Chapter 16: Covalent Bonding
Part 1 - Notes: Bonding (Metallic and Ionic) and Electron Dot Symbols

Objectives: List the characteristics of an ionic bond and classify bonds as ionic or not.
Use ionic bond characteristics to explain electrical conductivity of molten and aqueous ionic substances.
Draw electron dot structures of elements, elements, and ionic substances.
Define, identify, and explain: ionic bond, metallic bond, formula unit, electron dot picture, metallic bond, sea of electrons, and ionic substances.

Text Reference: Section 15.2 – pages 419-425

A ***chemical bond*** is . . .

There are three main types of bonding:

METALLIC BONDING

When metals "bond" with each other, they often form *metallic bonds*. Since metals have *low ionization energies*, they easily lose valence electrons. When metallic bonds are formed, metallic atoms tend to separate from valence electrons. The result is a band of positively charged nuclei with their core electrons surrounded by valence electrons that are free to move throughout the metal. These free flowing electrons are called a ***sea of electrons***. The sea of electrons is what makes it possible for metals to be hammered and stretched without breaking. The free flowing sea of electrons is also responsible for the high electrical and thermal conductivity of metals.

IONIC BONDING

An ionic bond is a chemical bond formed by the electrostatic attraction between a positive ion (cation) and a negative ion (anion). An electrostatic attraction is simply an attraction between particles with opposite charge. Remember: opposites attract.

A ***molecule*** is *two or more atoms bonded together through the sharing of electrons*. Since ions are held together by attractive forces, *not* by the sharing of electrons, **ionic compounds are not actually molecules**. There is no such thing as an ionic molecule. At room temperature, ionic compounds form solid crystal structures known as crystal lattices. An entity consisting of only one positive ion compounded with one negative ion does not exist. Ionic compounds exist as crystal lattices with repeating many patterns.

Let's examine the formation of the sodium chloride crystalline solid. Each sodium atom loses an electron to become a +1 ion and each chlorine atom gains an electron to become a –1 ion. As a sodium ion and a chloride ion approach each other, the attractive force between the two particles increases, and the potential energy of the system decreases. As more oppositely charged ions approach, the potential energy decreases even more, and the crystal begins to form. The crystal takes on a regular three-dimensional arrangement of positive and negative ions in a crystal lattice. For the crystal lattice to form, the attractive forces of the (+) and (-) ions must be greater than the repulsive forces felt between all the (+) ions and all the (-) ions. The crystal will continue to grow as long as positive and negative ions are available to be added to the crystal, and as long as the heat given off by the formation of the crystal can be effectively removed. Ionic crystals are very stable substances as indicated by their high melting points. Their stability comes from the strong attractive forces between oppositely charged ions composing the crystal lattice. Such substances are not described as molecular because **they do not exist as individual molecules in the solid state**. We refer to a single unit of an ionic solid as a ***formula unit***.

IONIC COMPOUND FORMATION

General Rules:

1. When an alkali metal reacts with a halogen, an ionic bond is formed.
2. When an alkaline earth metal reacts with a halogen, an ionic bond is formed.
3. When an alkali metal reacts with oxygen or nitrogen, an ionic bond is formed.
4. When an alkaline earth metal reacts with oxygen or nitrogen, an ionic bond is formed.

How ionic compounds are formed:

To form an ionic compound, a ***transfer of one or more electrons*** must occur. Metallic atoms lose electrons to become positive ions. Nonmetallic atoms gain electrons to become negative ions. Ions with opposite charges have an electrostatic attraction and form "ionic compounds."

One way to represent the formation of an ionic bond is by using electron configurations.

Example 1: Show the formation of the ionic compound lithium fluoride using electron configurations.

Another way to represent the formation of an ionic bond is by using ***electron dot symbols***.

ELECTRON DOT SYMBOLS

Rules for writing electron dot symbols:

1. The letter symbols represent the nucleus and core electrons of an atom.
2. Each dot represents one valence electron.
3. Dots are placed to indicate the electrons filling the orbitals.
4. For individual atoms, the s^2 fills then the p orbitals fill with each one getting one before anyone gets two. This means that one side of the square gets two dots, then the remaining three sides get electrons, one fore ach before you start doubling up.

   For particles that are part of a compound, each side gets one electron before any side of the square gets two. We will discuss why this happens later in the chapter.

Element	Electron Configuration	Orbital Diagram	Electron Dot Symbol
Sodium			
Chlorine			
Phosphorus			
Aluminum			
Carbon			

Electron dot symbols may also be used to represent ions.

Ion	Electron Configuration	Orbital Diagram	Electron Dot Symbol
Sodium			
Chloride			
Phosphide			
Aluminum			

Note: No dots are around the sodium and aluminum ions because only the core electrons from the sodium atom remain.

The formation of an ionic compound can also be represented using electron dot structures.

Example 2: Show the formation of the ionic compound sodium chloride using electron dot symbols.

Example 3: Show the formation of the ionic compound magnesium fluoride using electron dot symbols.

Question: Why are ionic compounds electrically neutral?

Chapter 16: Covalent Bonding

Part 1 - Assignment: Bonding (Metallic and Ionic) and Electron Dot Symbols

1. Write the electron dot notation for the following atoms:

Lithium	Sodium	Potassium	Calcium	Barium
Magnesium	Beryllium	Aluminum	Boron	Carbon
Fluorine	Chlorine	Neon	Bromine	Oxygen
Nitrogen	Iodine	Phosphorus	Gallium	Silicon

2. Write the electron dot notation for the following IONS.

Sodium	Magnesium	Calcium	Fluoride	Oxide
Nitride	Chloride	Bromide	Potassium	Barium

3. Use electron dot notation to represent the formation of calcium chloride from the **atoms** of calcium and chlorine.

4. Which of the following pairs will NOT form ionic compounds?

a. sulfur and oxygen
b. sodium and calcium
c. sodium and sulfur
d. oxygen and chlorine

5. Arrange the following from smallest to largest radius.

a. oxygen atom, oxide ion, sulfur ion, sulfide ion

b. sodium atom, sodium ion, potassium ion, potassium ion

Chapter 16: Covalent Bonding

Part 2 - Notes: Covalent Bonding and Polyatomic Ions

Objectives: Identify, define, and explain: covalent bond, single covalent bond, structural formulas, unshared pairs, double covalent bonds, triple covalent bonds, coordinate covalent bond, and chemical bond.
Use electron dot structures to show the formation of single, double, and triple covalent bonds in molecular compounds and polyatomic ions.

Text Reference: Section 16.1 – pages 437-447

A ***chemical bond in which atoms share electrons*** is known as a ***covalent bond***. Recall that in an *ionic bond*, electrons are *transferred* from one atom to another and form ions that are held together by electrostatic attractions. A transfer of electrons does NOT take place when atoms form a covalent bond. Instead, the atoms share one, two, or three pairs of electrons.

Covalent bonds exist as two varieties: polar covalent and nonpolar covalent. We discuss these distinctions later. For now, remember the following general rule:

Bonds that form between nonmetals are covalent.

The Octet Rule

Sharing of electrons occur if the atoms involved acquire the electron configurations of noble gases, which contain eight valence electrons (usually). So, atoms tend to form bonds to achieve a noble gas electron configuration. Atoms that achieve this configuration have eight valence electrons. The exception to his is hydrogen.

The Duet Rule

Hydrogen is the exception to the octet rule. Since hydrogen's single electron is in the first energy level, and the first energy level contains only a single s-orbital, obtaining enough electrons to have a full octet in this energy level is impossible. *Hydrogen tends to form covalent bonds to achieve the electron configuration of helium, its nearest noble gas, which has only two electrons.*

Electron Dot Structure

Electron dot structures may be used to describe covalent bonds and **structural formulas**. A **structural formula** *is the formula that shows the arrangement of atoms in a molecule or polyatomic ion*. Remember that the letter symbol of an element represents the nucleus and the core electrons of an atom and that each dot represents a valence electron.

We will use electron dot structures to show how two chlorine atoms combine to form a diatomic molecule.
Write the electron dot structure of two separate atoms of chlorine:

Note that each atom contains seven valence electrons. Greater stability is gained when each atom obtains the valence electron configuration of a noble gas (when each atom has eight electrons around it). This is achieved by ***sharing*** one pair of electrons. This shared pair of electrons is the bond between these two chlorine atoms in forming the diatomic molecule.

See how one pair of electrons is shared and how this allows each atom to have eight electrons around it:

This bond may also be represented by using orbital diagrams:

Double Bonds

Some atoms tend to share two pairs of electrons in order that they may achieve the electron configuration of a noble gas. The sharing of two pairs of electrons between two atoms is called a double bond. The atoms that have a tendency to form double bonds are:

Draw the electron dot symbols of two separate oxygen atoms:

See how two pairs of electrons are shared and how this allows each atom to have eight electrons around it:

Triple Bonds

Some atoms tend to share three pairs of electrons in order that they may achieve the electron configuration of a noble gas. The sharing of three pairs of electrons between two atoms is called a triple bond. The atoms that have a tendency to form triple bonds are:

Draw the electron dot symbols of two separate nitrogen atoms:

See how three pairs of electrons are shared and how this allows each atom to have eight electrons around it:

Now You Try It:

Write the electron dot structures to represent the diatomic molecules: hydrogen and iodine.

A problem will always give you the formula, the name, or the structure of the molecule. Whenever the term "molecule" is used, a big flashing sign in your head should go off. That sign should read "COVALENT BOND." Remember, an ionic molecule is not something that exists.

Steps for Writing Electron Dot Structures for Covalent Bonds:

1. From the name, formula, or structure of the molecule, determine how many atoms of each element are present.

2. Determine the number of valence electrons present in the molecule. This is the total of the valence electrons of all atoms of all elements.

3. Determine which atoms are connected to each other.
    a. For simpler molecules, there will be a single atom of one element and multiple atoms of another element. The single atom is the central atom in the structure.
    b. Hydrogen will not be a central atom because it only forms one bond with one other element, fulfilling the duet rule. Central atoms generally have at least eight electrons.
    c. For molecules that contain more than one element with only one atom, the central atom is either the first in the formula or it will be identified.
    d. Some molecules are complex and form chains that contain what looks like more than one central atom. If it is NOT a hydrocarbon, the position will be that listed in the formula.
    e. Hydrocarbons usually form molecules that form a chain of carbon atoms surrounded by hydrogen atoms (and possibly other elements).

4. Once the position of each of the atoms in the molecule is located, determine where the electrons belong. All the valence electrons must be used and they must be arranged so that every atom (except hydrogen) has (at least) eight electrons.

Now You Try It:

Use the steps above to draw the electron dot structures for: H_2O, $CHCl_3$, HN_3, and HCl.

You should remember that polyatomic ions are groups of covalently bonded atoms that carry a single charge. To draw a Lewis dot picture of a polyatomic ion, all you need to do is (A) add electrons to the total count for an anion or (B) subtract electrons from the total count for a cation.

Now Your Try It:

Draw Lewis dot pictures for: ammonium, NH_2^{-1}, and OH^{-1}.

Chapter 16: Covalent Bonding

Part 2 - Assignment: Covalent Bonding and Polyatomic Ions

Draw the electron dot structure to represent the bonds of each of the following molecules.

1. Hydrogen selenide
2. arsenic trihydride
3. silicon tetrahydride
4. monoxygen difluoride
5. phosphorus trichloride
6. carbon dioxide
7. hydrocyanic acid = HCN
8. methanol = CH_3OH
9. ethyne = C_2H_2
10. hydronium ion
11. acetate ion (Think – CH_3COO)
12. sulfate ion
    (Sulfur is the central atom)
13. Carbon monoxide – This illustrates a coordinate covalent bond. Remember, electrons are mobile.

Chapter 16: Covalent Bonding
Part 3 - Notes: Promotion and Hybridization

Objectives:	Identify the hybridization of the orbitals in molecules with expanded or condensed octets. Draw Lewis dot structures of atoms with expanded and condensed octets. Explain why hybridization of orbitals occurs and why octets may be expanded.
Text Reference:	Section 16.2 – pages 452-459

Promotion of electrons to higher energy orbitals frequently occurs in groups II, III, and IV. Similar promotion occurs in groups V, VI, and VII when d-orbitals are present. It requires energy to excite an electron from a *2s* to a *2p*, since *2p* is higher in energy than *2s*. However, the difference in energy between *2p* and *2s* is not very great, and the energy required is fairly low. There is an ***energy benefit*** when this promotion to higher energy levels occurs: ***extra bonds*** may be formed. These extra bonds give a molecule greater stability. So, while it may have taken energy to promote an electron to a higher energy level, but the result was well worth it as the molecule gains stability.

Beryllium has an electron configuration ending in $2s^2$. If both these electrons are in the same 2s orbital, no bonding may occur. However, if one of the electrons is promoted into the next higher 2p orbital (so the ending of its configuration is $2s^1\ 2p^1$), then beryllium has two unpaired electrons and it is able to form two bonds. A similar situation occurs with boron. Without promotion, boron's configuration of $2s^2\ 2p^1$ has only one unpaired electron and is only allowed to bond with one other atom. However, when one of the 2s electrons is promoted to one of the unoccupied 2p orbitals, the three unpaired electrons allow for boron to bond with three other atoms. Similarly, carbon has four valence electrons, $2s^2\ 2p^2$, with two electrons paired in the 2s orbital and two electrons, unpaired, in two or the 2p orbitals. When one of the 2s electrons is promoted into the last unoccupied 2p orbital, it gives the carbon four unpaired electrons and allows the carbon atom to bond with four other atoms. The extra energy that is ***released when the extra bonds are formed*** more than makes up for the energy required to promote an electron to the next highest orbital.

After promotion, the orbitals generally blend to form what are called ***hybrid orbitals***. These hybrid orbitals are a mixture of 2s and 2p orbitals; each orbital lending properties to these new hybrid orbitals. After hybridization, the orbitals are all equivalent; they can no longer be distinguished as s-orbitals or p-orbitals. The hybrid orbitals point in directions to satisfy the VSEPR principles. The directions in which the hybrid orbitals point generally determine the shape of the molecule.

The VSEPR principle states that electron pairs surrounding the central atom separate from each other as much as possible to minimize their mutual repulsions. How this principle is satisfied depends on the number of electron pairs (bonds and lone pairs) around the central atom. If there are four electron pairs (e.g., methane, carbon tetrachloride), the molecule is tetrahedral. If there are three bonds and a lone pair (total of four pairs), the molecule is pyramidal, equivalent to a tetrahedron missing one bond. Ammonia and nitrogen trichloride are good examples of pyramidal molecules. If there are three electron pairs surrounding a central atom, the shape is trigonal planar. If there are two electron pairs surrounding a central atom, the molecule is linear. The geometry of molecules will be discussed in more detail at a later point.

Additionally **–** Coordinate Covalent Bonds

Coordinate Covalent Bond is a bond where one atom in that bond contributes both of the electrons to the bond.

Group	I (1)	II (2)	III (13)	IV (14)	V (15)	VI (16)	VII (17)
Electronic Configuration-Ground state							
Dot Picture – Ground State							
Electronic Configuration-Promoted State	-----				-----	-----	-----
Dot picture after promotion	-----				-----	-----	-----
Hybridization of orbitals	-----				-----	-----	-----
Number of equivalent bonds	1	2	3	4	3	2	1
Geometry	Linear	Linear	Trigonal planar	Tetrahedral	Trigonal pyramidal	Bent	Linear
Compound with hydrogen							
Compound with chlorine							
Polar Molecule	Yes	No	No	No	Yes	Yes	Yes

Other types of Hybridization . . .

Chapter 16: Covalent Bonding
Part 3 - Assignment: Promotion and Hybridization

Draw the electron dot structure to represent the bonds of each of the following molecules.

1. phosphide
2. calcium ion
3. sulfuric acid (S is central)
4. cyanide
5. sulfur dioxide
6. phosphate ion
7. phosphorus pentabromide
8. sulfur hexachloride
9. boron trihydride
10. beryllium bromide
11. nitrate (has resonance)
12. iodine pentafluoride
(What's the hybridization?)

Chapter 16: Covalent Bonding
Part 4 - Notes: Electronegativity

Objectives: Use electronegativity values to classify bonds as polar covalent, nonpolar covalent, or ionic.
Identify, define, and explain: nonpolar covalent bond, polar covalent bond, ionic bond, and dipole.
Use delta notation to describe the polarity of bonds.
Text Reference: Section 16.3 – pages 460-466

Some of the general rules for identifying ionic and covalent bonds have been identified. Two different types of covalent bonds have been introduced: polar covalent and nonpolar covalent. To determine whether the individual bonds of a compound are ionic, polar covalent, or nonpolar covalent, an understanding of electronegativity is required.

Write the electron dot picture for the molecule HCl:

Although the hydrogen and chlorine atoms are sharing the electron pair between them, chlorine has a *greater attraction for the electron pair* and pulls it closer to its own nucleus.

This creates a *partial negative charge on chlorine* and a *partial positive charge on hydrogen*. The Greek letter "delta" is used to indicate the *equal but opposite charges* on each end of the molecule. It may be positive (+) of negative (-).

Greek letter – delta: δ

Use the delta notation to show the partial positive charge and partial negative charge on the H-Cl molecule:

A bond that has an unequal sharing of an electron pair is known as a ***dipole***.
A covalent bond that has a dipole is known as a ***polar covalent bond***.
Most covalent bonds formed by two different bonds are polar covalent.

A covalent bond that does not have a dipole – and the electrons are evenly shared – is known as a ***nonpolar covalent bond***.
Covalent bonds between two atoms of the same element are nonpolar covalent.

Electronegativity is the measure of the attraction an atom has for a shared pair of electrons in a bond.

In a polar covalent bond, the atoms with the *higher electronegativity* will take on the *partial negative charge.*
It has a stronger attraction for the shared pair of electrons – so it is the δ-.

The atom with the lower electronegativity has a weaker attraction for the shared pair of electrons.
It has a weaker attraction for the shared pair of electrons – so it is δ+.

DIPOLES ONLY OCCUR IN POLAR COVALENT BONDS!!!!!

Examine the table of electronegativities supplied. (It is also in your book on page 440 in table 13-2.)

How does electronegativity vary through a family? Why?

How does electronegativity vary through a group? Why?

Determining Bond Types from Electronegativity Values

The ***difference*** between the electronegativities of the two atoms involved in the bond will indicate whether the bond is ionic, polar covalent, or nonpolar covalent. Use the following guide.

Electronegativty difference		*Bond Type*
1.6 and greater	(difference ≥ 1.6)	Ionic bond
Between 0.5 and 1.5	($1.5 \geq$ difference ≥ 0.5)	Polar Covalent Bond
0.4 and less	(difference ≤ 0.4)	Nonpolar Covalent Bond

Write the electronegativity value for hydrogen:

Write the electronegativity value for chlorine:

Determine the difference between the two values.

What type of bond is between the H and the Cl?

What type of bond is found in H_2O?
What is the electronegativity value of hydrogen?

What is the electronegativity value for oxygen?

What is the difference in the electronegativities?

What type of bond is between the hydrogen and the oxygen?

Is the value of the electronegativity of hydrogen doubled to determine the answer to the above question?
Why or why not?

One of the most commonly used electronegativity scales, and the type of scale used in this course, is the *Pauling Electronegativity Scale*. There are other scales, but Pauling's scale has become the most widely used.

Question: What is the unit for electronegativity?
Answer:

Question: Why is this the unit for electronegativity?
Answer:

Chapter 16: Covalent Bonding
Part 4 - Assignment: Electronegativity

Directions:
Draw a Lewis dot structure for each compound.
Write the electronegativity value for each of the different types of atoms in the compound.
If the atom of an element appears more than once, you need to write it only once.
Determine the difference between the electronegativity values for each bond.
If the bond is repeated more than once, you need to write it only once.
Determine which bonds are ionic, polar covalent, and nonpolar covalent.
For polar covalent bonds, use a δ+ and δ- notation to indicate the partial charges on the dipole.

	Dot Picture	Electronegativity values	Electronegativity difference	Bond type	Dipole
KF					
O_2					
SCl_6					
K_2S					
AsH_3					
PCl_5					
HI					
NaBr					
CO_2					
NH_3					
H_2O					
NF_3					

Chapter 16: Covalent Bonding
Part 5 - Notes: Molecular Geometry

Objectives: Use VSEPR to predict the shapes of various covalent molecules.
Identify, define, and explain: bonding pair, nonbonding pair, tetrahedral, octahedral, trigonal pyramidal, trigonal bipyramidal, trigonal planar, bent, angular, linear, VSEPR, and bond angle.
State the bond angle between atoms in a covalent molecule.
List the number of bonding and nonbonding pairs in a molecule.
Text Reference: Section 16.2 – pages 452-459

Valence Shell Electron Pair Repulsion Theory

Example 1: Draw the electron dot notation for one atom of neon.

It is known that electron pairs orient themselves as far as they can from one another due to the repulsive forces that occur between particles of like charge. Look at the electron dot notation of neon, above. Have the electron pairs oriented themselves as far apart as possible? In our diagram, yes; however they are not actually 90° apart. They may appear to be 90° apart in two-dimensional space, but atoms (and the electrons around the nucleus) are in three-dimensional space. The electrons appear to be 90° apart because this is a two-dimensional representation of a three-dimensional object. In three-dimensions, the electrons are actually 109.5° apart.

In three-dimensional space, the electron pairs form a tetrahedral shape.

The ***Valence Shell Electron Pair Repulsion (VSEPR) Theory*** explains that *electrons orient themselves so that repulsion between electron pairs is minimized*. In other words, the electron pairs get as far away from one another as possible.

Molecular Shapes

A ***bond angle*** is the angle between the bonds connecting atoms of a molecule.

Tetrahedral Molecules

Example 2: Draw the electron dot structure for methane.

Note that on a two-dimensional piece of paper, the molecule appears to have bond angles of 90°. In actuality, the atoms in a molecule shaped like a tetrahedron have ***bond angles of 109.5°***. The shape of the molecule, and other molecules with this structure is said to be ***tetrahedral***.

Trigonal Pyramidal Molecules

Example 3: Draw the electron dot structure for the ammonia molecule, NH_3.

Three bonding locations exist around the central atom. Each location contains a pair of electrons and each is a ***bonding pair***.

The one unshared pair of electrons "belongs" to the nitrogen atom only has no interaction with the hydrogen atoms. This unshared pair of electrons is known as a ***lone pair*** or electrons or a ***nonbonding*** pair of electrons.

Lone pairs of electrons exert greater repulsive forces on the electrons in an atom than the forces exerted by bonding pairs of electrons.
Since the single lone pair of electrons associated with the nitrogen atom exerts a repulsive force greater than that of the three bonding pairs of electrons in the ammonia molecule, the nitrogen-hydrogen bonds are "pushed downward" and the angle is slightly smaller than in the normal tetrahedral molecule. The ***bond angle*** between the hydrogen atoms is ***107.0°***.

The molecular shape of the ammonia molecule, and other molecules with this structure, is described as ***trigonal pyramidal***. It is a three-sided pyramidal with triangular faces.

Bent Molecules

Example 4: Draw the electron dot structure for a water molecule.

Two bonding locations exist around the central atom plus two lone pairs of electrons.

The two lone pairs of electrons exert a greater repulsive force than do the two bonded pairs of electrons. Greater repulsive force by the lone pairs further reduces the bond angle in this type of molecule. The ***bond angle*** is ***104.5^o.***

The shape of the water molecule, and other molecules of this structure, is described as ***bent***.

Linear Molecule

Example 5: Draw the electron dot structure for carbon dioxide.

Two bonding locations exist around the central atom and NO lone pairs are present. Bonding electrons are at the bonding sites repelling each other to a maximum distance. The ***bond angle*** in this molecule is ***180^o***. Carbon dioxide, and other molecules of similar structure are described as ***linear***.

Exceptions to the Octet Rule

Exception 1: *One atom from group 2* covalently bonds with *two atoms of hydrogen*. This results in a molecule with only FOUR valence electrons around the central atom. The shape of such a molecule is described as ***linear*** and has a bond angle of ***180^o***.

Example 1:

Exception 2: *One atom from group 3* covalently bonds with *three halogen atoms*. This results in a molecule with only SIX valence electrons around the central atom. The shape of the molecule is described as ***trigonal planar*** and has bond angles of ***120^o***.

Example 2:

Exception 3: *One atom from group 15* covalently bonds with *5 halogen atoms*. This results in a molecule with TEN valence electrons around the central atom. The shape of this molecule is described as ***trigonal bipyramidal*** and it will have bond angles of ***90^o*** and ***120^o***. This violates the octet rule by having more than 8 electrons around the central atom. When there are more than eight atoms around the central atom, it is said to have an ***expanded octet***. (Only higher than nitrogen!!!)

Example 3:

Exception 4: *One atom from group 16* covalently bonds with *six halogen atoms*. This results in a molecule with TWELVE valence electrons around the central atom. The shape of the molecule is described as ***octahedral*** and it has bond angles of ***90^o.*** This is also a case of an expanded octet. (Only higher than oxygen.)

Example 4:

Chapter 16: Covalent Bonding

Part 5 - Assignment: Molecular Geometry

For each of the compounds listed below:

1. Determine the individual bond type.
2. Indicate the dipoles.
3. List Lone and Bonding PAIRS.
4. Draw the electron dot structure.
5. List the molecular shape and bond angle.

Compound	Bond Type(s)	Bond Dipoles	Lone & Bonding Pair	Electron Dot Structure	Shape and Angle
O_2			Lone: Bonding:		
H_2S			Lone: Bonding:		
$CaCl_2$			Lone: Bonding:		
CBr_4			Lone: Bonding:		
HCl			Lone: Bonding:		
NaCl			Lone: Bonding:		
NH_3			Lone: Bonding:		
BCl_3			Lone: Bonding:		
MgH_2			Lone: Bonding:		
PI_5			Lone: Bonding:		
SCl_6			Lone: Bonding:		
CH_2F_2			Lone: Bonding:		

Chapter 16: Covalent Bonding
Part 6 - Notes: Intermolecular Forces

Objectives: Name and describe the weak attractive forces that hold groups of molecules together.
Differentiate between intermolecular and intramolecular forces.
Identify, define, and explain: van der Waals forces, dispersion forces, dipole interactions, and hydrogen bonding.
Text Reference: Section 16.3 – pages 460-466

Until now, the concentration has been on the individual bonds that *make up* molecules and ionic units. The bonds may be ionic, polar covalent, or nonpolar covalent. Now, the study will turn to the forces that work ***between*** molecules and allow them to form solids, liquids, or gases at different temperatures and pressures. These forces are ***INTERMOLECULAR FORCES***.

Gases: When a substance is in the gaseous state, the intermolecular forces between the molecules arre so small that they are considered zero until they come very close to each other. Gases can expand and completely fill their container because their intermolecular forces are so weak.

Liquids: When a substance is in the liquid state, the intermolecular forces are strong enough to prevent its molecules from expanding away from each other, like a gas. The molecules are kept together. The intermolecular forces of a liquid, however, are weak enough to allow the individual molecules to move past one another. Liquids can flow and have no definite shape, but they do have a definite volume.

Solids: When a substance is in the solid state, the intermolecular forces are strong enough to prevent its molecules from moving past each other. Molecules in a solid can vibrate but they do not flow. Solids assume a definite shape and have a definite volume as well.

Condensation is the process by which a gas becomes a liquid. Recall that the temperature of a substance is a measure of the average kinetic energy of the particles in that substance.

Kinetic Energy:

When two molecules in the gaseous state come close to each other, a momentary attraction forms between them. *Whether the two particles condense depends upon their kinetic energy (temperature) and the strength of their intermolecular forces.* Increasing the pressure can also cause condensation by forcing the particles of a gas close enough to allow the intermolecular forces to hold them together.

If intermolecular forces are strong, condensation will occur at higher temperatures. In other words, a gas becomes a liquid at a high temperature and it is no longer a gas at that high temperature.

If intermolecular forces are weak, condensation will occur at lower temperatures. In other words, a gas becomes a liquid at a lower temperature and it is a gas until it hits this lower temperature.

van der Waals Forces

van der Waals forces are a combination of three different kinds of intermolecular forces. Those forces are dispersion forces, dipole-dipole forces, and hydrogen bonding. The ability to determine the polarity of the molecule is the key to determining how many of these forces are acting between molecules in a substance.

Dispersion Forces

Dispersion forces are the ***ONLY*** forces in ***nonpolar molecules***.

Within all molecules, momentary dipoles arise because electron distributions are constantly changing. When the electron distribution becomes unsymmetrical, it forms momentary dipoles. Dispersion forces occur when the partially negative end of a molecule's momentary dipole attracts the partially positive end of another molecule's momentary dipole.

An important fact to note is that the larger the molar mass of a ***nonpolar*** compound, the greater the dispersion forces between the molecules of that substance.

The greater the intermolecular forces in a substance, the higher the substance's boiling temperature will be. This is because, the stringer the intermolecular forces, the stronger the attraction of the molecules to one another and this means it will take more energy to separate these molecules. More energy to separate these molecules means that it will take a higher temperature before the molecules separate.

Dispersion forces can also occur between atoms. *Dispersion forces are the only forces of attraction between the atoms of a noble gas. Noble gases have very low boiling points.*

Dipole-Dipole Forces

Polar covalent molecules are subject to ***both*** dispersion forces and dipole-dipole forces. ***Dipole-Dipole forces*** arise from the electrostatic attraction between the opposite charges of (permanent) dipoles in different molecules. This signifies that these forces occur only in molecules that have molecular dipoles, in other words, molecules that are polar. Ionic and covalent bonds are about 100 times stronger than dipole-dipole forces.

Comparing two compounds of equal molar mass, one composed of nonpolar molecules and the other composed of polar molecules, the intermolecular forces are not the same, despite the molar mass. The intermolecular forces would be stronger in the polar compound because ***both*** dispersion and dipole-dipole forces are present in the polar compound as opposed to only dispersion forces in the nonpolar compound.

Hydrogen Bonding

Hydrogen bonding is an especially strong intermolecular force. It is a special type of dipole-dipole force. Hydrogen bonding occurs between molecules that have N-H, O-H, or F-H attractions. Nitrogen, oxygen, and fluorine are very electronegative and draw electrons to itself more strongly than most other atoms in a molecule. This creates an especially strong molecular dipole. This strong molecular dipole gives rise to an especially strong dipole-dipole force between two molecules. Hydrogen bonding is about one-tenth the strength of covalent bonding.

Note: Hydrogen bonding occurs ***BETWEEN*** molecules, *NOT WITHIN* molecules.

Chapter 16: Covalent Bonding
Part 6 - Assignment: Intermolecular Forces

For each of the following, give its shape, its *molecular polarity*, and the intermolecular forces that act upon the molecule.

	Compound	***Shape***	***P or NP?***	***Intermolecular Forces***
1	NaH			
2	CaH_2			
3	CH_2Cl_2			
4	H_2S			
5	C_4H_{10}			
6	H_2O			
7	SF_6			
8	NH_3			
9	SiH_4			
10	PI_5			
11	H_2Se			

Compare each of the following compounds pairs. Circle the one with the greatest boiling point. Explain the reason for your choice.

12. CH_4 or $CHCl_3$

13. CH_4 or C_4H_{10}

14. H_2O or H_2S