tightly stoppered, GENTLY tilt the flask back and forth.

### **Station 3: Mini Lava Lamp**

1. Fill 1/3 of a 2-L pre-form mini bottle with water, and 2/3 with mineral oil.
2. Add a couple of drops of food coloring and a pinch of glitter, if desired.
3. Screw on the top of the mini bottle and shake gently
4. You may take this science project with you!

### **Station 4: Liquid Art**

1. Pour milk into the Petri dish so it is about half full. Place two drops of each color of food coloring into the milk. The drops can be placed at 12, 3, 6, and 9, as if you were looking at the face of a clock.
2. Place a small drop of soap into the center of the Petri dish using the toothpick.
3. After you are finished, pour the liquids in the pie tin down the sink. Rinse out the Petri dish with water.

### **Station 5: Magic Salt**

1. Locate the plate with table salt (NaCl) on it.
2. Rub the FACE side of the balloon vigorously in your hair or on your clothes. If you use the hair, the hair must be clean and product-free.
3. Gradually bring the FACE side of the balloon closer to the salt.
4. Wipe off the salt with a paper towel.

### **Station 6: Magic Sand**

1. Fill up two beakers about half-way with water.
2. To the first beaker, add a couple of spoonfuls of magic sand. To the second beaker, add a couple of spoonfuls of regular sand.
3. Stir each mixture with a spoon.
4. When you are finished with your observations, pour off as much water as possible into the sink. Then dump the regular sand into the designated tray. Scoop the magic sand back into the labeled beaker.

### **Station 7: Stain Remover**

#### Part 1

1. Place 20 drops of turpentine and 20 drops of kerosene in a test tube.
2. Swirl gently. Are they miscible?
3. Pour results into waste jar. Nothing goes down the drain!
4. Next, place 20 drops of kerosene and 20 drops of water in a test tube.
5. Swirl gently. Are they miscible?
6. Pour results into waste jar. Nothing goes down the drain!

#### Part 2

1. Obtain 4 pieces of cloth.
2. Label each square with one of the following letters:
  - a. Alcohol
  - b. Turpentine
  - c. Kerosene
  - d. Water
3. Lay out the 4 squares with paper towel under them at your lab station.
4. Place one drop of used motor oil on each square cloth.
5. Directly on top of the motor oil on each square cloth.
6. Directly on top of the motor oil, place 10 drops of one of the substances listed above. (Example, put 10 drops of alcohol on the square labeled "a")
7. Allow each square to sit for 5 minutes.
8. Scrub each square thoroughly under running water for 30 seconds.
9. Spread each square out on the lab table.
10. Rank your substances 1-4 based on the effectiveness of how well the stain was removed.
11. Throw the squares in the trash.

### **Station 8: Penny Stacker**

1. Place two pennies on the paper towel.
2. On one penny, count how many drops of oil fit on the penny before the oil spills onto the paper towel.
3. On the other penny, count how many drops of water fit on the penny before the water spills onto the paper towel.

## Questions and Conclusions for Polarity Stations

### Station 1: Bending Water

1. Sketch the behavior of the water with the balloon.
2. By charging the balloon you are collecting electrons on the side of the balloon that was rubbed. Use this information, along with the fact that water is a polar molecule, to explain which part of the water molecule is attracted to the balloon. Use can use words or draw, but both explanations should include knowledge about the polarity of a water molecule.

### **Station 2: Floating Chads (Chads—you know, the hole punched out of paper in a hole punch)**

1. Is the substance in the flask composed of multiple layers? \_\_\_\_\_
2. Describe the orientation of the chads in the Erlenmeyer flask. Which liquid is the grey side facing? Which liquid is the white side facing?
3. Can you get the chads to flip over (change orientation) and stay that way? If so, you must show me.
4. The chads are pieces of paper covered in pencil, which is actually just graphite composed of carbon-carbon bonds. Now that you know the gray side of the chads are covered in carbon-carbon bonds (hmmm...what kind of bond type is that?), explain what is going on at the molecular level. FYI: One layer of the substance in the flask is hexane, the other layer is water.

### **Station 3: Mini Lava Lamps**

1. After you shake the contents of the mini bottle, what do you observe?
2. Which liquid does the food coloring mix with? Why do you think it only mixes with this substance?
3. Based on your knowledge of polarity and density, explain the interaction between the mineral oil and water.

### **Station 4: Liquid Art**

1. Describe the interaction between the food coloring and milk.
2. Describe the interaction of the food coloring and milk once the soap has touched the surface.
3. After mixing, do the substances remain mixed? .
4. Give an explanation to why this phenomenon occurs.

### **Station 5: Magic Salt**

1. Use a picture or words to describe the interaction between the salt and the charged balloon.
2. Is salt a covalent or ionic compound? Based on this information (and the fact that by rubbing the balloon you collected electrons on the face side of the balloon), why would this interaction occur? Which end of the NaCl compound is attracted to the charged balloon?

### **Station 6: Magic Sand**

1. Describe the difference between the interactions of water and the two types of sand.
2. Explain how you think magic sand works.
3. Describe a way that you could possibly make your own magic sand.

### **Station 7: Stain Remover**

1. Which substances were miscible? Which were immiscible?
  
  
  
  
  
  
  
  
  
  
2. Which solvent removed the oil stain the best? Explain why this solvent worked the best by the concept of polarity.

### **Station 8: Penny Stacker**

1. How many drops of oil fit on the penny? How many drops of water?
  
  
  
  
  
  
  
  
  
  
2. Why were you able to fit more drops of \_\_\_ on the penny than drops of \_\_\_? Your response should include the concept of "hydrogen-bonding".

# Bonding Unit

## Activity #6: Polar Bears and Penguins

### Questions

What is polarity?

How do you determine the polarity of molecules?

How does polarity relate to electronegativity?

How do you determine which end is positive and negative in polar molecules?

### Objective

In this lesson you will be exploring polarity and bonding between atoms in greater detail. A comic book will provide new information about these topics and will introduce you to the concept of electronegativity, which helps us to understand partial charges.

### Teacher Notes

For more informational background and detailed lesson plan see the following website from Key Curriculum Press, "Living by Chemistry" series.

<http://www.keypress.com/x7785.xml> (Click on "Sample lesson – Teacher)

### Source

Stacy, Angelica. LHS Living By Chemistry. Emeryville, CA: Key Curriculum Press, 2006.  
Website: <http://www.keypress.com/x10.xml>

## Questions to answer

Use the cartoon called "The Bare Essentials of Polarity" to answer the following questions.

1. How does the comic book define a "polar molecule?"
2. Define electronegativity as you understand it, after reading the first two pages of the comic book.
3. Interpret the picture at the bottom of page 1. Explain how the iceberg, penguins, and polar bears represent trends in electronegativity.
4. What is the artist trying to represent when there are two polar bears arm wrestling together or two penguins arm wrestling together?
5. What three types of bonds are represented on page 3 of the comic book? What happens to the bonding electrons in each type of bond?
6. Explain why there are four scoops of ice cream in the illustration of  $O_2$  on page 3.
7. What do the six scoops of ice cream represent in the illustration of  $N_2$  on page 4?
8. Describe what you think is happening to the penguin in the  $CO_2$  molecule in the picture on page 4.
9. Name three things that the picture of  $CO_2$  on page 4 illustrates about the molecule.
10. Describe what you think is happening to the penguins in the illustration of  $H_2O$  on page 4.
11. Explain what you think the crossed arrow represents in the comic book.
12. What are the two definitions of "dipole" given in the comic book?

## Making sense

What does electronegativity have to do with polarity?

If you finish early...

Using polar bears and penguins create an illustration showing a hydrogen sulfide molecule,  $H_2S$ . (Hint: You may wish to start with a Lewis dot structure.)

# The BARE ESSENTIALS of POLARITY

by David R. Uddley

You don't have to go to the ends of the earth to find **POLAR MOLECULES**. They're all over the place. A polar molecule is just a molecule with a difference in electrical charge between two ends.

partial positive charge

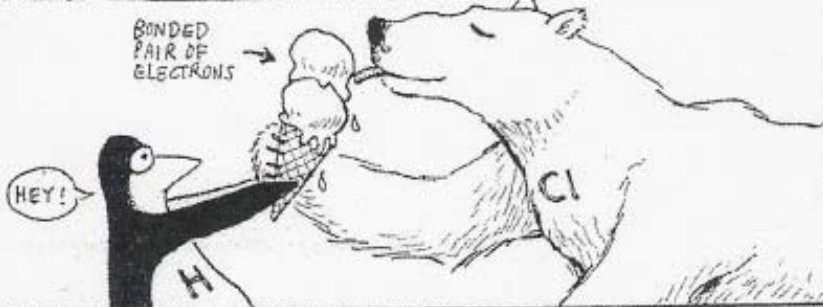
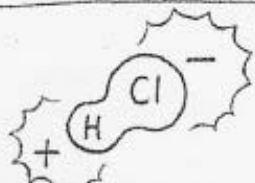


"Dipole vector" pointing toward negative charge.

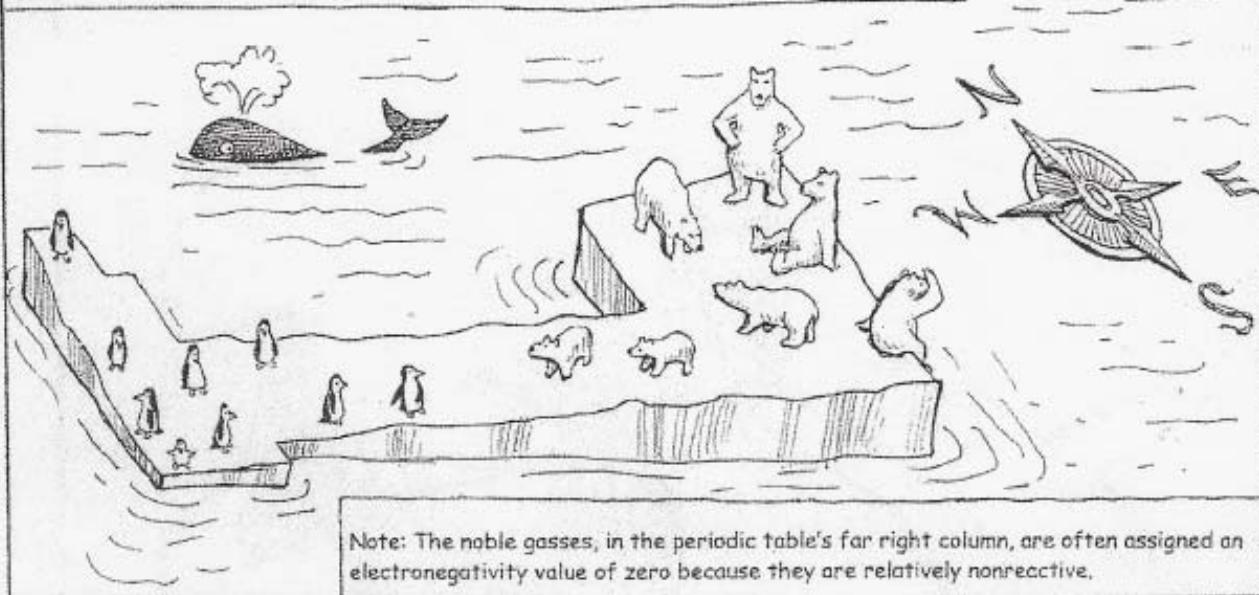
partial negative charge

The electrical imbalance of **POLARITY** is caused by differences in **ELECTRONEGATIVITY** between atoms. Electronegativity is the ability of an atom/nucleus to attract bonding electrons toward itself.

In HCl, the bonded pair of electrons spends more time near the chlorine's nucleus because chlorine is more electronegative than hydrogen.



The periodic table shows a general trend in the electronegativity of the elements. Electronegativity tends to rise as you move "northeast" on the periodic table, and fall as you move "southwest."

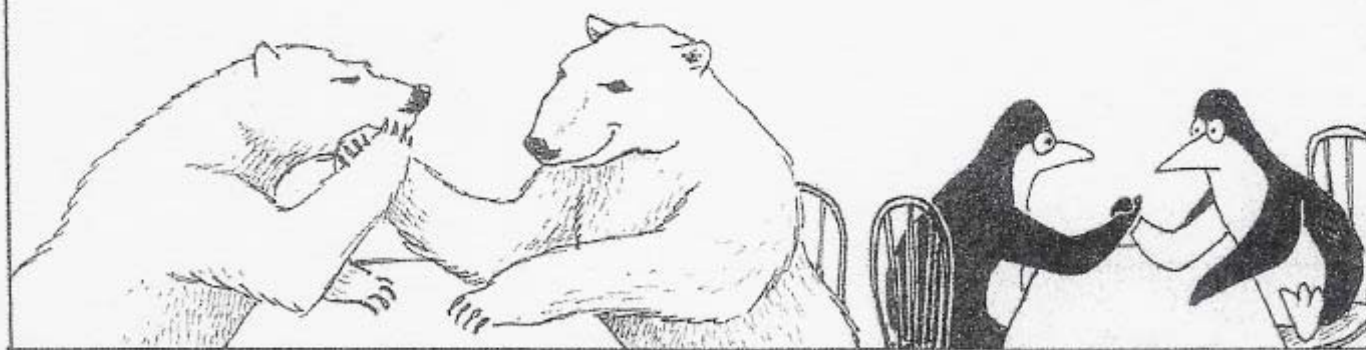


Note: The noble gases, in the periodic table's far right column, are often assigned an electronegativity value of zero because they are relatively nonreactive.

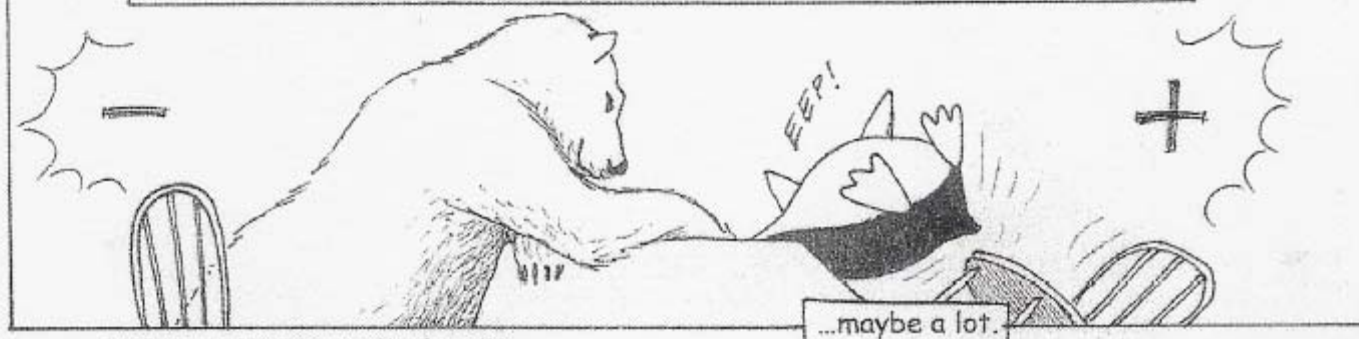
When two atoms with unequal electronegativity values bond, they do not share the bonding electrons evenly. The bonding electrons spend more time around the more electronegative atom, creating a PARTIAL NEGATIVE CHARGE on that atom. The other atom then has a PARTIAL POSITIVE CHARGE, and the bond is polar.



So the polarity of a bond is a function of the difference between the electronegativity values of two bonding atoms. Bonded atoms with equal electron-attracting strength will have nonpolar bonds.

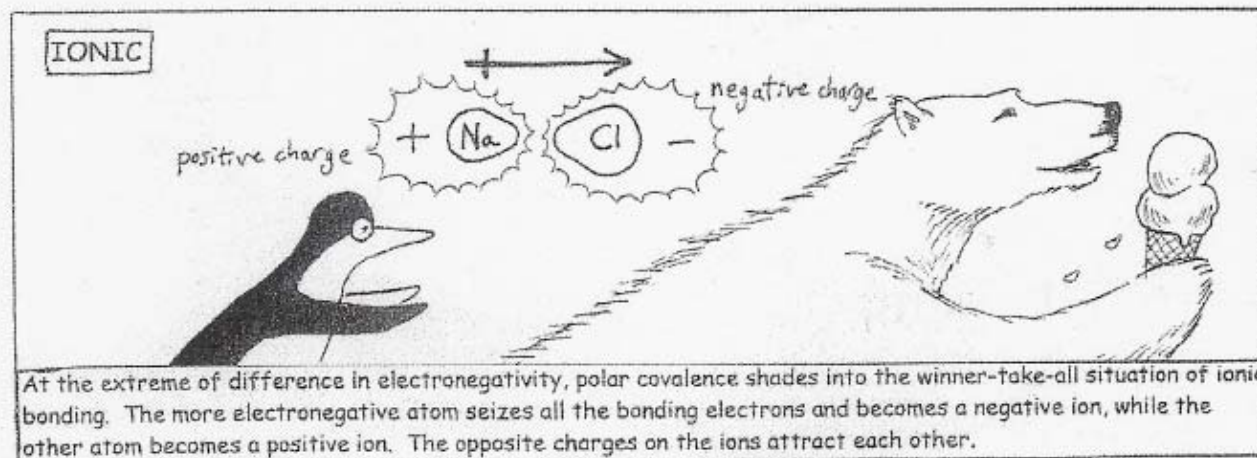
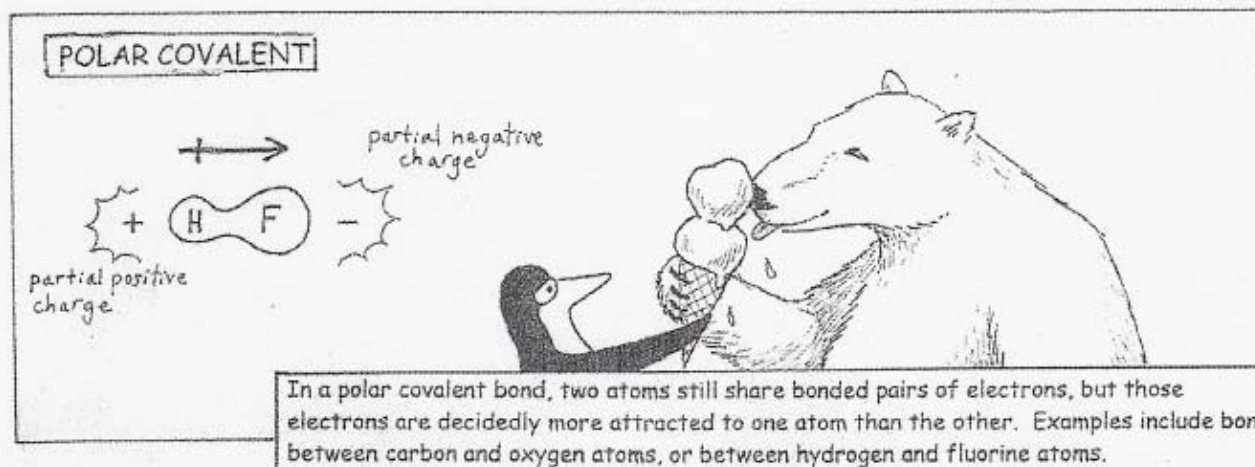
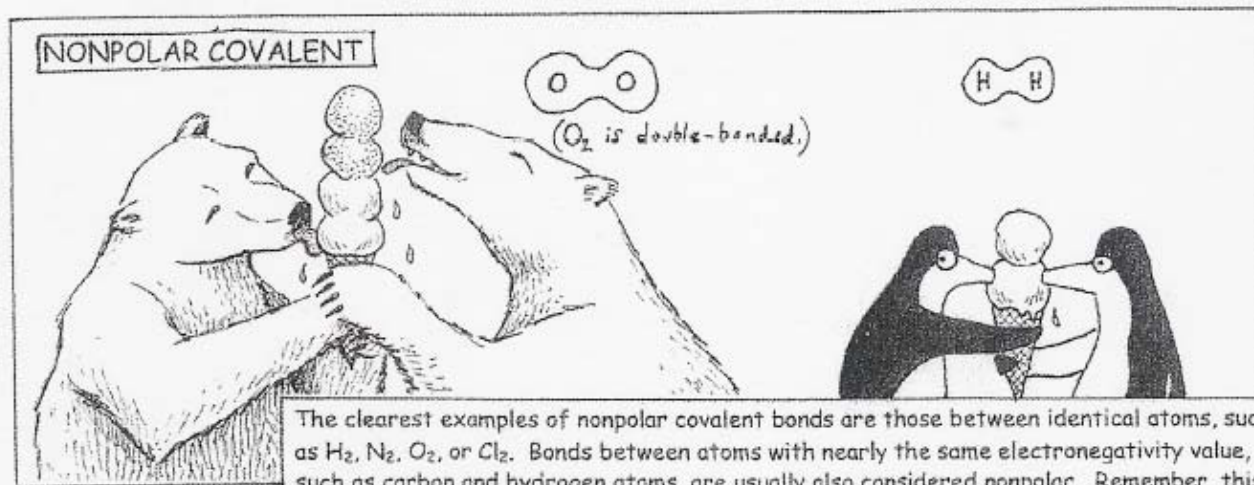


However, if the electronegativity of two bonded atoms is unequal, then their bond will be polarized—maybe a little...



...maybe a lot.

Because the elements have such varying electronegativities and can come together in so many different combinations, there is really a **CONTINUUM OF POLARITY IN BONDING**. For convenience, we can break the continuum down into three categories: (1) nonpolar covalent, (2) polar covalent, and (3) ionic.



Polar bonds between atoms constitute **DIPOLES**. Actually, the word "dipole" can refer to several different things that are relevant here: (1) the polarity of an individual polar bond between atoms, (2) the net polarity of a polar molecule that may have several polar covalent bonds within it, and (3) the polar molecule itself.

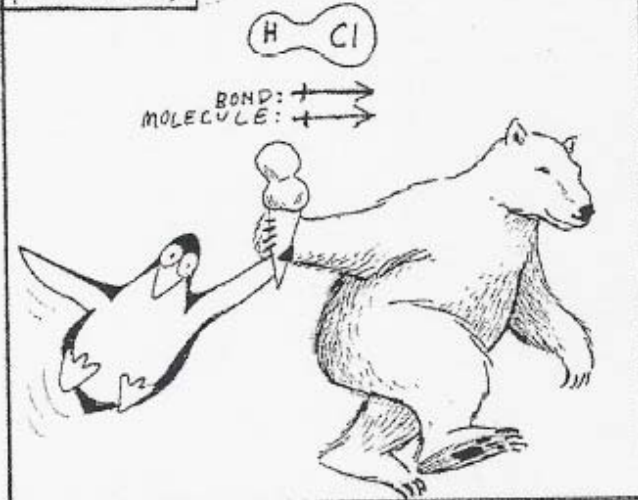


Confusing? Let's look at some examples:

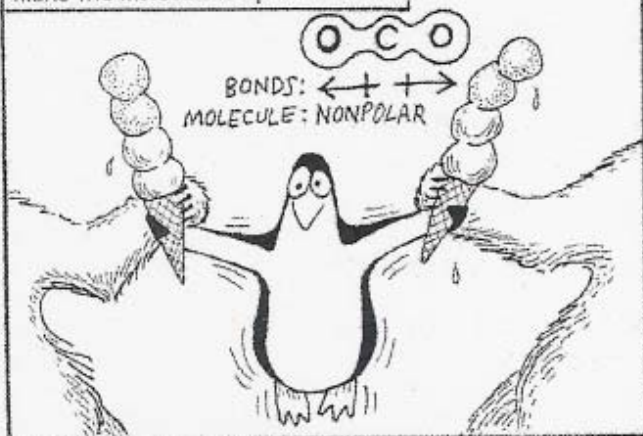
An  $N_2$  molecule isn't a dipole (it's not a polar molecule), and it doesn't have any dipoles (polar bonds) within it.



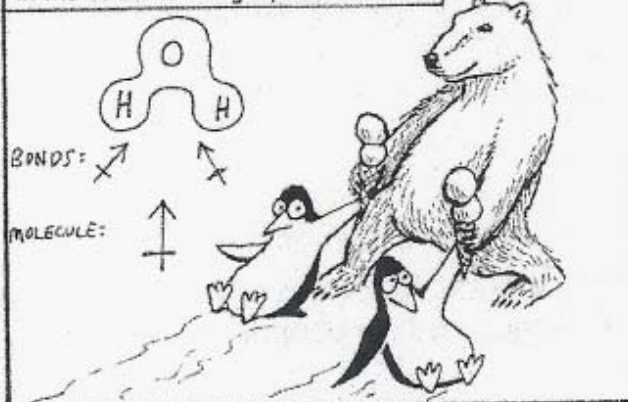
$HCl$  has a dipole (a polar bond) and it is a dipole (a polar molecule).



On the other hand,  $CO_2$  has two dipoles (two polar bonds), but the  $CO_2$  molecule itself is not a dipole because its polar bonds cancel each other out and make the molecule nonpolar overall.



Like  $CO_2$ ,  $H_2O$  has two dipoles (two polar bonds). But because of  $H_2O$ 's bent shape (caused by lone pairs of electrons on the oxygen atom),  $H_2O$  also has a dipole in the sense of an overall polarity. So  $H_2O$  is a dipole in the sense of being a polar molecule.



The polarity of molecules can affect many of their other properties, such as their solubility, their boiling and melting points, and their odor.



# Bonding Unit

## Teaching Tools: Covalent Bonds and Polarity

### Purpose

This demonstration is used to help visualize a polar versus nonpolar covalent bond.

### Construction

Straighten a large paper clip. Pass it through the centers of two Styrofoam balls. Bend a hook at one end of the wire, slide the balls onto the hook, bend a hook at the other end of the wire close to the ball, and cut off the excess wire. Find two identical rubber bands, and attach one to each hook. As the balls will represent electrons, you might wish to draw minus signs on each. Prepare a second identical model, except that one rubber band must be strong and one must be weak, but the same length.

### Procedure

To demonstrate a nonpolar covalent bond, mark the back of each hand with an "H". Grasp each rubber band of the first model (the one with identical rubber bands) with one finger, and pull to stretch the rubber bands. Notice that the "electrons" are attracted equally by both atoms.

To demonstrate a polar covalent bond, mark the back of your right hand with "H" and the back of your left hand with "F". Grasp the second model so that the weaker rubber band is held by the right "H" hand and the stronger rubber band is held by the left "F" hand. Notice that when you pull to stretch the rubber bands, the "electrons" are pulled closer to the F atom.

You may want to point out the limitations of the model; the chemical bonds result from electrical attraction (not rubber bands) and that the electrons are constantly moving at a high speed in random motion. The model is not to scale.

### Source

Guenther, Al. Practical Activities for Strengthening Your Teaching of Chemistry Concepts in Your Science Classes. Bellevue: Bureau of Education and Research, 2003.

## Bonding Unit

### Bond Length and Strength: Discussion Questions

Refer to Figure 1 for questions 1-3.

1. Which type of bond is the strongest: single, double, or triple bonds? Describe what you think may affect how strong a bond is between atoms.
2. Which type of bond is the longest: single, double, or triple bonds? Describe what you think may affect the length of bond between atoms.
3. State the relationship between bond strength and bond length. List any exceptions that you see.

Refer to Figure 2 for questions 4 and 5.

4. What type of bond is stronger, ionic or covalent bonds?
5. Describe the relationship between the charge of the ion and strength of bond in ionic compounds.

# CHEMICAL BONDING

## BOND LENGTH AND STRENGTH

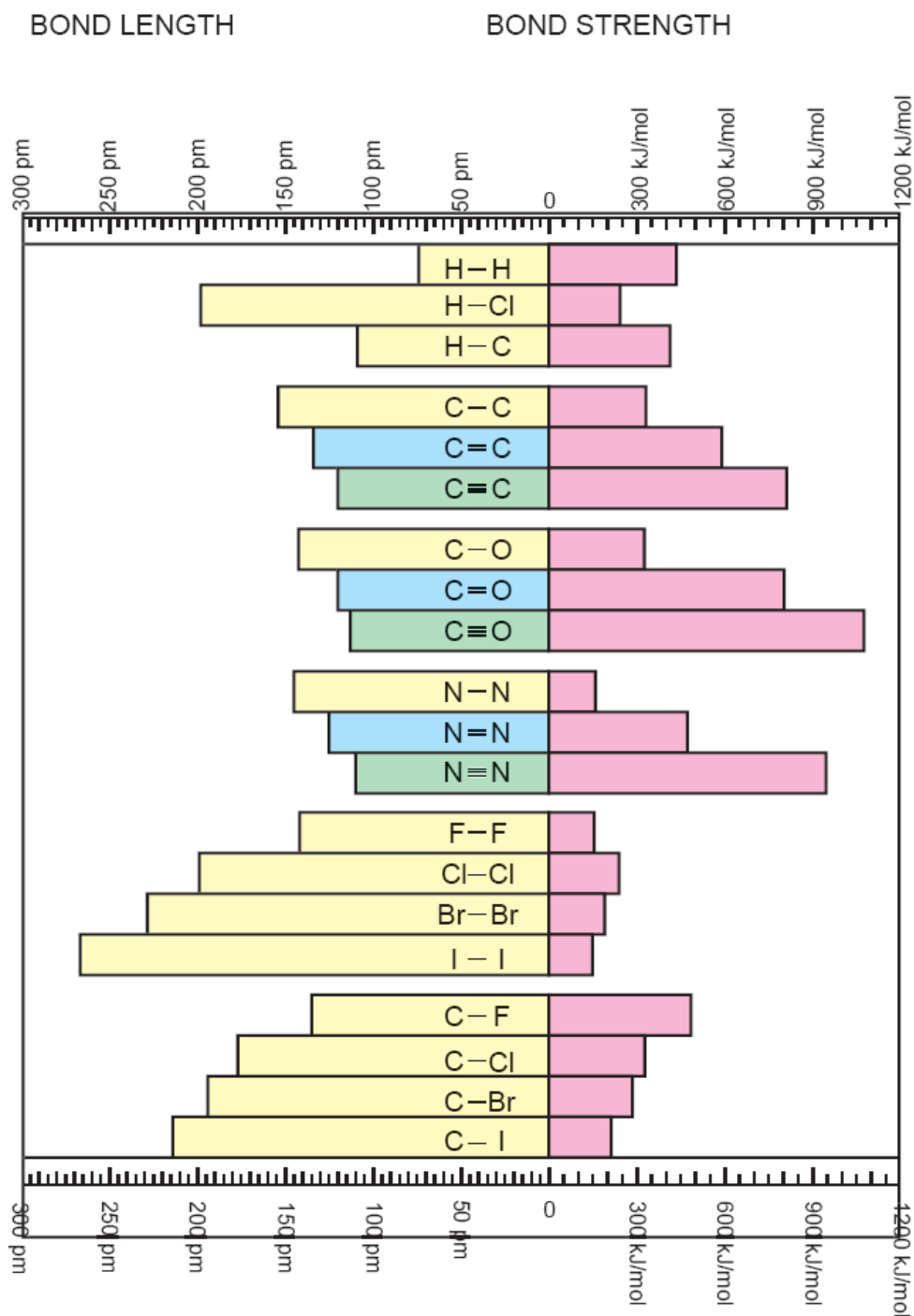


FIGURE 1

## STRONG AND WEAK BONDS

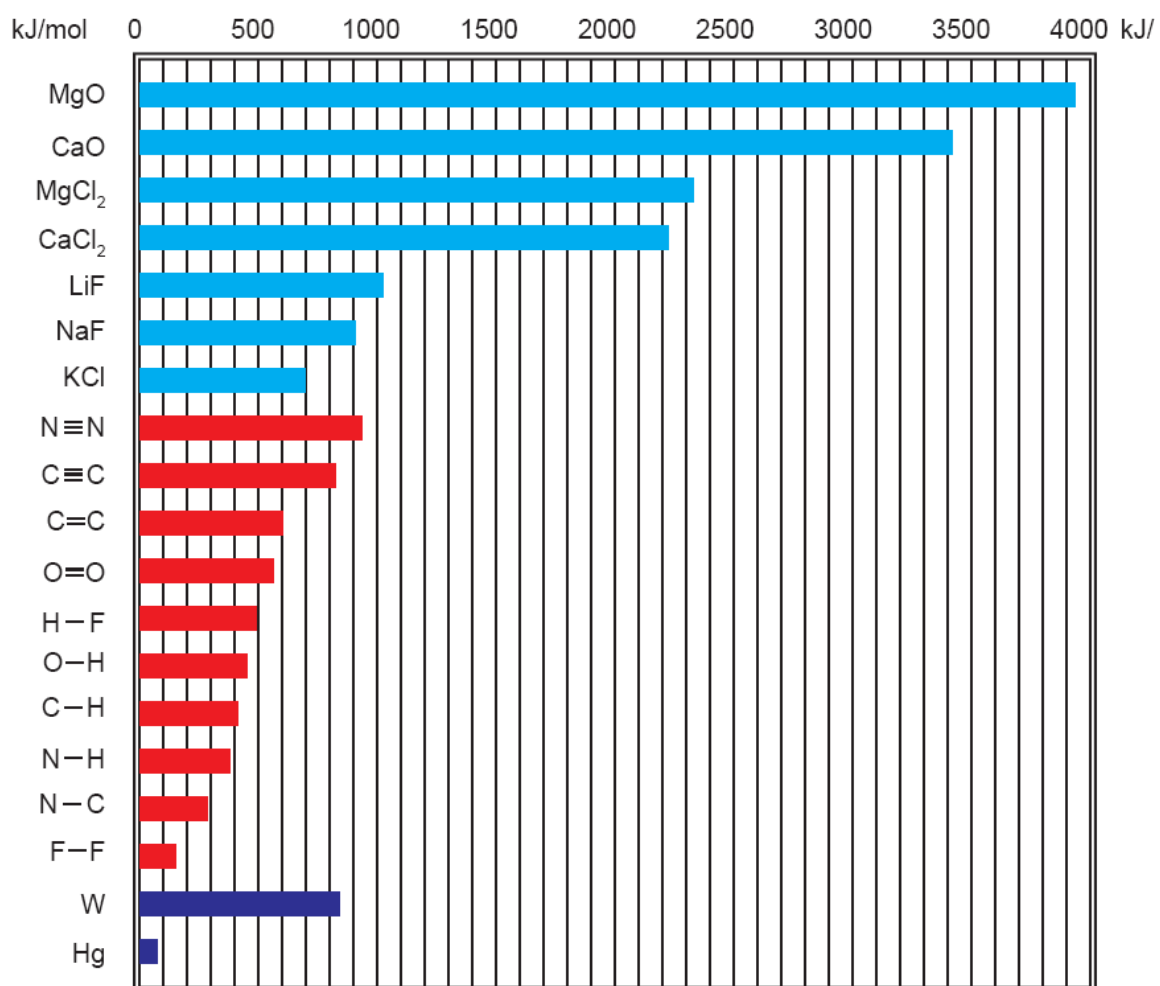
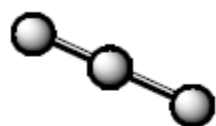


FIGURE 2

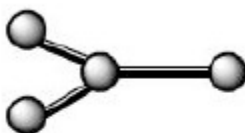
## Reference Charts and Diagrams

### Molecular Shapes and Bond Angles



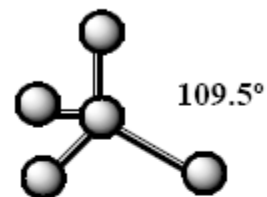
180°

**AX<sub>2</sub>**  
**Linear**  
**(BeF<sub>2</sub>)**

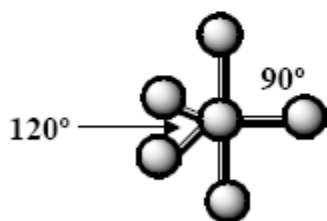


120°

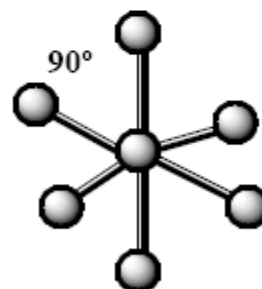
**AX<sub>3</sub>**  
**Trigonal planar**  
**(BF<sub>3</sub>)**



**AX<sub>4</sub>**  
**Tetrahedral**  
**(CH<sub>4</sub>)**

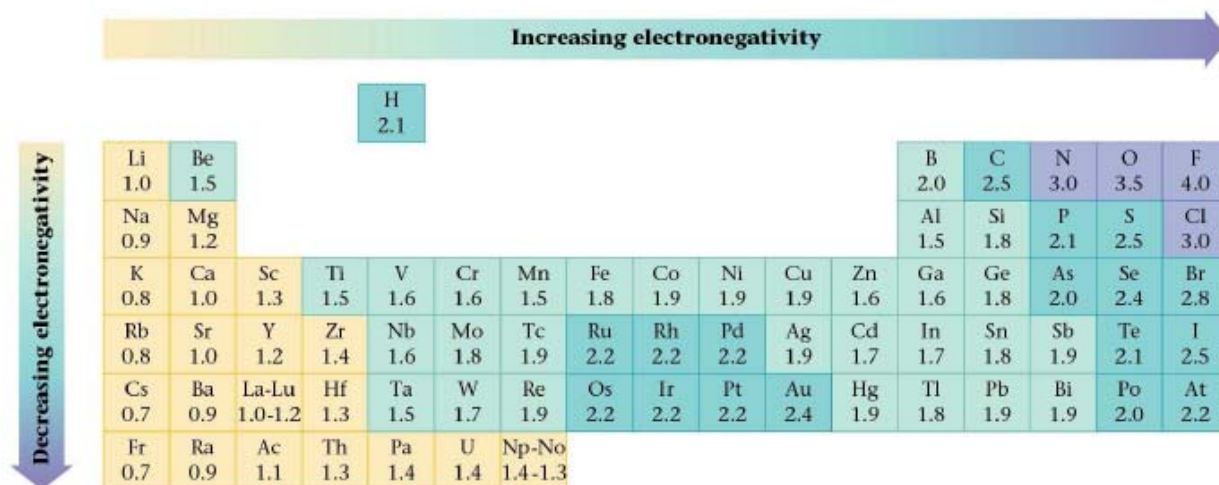


**AX<sub>5</sub>**  
**Trigonal**  
**bipyramidal**  
**(PCl<sub>5</sub>)**



**AX<sub>6</sub>**  
**Octahedral**  
**(SF<sub>6</sub>)**








# Electronegativity Chart



Key

- < 1.5
- 1.5-1.9
- 2.0-2.9
- 3.0-4.0

## Molecular Shapes

Molecular Shapes						
Example	Total pairs	Shared pairs	Lone pairs	*Molecular shape	Bond Angle	Hybrid Orbitals
BeCl <sub>2</sub>	2	2	0	 Linear	180°	sp
AlCl <sub>3</sub>	3	3	0	 Trigonal planar	120°	sp <sup>2</sup>
CH <sub>4</sub>	4	4	0	 Tetrahedral	109.5°	sp <sup>3</sup>
PH <sub>3</sub>	4	3	1	 Trigonal pyramidal	107.3°	sp <sup>3</sup>
H <sub>2</sub> O	4	2	2	 Bent	104.5°	sp <sup>3</sup>
NbBr <sub>5</sub>	5	5	0	 Trigonal bipyramidal	(vertical to horizontal) 90° / 120° (horizontal to horizontal)	sp <sup>3</sup> d
SF <sub>6</sub>	6	6	0	 Octahedral	90°	sp <sup>3</sup> d <sup>2</sup>

\*Balls represent atoms; sticks represent bonds; and lobes represent lone pairs of electrons.

## Bonding Unit

### Worksheet: Lewis Structures, VSEPR, Polarity, IM Forces

For each of the following molecules, draw the Lewis structure (with any resonance structures, if applicable), indicate the molecular shapes, indicate the molecular polarity (if any), and identify the major intermolecular force in each compound.

1) carbon tetrafluoride

2)  $\text{BF}_3$

3)  $\text{NF}_3$

4)  $\text{H}_2\text{CS}$

5) carbonate ion

6)  $\text{CH}_2\text{F}_2$

7) nitrate ion

8)  $O_2$

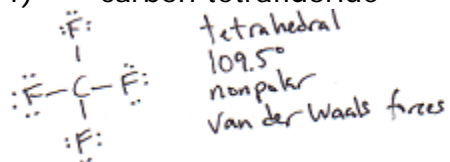
9)  $PF_3$

10)  $H_2S$

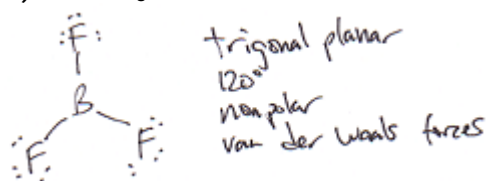
## Answer Key: Lewis Structures, VSEPR, Polarity, IM Forces -

For each of the following molecules, draw the Lewis structure (with any resonance structures, if applicable), indicate the molecular shapes, indicate the molecular polarity (if any), and identify the major intermolecular force in each compound.

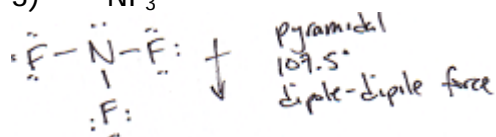
1) carbon tetrafluoride



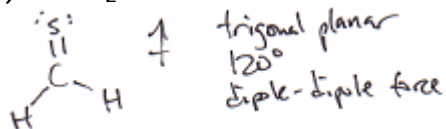
2) BF<sub>3</sub>



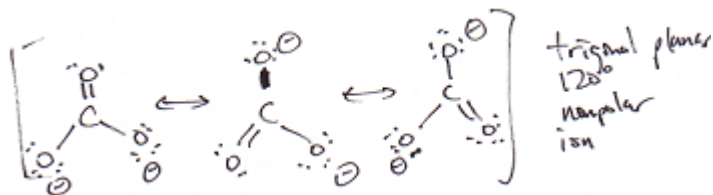
3) NF<sub>3</sub>



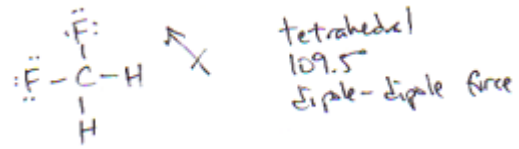
4) H<sub>2</sub>CS



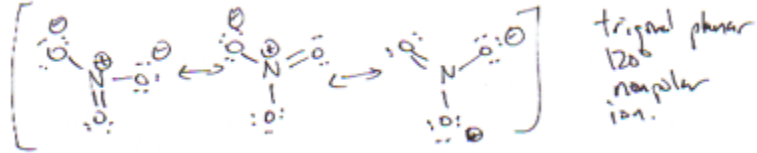
5) carbonate ion



6)  $\text{CH}_2\text{F}_2$



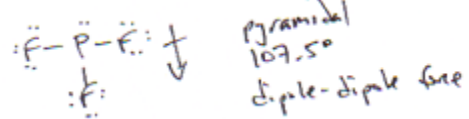
7) nitrate ion



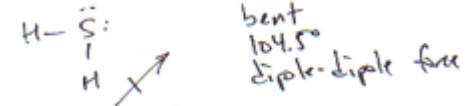
8)  $\text{O}_2$



9)  $\text{PF}_3$



10)  $\text{H}_2\text{S}$



## Worksheet: Creative Compounds

- 1) List three covalent objects that can be found in this classroom:  
a) \_\_\_\_\_ b) \_\_\_\_\_ c) \_\_\_\_\_
  
- 2) Haikus are three line Japanese poems where the first line has five syllables, the second has seven syllables, and the third has five syllables. Write a haiku about how the electrons behave when a covalent compound is formed.
  
  
  
  
  
  
  
  
  
  
- 3) Describe several tests you could conduct to determine whether a compound is ionic or covalent.
  
  
  
  
  
  
  
  
  
  
- 4) Write the names of everyday objects that have the same shapes as the following molecular shapes.  
linear: \_\_\_\_\_  
bent: \_\_\_\_\_  
trigonal pyramidal: \_\_\_\_\_  
trigonal planar: \_\_\_\_\_  
tetrahedral: \_\_\_\_\_
  
- 5) If you were a chemical, would you want to be ionic or covalent, and why?

### Answer Key: Creative Compounds

- 1) List three covalent objects that can be found in the sinks in this classroom:

**Any three covalent compounds would be fine. Common things include soaps, sponges, paper towels, pencils, pens, and other covalent objects.**

- 2) Haikus are three line Japanese poems where the first line has five syllables, the second has seven syllables, and the third has five syllables. Write a haiku about how the electrons behave when a covalent compound is formed.

**As long as the poem is three lines long with the correct number of syllables and emphasizes that covalent bonds involve the sharing of electrons, give this credit.**

- 3) Describe several tests you could conduct to determine whether a compound is ionic or covalent.

**Covalent compounds could be liquids or gases; ionic are solids at room temperature. Covalent compounds are usually softer, whereas ionic compounds are brittle. Ionic compounds have higher melting and boiling points. Ionic compounds dissolve in water and are electrolytes.**

- 4) Write the names of everyday objects that have the same shapes as the following molecular shapes.

linear: **baseball bat, golf club, pencil, pen, etc.**

bent: **boomerang, barbecue tongs, forceps, etc.**

trigonal pyramidal: **be generous when giving credit for this one, as there really aren't very many trigonal pyramidal objects in everyday use.**

trigonal planar: **Mercedes hood ornament, etc.**

tetrahedral: **some jacks, pyramids, etc.**

- 5) If you were a chemical, would you want to be ionic or covalent, and why?

**You will get some very peculiar answers for this one!**

#### **Source**

Guch, "Covalent Compound Scavenger Hunt." Worksheet Involving Ionic and Covalent Compounds. 27 Jun 2008 <<http://misterguch.brinkster.net/ioniccovalentworksheets.html>>.

