Chemical Bonding and Intermolecular Forces
Drawing Lewis Structures to Determine Molecular Geometry, Hybridization, and Molecular Polarity

OBJECTIVE

Students will identify characteristics for the three most common types of chemical bonds: ionic, covalent and metallic. Students will learn to draw Lewis structures and use them to determine the molecular geometry, hybridization and polarity of compounds and polyatomic ions. The concept of intermolecular forces will also be introduced.

LEVEL

Chemistry

NATIONAL STANDARDS

UCP.2, UCP.5, B.2

CONNECTIONS TO AP

AP Chemistry:
    I. Structure of Matter B. Chemical bonding 1. Binding forces a. Types: ionic, covalent, metallic, hydrogen bonding, van der Waals (including London dispersion forces) b. Relationships to states, structure, and properties of matter c. Polarity of bonds 2. Molecular models a. Lewis structures b. Valence bond: hybridization of orbitals, resonance, sigma and pi bonds c. VSEPR 3. Geometry of molecules and ions

AP Biology:
    I. Molecules and Cells A. Chemistry of Life

TIME FRAME

90 minutes class lecture; 45 minutes homework

MATERIALS
    periodic table
    molecular models

TEACHER NOTES

This lesson should precede the *Molecular Geometry* and the *Don't Flip Your Lid* activities found in this guide. This lesson equips students with the necessary skills to predict molecular geometries, hybridizations and molecular polarity. The students should actively participate as you introduce the different types of bonds and examples of each bond type. For each of your example compounds, be sure to point out each element's placement on the periodic table. Students should be able to deduce a good deal about the type of chemical bond formed within a compound by examining the positions of the

elements on the periodic table. If the elements are from opposite sides of the periodic table, the compound tends to be ionic; elements found closer together (most often both nonmetals) tend to form covalent compounds.

Students should also have an awareness regarding the strength of bonds. Specifically, that ionic attractions are among the strongest. Students may have questions regarding the strength or hardness of diamond. Diamonds are covalently bonded as a network, which makes them an exception to the general trends in bonding discussed here. Diamonds have very strong directional covalent bonds and are actually grouped into a class known as atomic molecular solids. Some metals, like iron also have very strong attractions and are therefore very strong while others metals, like mercury, have weaker attractions making it a liquid at room temperature.

Examples for types of compounds:
Ionic: table salt, sodium chloride, NaCl; baking soda, sodium bicarbonate, $NaHCO_3$; copper(II) sulfate, $CuSO_4$ (available at hardware stores as a root killer); lye, sodium hydroxide, NaOH (available at hardware stores)

Covalent: table sugar, sucrose, $C_{12}H_{22}O_{11}$; ammonia, NH_3; water, H_2O; any plastic material (long hydrocarbon chains)

Metallic: aluminum foil, Al; gold, Au; silver, Ag; lead, Pb (point out that pencil "lead" is not Pb but is really graphite, carbon, C)

Lewis Structures and Molecular Geometries
When discussing Lewis structures, be sure to continually refer to the periodic table and point out that the Roman numeral at the head of each column for the main group (representative) elements is in fact the number of valence electrons. Demonstrate the regions of electron density on the chalkboard, overhead or with models so that students have a concrete understanding of this concept. Remember that a region of electron density may consist of a bond or lone pair and that double and triple bonds count as only one region of electron density.

Molecular models representing each type of molecular geometry mentioned in the notes should be prepared ahead of time. Students benefit from seeing the models before being asked to construct their own. Molecular model kits may be purchased from any science supply company or may be constructed using gumdrops and toothpicks as in the *Molecular Geometry* activity found in this guide. The advantage to the purchased model sets is that bond angles are pre-established and correct.

Intermolecular attractive forces: IMF's
- **Dipole-dipole**—forces of attraction between polar molecules. Polar molecules are those which have an uneven charge distribution. Hydrochloric acid molecules are held to each other by this type of force. HCl—the chlorine pulls the electrons in the bond with greater force than hydrogen so the molecule is polar in terms of electron distribution. Two neighboring HCl molecules will align their oppositely charged ends and attract one another.

- **Hydrogen bonds**—strongest IMF—exist between H and an unshared electron pair on F, O, or N. The IMF's give rise to many unique properties. As a result of these attractions, water has a high boiling point, high specific heat, and many other unusual properties.

- **Dipole-Induced dipole**—the force of attraction that exists between a polar molecule and a nonpolar molecule (oil and water do not mix well with each other).

- **Induced dipole-Induced dipole forces or London dispersion forces (LDF)**—the force of attraction that exists between two nonpolar molecules. Liquid nitrogen, N_2, used to "burn" off warts is held by this type of force. Although easily broken, this is the predominant attractive force and exists between all types of molecules. The strength of this force increases as the number of electrons increases. We say that the strength of the LDF increases as the molecule increases in polarizability; the chance for electron distribution to become unbalanced.

Students should be able to look at the structures that are drawn and determine whether they are polar (unequal charge distribution around the central atom) or nonpolar (equal charge distribution around the central atom). Therefore, the type of IMF can be predicted.

POSSIBLE ANSWERS TO THE SELF-CHECK EXERCISES

Self-Check #1:
Predict the type of bonding found in the following compounds:
1. NaCl — ionic (a metal with a nonmetal)
2. H_2O — covalent (a nonmetal with a nonmetal)
3. Ca — metallic (only one type of atom and it is a metal)

Self-Check #2:
Predict the number of valence electrons for the following elements: Remind students that the Roman numeral group number of the representative elements will be the number of valence electrons. You may have to remind them to use the electron configurations as a sure alternative.

Element	Valence Electrons	Electron Configuration
Li	1	$1s^2 2s^1$
Ba	2	$[Xe]6s^2$
B	3	$1s^2 2s^2 2p^1$
Si	4	$[Ne]3s^2 3p^2$
N	5	$1s^2 2s^2 2p^3$
S	6	$[Ne]3s^2 3p^4$
Br	7	$[Ar]4s^2 3d^{10} 4p^5$
Ne	8	$1s^2 2s^2 2p^6$

Laying the Foundation in Chemistry

T E A C H E R   P A G E S

Self-Check #3:
Draw Lewis structures AND predict the molecular geometry of the following compounds or polyatomic ions:

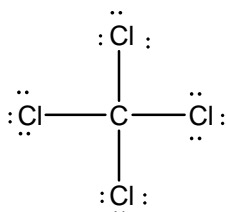1. CCl$_4$ Tally the valence electrons   C $= 1 \times 4 = 4$
    Cl $= 4 \times 7 = 28$
    Total: 32 valence electrons or 16 pair

C is in center with 4 Cl's around—this uses 8 electrons or 4 pair. We still have 12 pair—place 3 pair on each terminal Cl atom. Count and make sure that all atoms have 8 electrons.
The molecular geometry is tetrahedral. The molecule is nonpolar with sp^3 hybridization and LDF attractions.

2. BF$_3$ Tally the valence electrons.   B $= 1 \times 3 = 3$
    F $= 3 \times 7 = 21$
    Total: 24 valence electrons or 12 pair
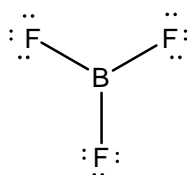
B is in the center with the 3 F's around at angles of 120º. This uses 6 electrons or 3 pairs—use the other 9 pairs and place them around each terminal fluorine atom. Boron has only 6 electrons. It is often an octet exception due to its small size and having only 5 protons. Each F is filled with eight.
The molecular geometry is trigonal planar. The molecule is nonpolar with sp^2 hybridization and LDF attractions.
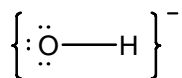
3. OH$^-$ Tally the valence electrons.   O $= 1 \times 6 = 6$
    H $= 1 \times 1 = 1$
    Negative one charge $= 1$
    Total: 8 valence electrons or 4 pair

O bonds to H with a single bond. This uses one pair. The other 3 pairs are found around the O atom. Place the entire ion in brackets and place the negative sign as a superscript on the right hand bracket. The molecular geometry may be recorded as linear. The ion is polar. Its IMF would be ion-ion or dipole-dipole.

QUESTIONS

For each of the following molecules, draw a Lewis structure and predict the molecular geometry, hybridization, overall polarity, and IMFs of the molecule. Draw your final diagram as a 3-dimensional structure. The following geometries may be used: linear, trigonal planar, tetrahedral (as well as the variations of tetrahedral—trigonal pyramidal and bent), trigonal bipyramidal, and octahedral.

1. BeH_2

H——Be——H
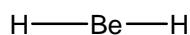
Valence electrons:

Be = $1 \times 2 = 2$

H = $2 \times 1 = 2$

Total: 4 electrons or 2 pair

Molecular Geometry: linear; nonpolar; sp; LDF

2. $SiBr_4$

Valence electrons:
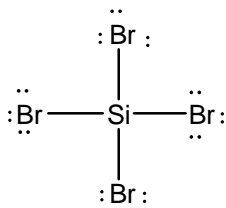
Si = $1 \times 4 = 4$

Br = $4 \times 7 = 28$

Total: 32 electrons or 16 pair

Molecular Geometry: tetrahedral; nonpolar; sp^3; LDF

3. BF_3

Valence electrons:

B = $1 \times 3 = 3$

F = $3 \times 7 = 21$

Total: 24 electrons or 12 pair

Molecular Geometry: trigonal planar; nonpolar; sp^2; LDF

4. CO_2

Valence electrons:
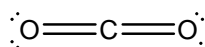
C = $1 \times 4 = 4$

O = $2 \times 6 = 12$

Total: 16 electrons or 8 pair

Molecular Geometry: linear; nonpolar; sp; LDF

5. $*SO_4^{2-}$

Valence electrons:

S = $1 \times 6 = 6$

O = $4 \times 6 = 24$

2^- charge = 2

Total: 32 electrons or 16 pair

Molecular Geometry: tetrahedral; nonpolar; sp^3, [ion-ion or dipole-dipole]

* There are several correct yet nonequivalent Lewis structures that *model* the sulfate ion's molecular structure. Students will most likely draw the one above, but a chemist will argue that the one given below is more accurate due to the concept of formal charge. Formal charge = (number of valence electrons on the free atom) − (number of valence electrons assigned to the atom in the molecule). The number of valence electrons assigned to the atom in the molecule = (number of lone pair electrons) + ½(number of shared electrons). The formal charge for S in the structure above = $6 − [0 + ½ (8)] = 2$. The formal charge for S in the structure below = $6 − [0 + ½(12)] = 0$. The formal charge for O in the structure above = $6 − [6 + ½(2)] = −1$. The formal charge for O atom in the *single* bonds in the structure below = $6 − [6 + ½(2)] = −1$ while the formal charge for the O atoms in the *double* bonds = $6 − [4 + ½(4)] = 0$.

There are two fundamental assumptions about formal charges used to evaluate Lewis structures:

1. Atoms in molecules try to achieve formal charges as close to zero as possible.

2. Any negative formal charges are expected to reside on the most electronegative atoms.

Applying both of these assumptions, we would expect that the resonance structure below would more closely describe the bonding in the sulfate ion than the Lewis structure with only single bonds. To best illustrate this, consider the individual dot structure of the S and O atoms shown on the left. The equivalent resonance Lewis structures are shown on the right.
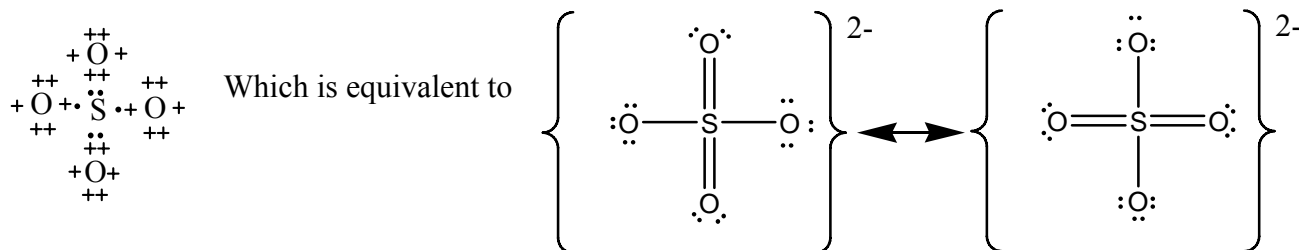
6. SF$_6$
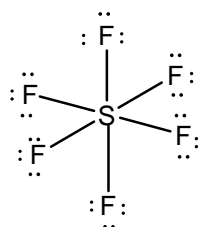
Valence electrons:

S $= 1 \times 6 = 6$

F $= 6 \times 7 = 42$

Total: 48 electrons or 24 pair

Molecular Geometry: octahedral; nonpolar; sp^{3}d^2; LDF

7. NH$_3$
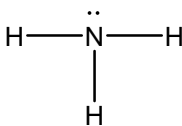
Valence electrons:

N $= 1 \times 5 = 5$

H $= 3 \times 1 = 3$

Total: 8 electrons or 4 pair

Molecular Geometry: trigonal pyramidal; polar; sp^3; dipole-dipole

8. H$_2$S
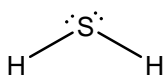
Valence electrons:

H $= 2 \times 1 = 2$

S $= 1 \times 6 = 6$

Total: 8 electrons or 4 pair

Molecular Geometry: bent or V-shape; polar; sp^3; dipole-dipole

9. PCl_5
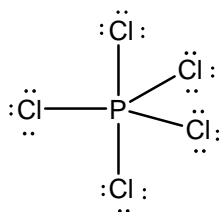
Valence electrons:

P = 1 × 5 = 5

Cl = 5 × 7 = 35

Total: 40 electrons or 20 pair

Shape: trigonal bipyramidal; nonpolar; sp^3d; LDF

10. CH_3Cl
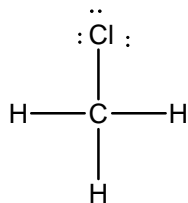
Valence electron:

C = 1 × 4 = 4

H = 3 × 1 = 3

Cl = 1 × 7 = 7

Total: 14 electrons or 7 pair

Shape: tetrahedral; polar; sp^3; dipole-dipole

Chemical Bonding and Intermolecular Forces
Drawing Lewis Structures to Determine Molecular Geometry, Hybridization, and Molecular Polarity

Have you ever wondered why some substances at room temperature exist as solids, others as liquids and still others as gases? The *intra*molecular and *inter*molecular forces of attraction can explain almost every observable property of a substance. The prefix "*intra*" means "within", thus *intra*molecular forces of attraction are otherwise known as chemical bonds which are the attractive forces between atoms within a compound. The prefix "*inter*" means "between", thus *inter*molecular forces of attraction are those occurring between molecules and are ultimately responsible for properties such as boiling point, freezing point, and physical state.

PURPOSE
In Part I of this lesson you will study the main types of *intra*molecular forces that exist within a molecule. These forces bind atoms together. You will also learn to draw Lewis structures which allow you to predict the molecular geometry, hybridization and polarity of molecules. In Part II of this lesson you will be introduced to the types of *inter*molecular forces of attraction that occur between neighboring molecules.

MATERIALS
    periodic table
    molecular models

PART I: CLASS NOTES ON BONDING
What is a chemical bond?
Chemical bonds are forces of attraction that hold atoms or groups of atoms together and allow them to function as one unit. The type of attractive force or bond formed between atoms within a molecule is directly related to the molecule's physical properties such as melting point, hardness, electrical and thermal conductivity, as well as solubility characteristics. If you think about it, most of the chemical substances you can name or identify are *not* elements. They are compounds. Matter tends to favor systems that have positions of lowest energy. That means being bound requires less energy than existing in the elemental form. It also means that energy *was released* from the system when the atoms joined together. This is a *huge* misconception that most students have—it takes energy to break a bond, not make a bond! Energy is *released* when a bond is formed (an exothermic process); therefore, it *requires* energy to break a bond (an endothermic process). Just remember "breaking up is hard to do".

Types of Chemical Bonds and Identifying Characteristics

Ionic
Characteristics of ionic substances usually include:
- electrons that are transferred between atoms having large differences in electronegativity (greater than 1.67)

- compounds containing a metal and a nonmetal (Remember that metals are located on the left side of the periodic table and nonmetals are to the right of the "stairs.")
- strong electrostatic attractions between positive and negative ions
- formulas given in the simplest ratio of elements (empirical formula; NaCl)
- crystalline structures that are solids at room temperature
- ions that form a crystal lattice structure as pictured in Figure 1
- compounds that melt at high temperatures
- substances that are good conductors of electricity in the molten or dissolved state
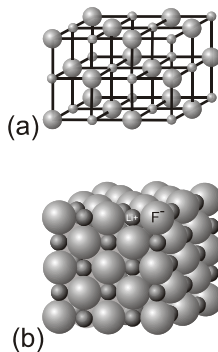
(a)

(b)

Figure 1

Covalent

Characteristics of covalent substances usually include:
- the sharing of electrons between atoms having small differences in electronegativities (less than 1.67)
- nonmetals attracted to other nonmetals
- formulas that are given in the true ratios of atoms (molecular formulas; $C_6H_{12}O_6$)
- substances that may exist in any state of matter at room temperature (solid, liquid, or gas)
- compounds that melt at low temperatures
- substances that are nonconductors of electricity

Metallic

Characteristics of metallic substances usually include:
- substances that are metals
- a "sea" of mobile or delocalized electrons surrounding a positively charged metal center
- an attraction between metal ions and surrounding electrons
- formulas written as a neutral atom (Mg, Pb)
- solids with a crystalline structure at room temperature
- a range of melting points—usually depending on the number of valence electrons
- substances that are excellent conductors of electricity since the electrons in the "sea" are free to move

Most chemical bonds are in fact somewhere between purely ionic and purely covalent.

Complete Self-Check #1 on your student answer page.

What part of the atom is involved in chemical bonding?
The valence electrons, the outermost electrons, are the only subatomic particle involved in bonding. The electron configuration for each element allows us to predict the number of valence electrons that element has available for bonding. Since the periodic table is arranged by electron configurations, it serves as an essential tool when drawing proper Lewis structures.

Complete Self-Check #2 on your student answer page.

Drawing Lewis Structures
Lewis structures allow us to predict the arrangement of atoms within a molecule, determine its molecular geometry, and help us understand how molecules interact with each other. Lewis structures are only drawn for covalently bonded molecules. Ionically bonded substances do not share electrons. They are the result of strong electrostatic attractions between + and – ions and usually form very large 3-dimensional crystalline structures. Metals are also not represented by Lewis structures. There is no true sharing of electrons because the delocalized electrons are available for any positive metal center.

Types of Covalent Bonds
- **Single bond**—one pair of electrons shared; represented by a single line drawn connecting the two atoms. Example: C——C
- **Double bond**—two pairs of electrons shared; represented by two lines drawn connecting the two atoms. Example: C═══C
- **Triple bond**—three pairs of electrons shared; represented by three lines drawn connecting the two atoms. Example: C≡══C
- **Multiple bonds**
  - are commonly formed by carbon, nitrogen, oxygen, phosphorous and sulfur. The acronym "C-NOPS" will help you remember the symbols of those elements that can form multiple bonds with themselves and with each other.
  - are stronger than single bonds. Consequently, it takes more energy to break a double bond than a single bond and still more energy to break a triple bond.
  - increase the electron density between two nuclei. As the electron density increases, the repulsion between the two nuclei decreases. Also, as the electron density increases, the attraction *each* nucleus has for the additional bonding electron pairs increases. Both of these effects draw the nuclei closer together, shortening the bond length. This means that the distance between the two nuclei is shortest for a triple bond.

Rules for Drawing Lewis Structures
1. Determine the total number of valence electrons in the molecule or polyatomic ion. Use a periodic table that has the representative elements numbered as groups IA−VIIIA. For example, C has 4 valence electrons since it is in group IVA and F has 7 valence electrons since it is group VIIA. If the structure is a polyatomic ion add the number of electrons equal to the charge for negative ions and subtract the number of electrons equal to the charge for positive ions.
2. Divide the total number of electrons by 2 to get the number of electron pairs available.
3. The element with the lowest electronegativity is the central atom in the molecule or polyatomic ion.
4. Hydrogen is always a terminal atom. It may only connect to one other atom since it can only form one bond.
5. Place one pair of electrons, as a dash representing a single bond, between each pair of bonded atoms.

6. Subtract the number of pairs used from the total the number of pairs you had available.
7. Place lone pairs around each terminal atom (except for H) to satisfy the octet rule. Left over pairs are assigned to the central atom. (If the central atom is from periods 3−7, it can accommodate more than four electron pairs due to the presence of d-orbitals.)
8. If the central atom is not yet surrounded by four electron pairs, convert one or more terminal atom lone pairs into double or triple bonds. Remember, not all elements form multiple bonds; only C, N, O, P, and S.

Exceptions to the Octet Rule

A few elements will have fewer than eight electrons.

- Hydrogen, as mentioned above, has only two electrons.
- Beryllium has a maximum of four valence electrons.
- Boron compounds are stable with only six valence electrons.

Also mentioned above, elements in periods 3 through 7 are often able to expand their octet and may have five or six electron pairs around the central atom. This is due to the availability of d-orbitals that begins with the third energy level. Even if the original atom has no electrons occupying d-orbitals before bonding, these empty d-orbitals can be drawn into the hybridization process.

Using Lewis Structures to Predict Molecular Geometry

The atoms involved in bonding will arrange themselves to minimize electron pair repulsions. This is known as the **VSEPR Theory**—Valence Shell Electron Pair Repulsion Theory. Molecular geometry is determined by the areas of electron density surrounding the central atom. If the central atom is surrounded by four bonding pairs of electrons, the molecular geometry is tetrahedral with a 109.5° bond angle. The molecular geometry is altered when one or more of the areas of electron density surrounding the central atom are occupied by an unshared electron pair (also called a lone pair). Unshared electron pairs have greater repulsive force than shared electron pairs and thus occupy more space. The increased repulsion causes the shared pairs of electrons to be "pushed" closer together. The presence of one lone pair reduces the bond angle to approximately 107°, while the presence of two lone pairs reduces it to approximately 104.5°.
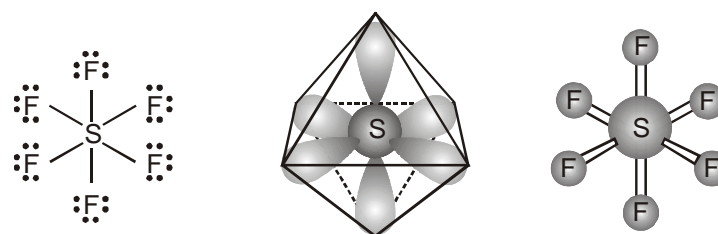
The Basic Shapes with No Unshared Electron Pairs

Linear

Trigonal Planar

Tetrahedral

Laying the Foundation in Chemistry
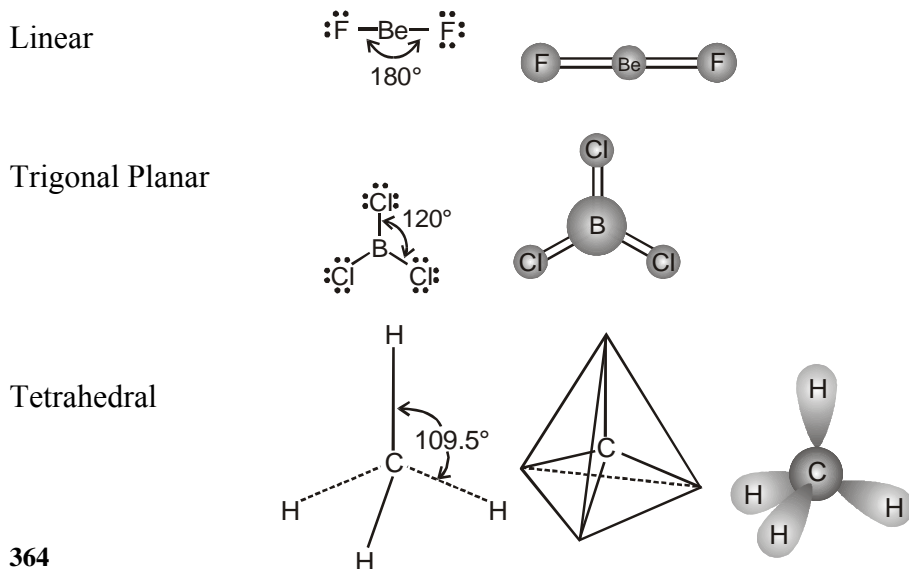
Trigonal bipyramidal
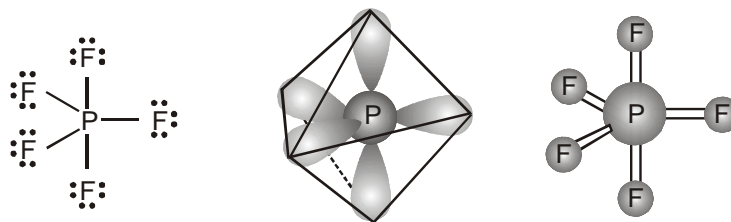
Octahedral

Variations in Molecular Geometry as a Result of Lone Pair Electrons

Ammonia, NH_3, is predicted to have a tetrahedral shape based solely on the positions of the electron pairs surrounding the central atom. Since only three of these regions contain *shared* electron pairs, with the fourth site being occupied by a *lone* pair, its molecular geometry is in fact trigonal pyramidal. The lone pair on nitrogen causes the H—N—H bond angle, to decrease from 109.5º to 107º.

(a)   (b)   (c)
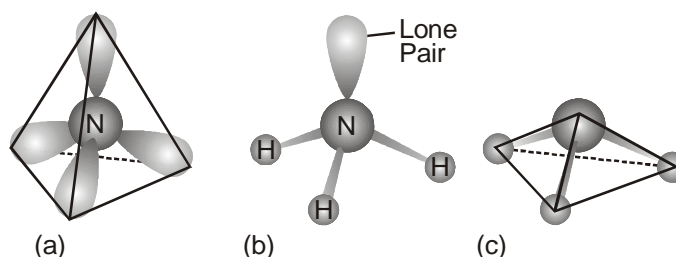
Water, H_2O, is also predicted to have a tetrahedral shape since it has four sites of electron density around the central atom. However, it exhibits a molecular geometry variation known as V-shape or bent. The two lone pairs "gang up" on the shared pairs and decrease the H—O—H bond angle from 109.5º to 104.5º.

(a)   (b)   (c)
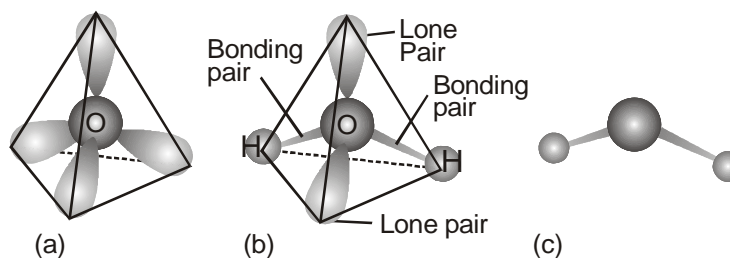
In summary, lone pairs have greater repulsive force than shared pairs and cause distortion of the remaining bond angles. The molecular geometry is determined *only* by the positions of the *atoms* surrounding the central atom. The presence of a lone pair with its increased repulsion determines the positions of the other electron pairs and thus the attached atoms. Because they are not attracted to another atom they appear "invisible" when naming the shape.
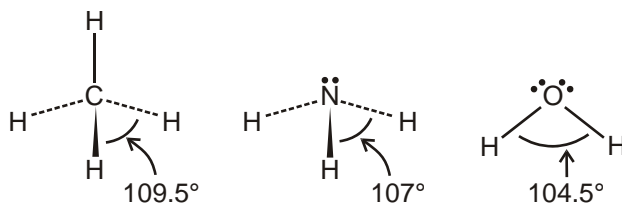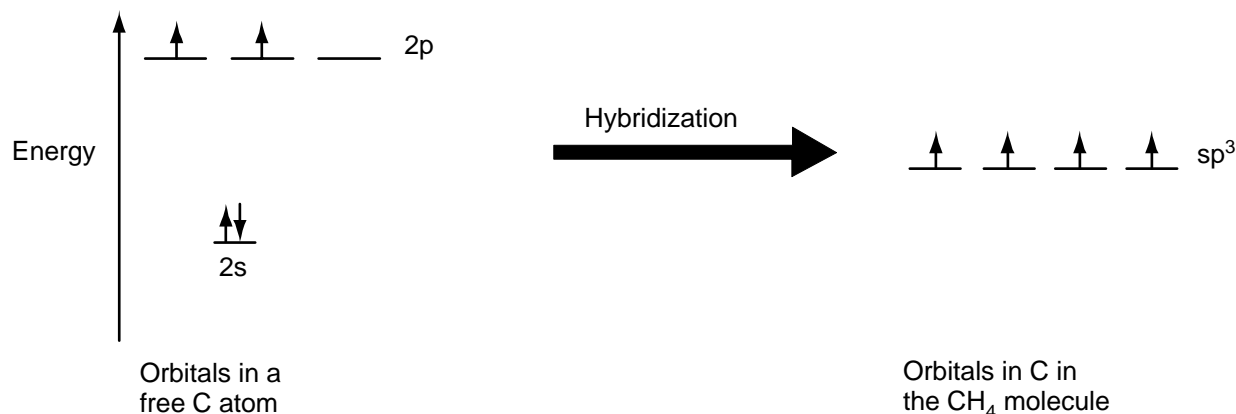

The same type of phenomenon occurs when unshared pairs occupy positions around the central atom in the trigonal bipyramidal and octahedral molecular geometries. The names of the resulting molecular geometry variations are shown in Table 10.1.
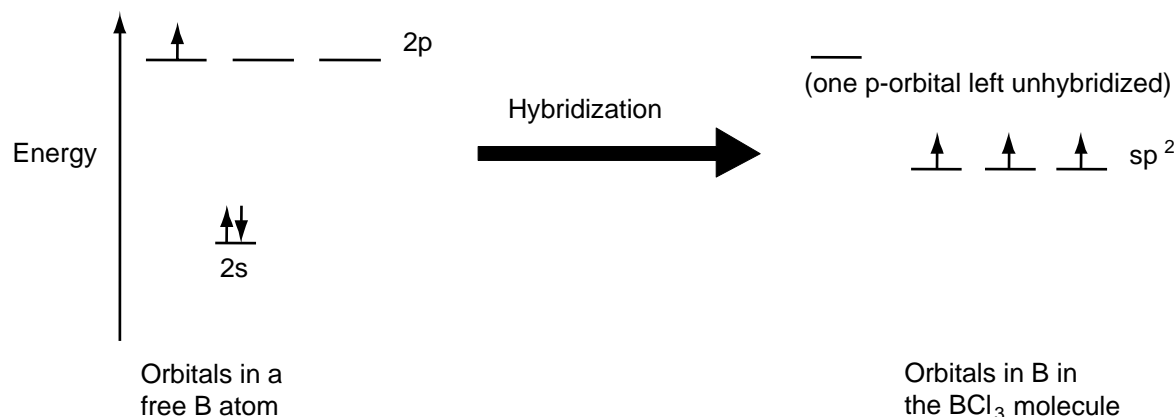
Valence Bond Theory
A second bonding theory which attempts to describe *how* the bonding occurs is known as the **valence bond theory**. The **VB** theory focuses on the atomic orbitals that must have overlapped or blended in order to obtain a particular molecular geometry. The two theories compliment each other to better explain molecular geometry.

Consider a molecule of methane, CH_4. Examine the electron configuration for the central carbon atom, $1s^2 2s^2 2p^2$. It is clear that the valence electrons are the 2s and 2p electrons. However, the 2s orbital is full, two of the p-orbitals are half-full, and the third p-orbital remains empty. How can the central C atom be using its four valence electrons to participate in four bonds? And more importantly, how can we explain the experimental evidence that shows that all four bonds have identical characteristics? The valence bond theory explains just this—the second energy level has two types of orbitals available for bonding, s and p. By blending the one s-orbital and the three p-orbitals we can create four "spaces" for each valence electron of the central C to singly occupy. The idea is similar to mixing pigments. When red and white pigments are mixed, the result is a mixture of pigments we have named "pink" which is a hybrid, or blend, of the two original pigments. The new orbitals are not **s** and not **p**; they are four hybrid orbitals, each known as **sp³**. The new name for *each* of the four newly-formed orbitals of equal energy is taken from the names of the atomic orbitals that formed it, sppp or sp³ for short. These sp³ orbitals are all equal in energy.

↑   ↑   ___   2p

Energy

Hybridization →   ↑   ↑   ↑   ↑ sp^3

↑↓
2s

Orbitals in a
free C atom

Orbitals in C in
the CH_4 molecule

Let's examine another molecule, consider BCl_3. The electron configuration for the central atom, B, is $1s^2 2s^2 2p^1$. How are three bonds of equal energy to be explained here? Again, by blending orbitals we can create a hybrid where all of the bonding electrons occupy areas of equal energy resulting in minimum overall energy usage.

↑   ___   ___   2p

___
(one p-orbital left unhybridized)

Energy

Hybridization →   ↑   ↑   ↑ sp^2

↑↓
2s

Orbitals in a
free B atom

Orbitals in B in
the BCl_3 molecule

Since boron can violate the octet rule by having only three bonds surrounding it as the central atom (due to the fact that it only has 5 protons and is extremely small), it forms three orbitals of equal energy by blending one s-orbital with two of its p-orbitals. These three orbitals are named spp or sp^2 for short.

The same pattern occurs when expanding the octet to form trigonal bipyramidal and octahedral molecular geometries. Once the one s-orbital is blended with the three p-orbitals, the d-orbitals are brought into the mix to yield five sp^3d orbitals and six sp^3d^2 orbitals, respectively. An easy way to determine the hybridization an atom is exhibiting is to count the regions of electron density around the central atom and then use that number of orbitals to build your hybrid. A multiple bond counts as *one* region of electron density as does a lone, unshared pair of electrons. Use Table 10.1 as a guide to predict the molecular geometry and hybridization of various molecules.

TABLE 10.1			
Areas of Electron Density on Central Atom	**Number of Lone Pairs on Central Atom**	**Predicted Geometry of the Molecule**	**Hybridization**
2	0	Linear	sp
3	0	Trigonal planar	sp^2
4	0	Tetrahedral	sp^3
4	1	Trigonal pyramidal	sp^3
4	2	V-shape (or bent or angular)	sp^3
5	0	Trigonal bipyramidal	sp^3d
5	1	See-saw	sp^3d
5	2	T-shape	sp^3d
5	3	Linear	sp^3d
6	0	Octahedral	sp^3d^2
6	1	Square pyramidal	sp^3d^2
6	2	Square planar	sp^3d^2

Complete Self-Check #3 on your student answer page

PART II: CLASS NOTES ON INTERMOLECULAR FORCES AND POLARITY

Molecular Polarity
Covalent chemical **bonds** between atoms are considered *polar* if the electronegativity difference between the two atoms is greater than 0.4, but less than 1.67. The term polarity refers to a separation of charge.

A **molecule** is considered polar if there are uneven electron forces acting on the central atom. Uneven forces (called dipole moments) are generally caused by the presence of a lone pair on the central atom or different kinds of atoms being bonded to the central atom. A molecule can obey the octet rule, have four polar bonds and be a *non*polar molecule since all four bonds exert an equivalent dipole moment on the central atom, canceling each other out.

Intermolecular Attractive Forces (IMF's)
Intermolecular forces are those forces that exist *between* molecules. As stated earlier, the prefix "inter-" means between. You are most likely familiar with the term "interstate" which describes the highways that connect one state to another; they are highways that allow travel between states. Intermolecular attractive forces are *not* chemical bonds. They are simply forces of attraction between neighboring molecules.

Physical properties such as melting points and boiling points can be attributed to the strength of the intermolecular attractions present between molecules. It works like this: the lower the boiling point, the weaker the intermolecular attractions; the higher the boiling point, the stronger the intermolecular attractions. For example, gasoline evaporates much more quickly than water. Therefore, the intermolecular attractive forces that hold one gasoline molecule to another are much weaker than the forces of attraction that hold one water molecule to another water molecule. In fact, water molecules are held together by the strongest of the intermolecular attractive forces, hydrogen bonds. Hydrogen bonds are not true bonds—they are just forces of attraction that exist between a hydrogen atom on one molecule and the unshared electron pair on fluorine, oxygen or nitrogen atoms of a neighboring molecule. The strands of DNA that make up our genetic code are held together by this type of intermolecular attraction.

THE TYPES OF INTERMOLECULAR FORCES IN DECREASING ORDER
- **Dipole-dipole**—the force of attraction that enables two polar molecules to attract one another. Compounds exhibiting this type of IMF have higher melting and boiling points than those exhibiting weaker IMFs.
  - **Hydrogen bonding**—the force of attraction between the hydrogen atom of one molecule and an unshared electron pair on F, O, or N of a neighboring molecule (a special case of dipole-dipole). Never confuse hydrogen bonding with a bonded hydrogen. The unique physical properties of water are due to the fact that it exhibits hydrogen bonding.
- **Dipole-induced dipole**—the force of attraction between a polar molecule and a nonpolar molecule. The polar molecule induces a temporary dipole in the nonpolar molecule. Larger molecules are more polarizable than smaller molecules since they contain more electrons. Larger molecules are more likely to form induced dipoles.
- **Induced dipole-induced dipole or London dispersion forces**—the force of attraction between two non polar molecules due to the fact that they can form temporary dipoles. Nonpolar molecules have no natural attraction for each other. This IMF is known by both names.

To better understand the induced dipole-induced dipole IMF, examine the halogens. Halogens exist as diatomic molecules at room temperature and atmospheric conditions. F_2 and Cl_2 are gases, Br_2 is a liquid and I_2 is a solid. Why? All of these molecules are completely nonpolar and according to theory, not attracted to each other, so one might predict they should all be gases at room temperature. Bromine exists as a liquid at room temperature simply because there is a greater attractive force between its molecules than between those of fluorine or chlorine. Why? Bromine is larger than fluorine or chlorine; it has more electrons and is thus more polarizable. Electrons are in constant motion so it is reasonable that they may occasionally "pile up" on one side of the molecule making a temporary negative pole on that end, leaving a temporary positive pole on the other end. This sets off a chain-reaction of sorts and this temporary dipole *induces* a dipole in its neighbors which induces a dipole in its neighbors and so on. Iodine is a solid since it is larger still, has even more electrons, is thus even more polarizable and its attractive forces are even greater leading to a solid structure.

Name_____

Period _____

Chemical Bonding and Intermolecular Forces
Drawing Lewis Structures to Determine Molecular Geometry, Hybridization, and Molecular Polarity

SELF-CHECK EXERCISES

Self-Check #1:
Predict the bond type found in the following compound:

1. NaCl _____

2. H_2O _____

3. Ca _____

Self-Check #2:
1. Predict the number of valence electrons for an atom of the following elements:

Li _____

Ba _____

B _____

Si _____

N _____

S _____

Br _____

Ne _____

Self-Check #3:
Draw Lewis structures AND predict the molecular geometry and IMF exhibited by each of the following compounds or polyatomic ions:
1. CCl_4

2. BF_3

3. OH^-

QUESTIONS

For each of the following molecules, draw a Lewis structure and predict the molecular geometry, hybridization, overall polarity and IMF exhibited by the molecule. Draw your final diagram as a 3-dimensional structure. The following geometries may be used: linear, trigonal planar, tetrahedral (as well as the variations of tetrahedral—trigonal pyramidal and bent), trigonal bipyramidal, and octahedral.

Initial Lewis Structure	**Shape & Hybridization**	**Final 3-D Structure**

1. BeH_2

   Shape: _____

   Polarity: _____

   Hybridization: _____

   IMF exhibited:_____

2. $SiBr_4$

   Shape: _____

   Polarity: _____

   Hybridization: _____

   IMF exhibited:_____

3. BF_3

   Shape: _____

   Polarity: _____

   Hybridization: _____

   IMF exhibited:_____

4. CO_2

   Shape: _____

   Polarity: _____

   Hybridization: _____

   IMF exhibited:_____

5. SO_4^{2-} Shape: _____

                            Polarity: _____

                            Hybridization: _____

                            IMF exhibited:_____

6. SF_6 Shape: _____

                            Polarity: _____

                            Hybridization: _____

                            IMF exhibited:_____

7. NH_3 Shape: _____

                            Polarity: _____

                            Hybridization: _____

                            IMF exhibited:_____

8. H_2S Shape: _____

                            Polarity: _____

                            Hybridization: _____

                            IMF exhibited:_____

9. PCl_5 Shape: _____

                            Polarity: _____

                            Hybridization: _____

                            IMF exhibited:_____

10. CH$_3$Cl   Shape: _____

Polarity: _____

Hybridization: _____

IMF exhibited:_____