Unit 5: Bonding, Molecular Geometry, IMFs

Big Ideas:

Atoms form bonds with other atoms when their valence electrons are shared or transferred. The type of bond that is formed between atoms can be predicted from an elements location on the periodic table. The shape of molecules can be predicted by drawing Lewis Structures and applying VSEPR theory. Intermolecular forces are the attractions that hold molecules together. The type of intermolecular force a molecule has can be determined by its Lewis Structure and shape.

Monday	Tuesday	Wednesday	Thursday	Friday

5-2 Table of Contents, Learning Objectives, Old Skillz

Learning Objectives for Unit 5—Bonding, Molecular Geometry and Intermolecular forces	
E1	I can determine/describe bond type and draw Lewis Structures for compounds.
E2	I can predict the molecular geometry of compounds using VSEPR theory. I can determine the polarity and bond angles of molecules using Lewis structures and VSEPR.
E3	I can determine the intermolecular forces between molecules and predict relative physical states and properties.

Previous Acquired Skills That You Need to be Successful

I can determine the number of valence electrons for an atom.

I can determine the charge of common ions.

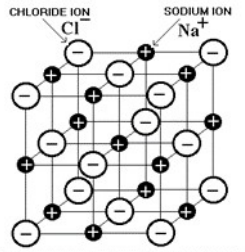
I can compare electronegativities between two atoms.

5-3 Read/Write: Bond Types

The electromagnetic force that hold atoms together are called bonds. There are three major types of bonding patters:

Ionic Bonds Covalent Bonds Metallic Bonds

Ionic Bonds — The ionic bond is formed by the attraction between oppositely charged ions. Ionic bonds can be formed between metals and non-metals. Remember that metal atoms lose one or more valence electrons in order to achieve a stable electron arrangement. When a metal atom loses electrons it forms a positive ion or cation. When non-metals react they gain one or more electrons to reach a stable electron arrangement. When a non-metal atom gains one or more electrons it forms a negative ion or anion. The metal cations donate electrons to the non-metal anions so they stick together in an ionic compound. This means that ionic bonds are formed by the complete transfer of one or more electrons.



SODIUM CHLORIDE LATTICE STRUCTURE

A structure with its particles arranged in a regular repeating pattern is called a crystal. Because opposite charges attract and like charges repel, the ions in an ionic compound stack up in a regular repeating pattern called a **crystal lattice**. The positive ions are pushed away from other positive ions and attracted to negative ions so this produces a regular arrangement of particles where each ion is surrounded by ions of the opposite charge. Each ion in the crystal has a strong electrical attraction to its oppositely charged neighbors so the whole crystal holds together as one giant unit. There are no individual molecules in an ionic compound, just the regular stacking of positive and negative ions. Ionic compounds are represented as **formula units** – the lowest whole number ratio in which ions combine.

At room temperature ionic compounds are high melting point solids. They are usually white except for compounds of the transition metals that may be colored. They are brittle (break easily). They do not conduct electricity as solids, but do conduct electricity when melted (molten) or dissolved in water.

Now turn to page 5-4 in the notes and follow the instructions concerning Ionic Bonds

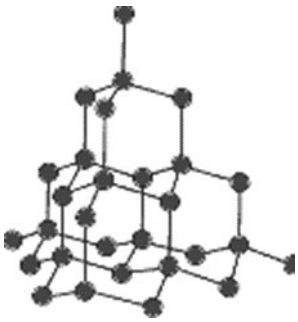
Covalent Bonds — A covalent bond is formed between non-metal atoms. The non-metals are connected by a shared pair of valence electrons forming a **molecule**. A molecule is a neutral particle of two or more atoms bonded to each other. Molecules may contain atoms of the same element such as N_2, O_2, and Cl_2 or they may contain atoms of different elements like H_2O, NH_3, or $C_6H_{12}O_6$.

Covalent substances have low melting points and boiling points compared to ionic compounds or metals. At room temperature, covalent substances are gases, liquids or low melting point solids (think of fats and greases). They do not conduct electricity as solids or when molten and usually do not conduct when dissolved in water.

There are several types of covalent bonds. A **single** covalent bond is formed when two atoms share one pair of valence electrons. A **double** covalent bond is when two atoms share two pairs of valence electrons. A **triple** covalent bond is when two atoms share three pairs of valence electrons

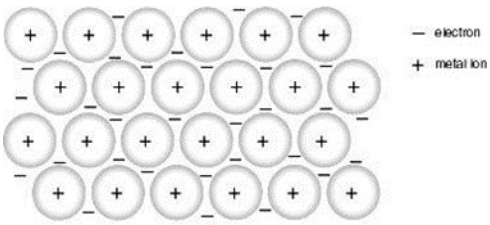
Now turn to page 5-6 in the notes and follow the instructions concerning Covalent Bonds

Network Covalent Bonds —There is one last type of covalent bonding—the bonding in network solids (macromolecules). In this type of bonding, atoms share valence electrons but the atoms are arranged in a regular crystalline pattern in which each atom is covalently bonded to its neighbors in all directions. Therefore, you do not have a collection of small molecules that are easy to separate from each other; the whole system is one giant molecule or a macromolecule held together by this network of strong covalent bonds. Network solids are extremely hard, brittle, solids that do not conduct electricity. Diamonds (a form of pure carbon (see figure)), carborundum (silicon carbide) and quartz (silicon dioxide) are examples of macromolecules.



Now turn to page 5-5 in the notes and follow the instructions concerning Network Covalent Bonds

Metallic Bonds — A metallic bond forms between multiple metal atoms. The metallic bond is formed by the mutual attraction of a positively charged cation to "delocalized" valence electrons. Most metal atoms have only one or two valence electrons and these are not tightly bound to the atoms (low ionization energy). In a piece of metal these valence electrons do not seem to belong to any one of the atoms but are able to move freely through the structure from one atom to another. Metals can be thought of as positive ions (the nucleus and inner shells of electrons—all of the atom except the valence electrons) in a **"sea" of loosely bound valence electrons.** The metal ions line up in a regular repeating pattern (a crystal lattice) and their loose valence electrons move through this crystal acting as electron glue (see figure). Each of the ions is strongly attracted to all of the loose electrons surrounding it so the whole metal holds together as a crystal.



These electrical attractions for the electron glue are strong and hard to break so metals are high melting point crystalline solids. Since there are charged particles free to move metals are good conductors of heat and electricity as solids and as liquids. Because the "electron glue" is free to move, if we hammer or pull the cations to new positions the electron glue flows right along with the cations and holds the structure together in the new position. Thus, metals are malleable (can be hammered into sheets) and ductile (can be stretched into wires) and have a high tensile strength (can be stretched without breaking). Metals also tend to have high luster. This is a direct consequence of loosely bound valence electrons.

Metallic bonding is found in elemental metals and in **mixtures of metals called alloys.**

Now turn to page 5-5 in the notes and follow the instructions concerning Metallic Bonds

5-4 Ionic Bonding, Ionic Compounds

Define the following terms:

ionic bond –

cation –

anion –

crystal lattice–

formula unit –

What are the smallest units of an ionic bond?

List several properties of ionic compounds:

Under what conditions do ionic compounds conduct electricity?

5-5 Metallic Bonding, Network Covalent Compounds

Sea of Electrons

What is a metallic bond? Explain how the ions and electrons are arranged. Pictures are helpful.

List some properties of metallic bonds:

What is an alloy?

What is a network solid?

What type of bonding exists in network solids?

What are some properties of network solids?

What are some examples of network solids?

5-6 Covalent Bonding, Covalent Compounds

Define the following terms:

covalent bond—

Molecule —

single covalent bond —

double covalent bond —

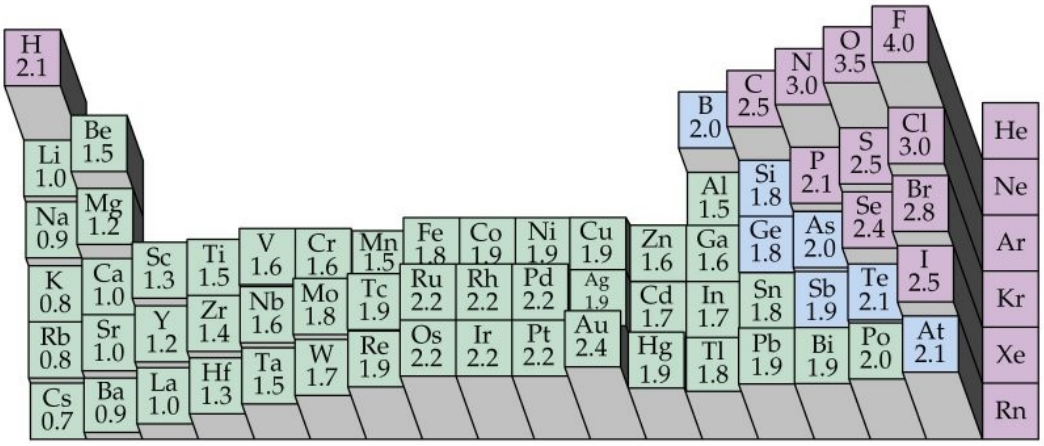
triple covalent bond —

List several properties of covalent compounds:

Valence Bond Theory—

5-7 Polarity of Covalent Bonds

Electronegativity —



Polar Covalent Bond —

Non-Polar Covalent Bond —

Polar Bond vs. Polar Molecule

Looking Ahead: Why might polar molecules want to "stick" to each other? What are some consequences of polar molecules sticking to each other?

5-8 Lewis Structures—How to Draw Covalent Molecules

	Ammonia (NH_3)	Carbonate (CO_3^{2-})	Krypton tetrafluoride (KrF_4)
Step 1: Add up all valence and charge electrons			
Step 2: Draw the skeleton structure with the surrounding atoms connected to the central atom by single bonds.			
Step 3: Add three pairs of electrons to all surrounding atoms except for hydrogen. Never add electrons to hydrogen			
Step 4: Count up the total amount of electrons you have shown and subtract from the total number of valence electrons. Add any leftover electrons as pairs to the central atom.			
Step 5: If all atoms have satisfied valence shells then you are done. If not, move two electrons from a surrounding electron to form a double bond. Repeat if necessary to form a triple bond.			
Step 6: Check to see if it's possible to draw resonance structures. This happens when you can put a double bond in more than one place. You must draw all possibilities.			

5-9 Lewis Structures—Practice

1. PH_3 # Valence electrons =	6. SO_4^{-2} # Valence electrons =	11. SO_3^{-2} # Valence electrons =
2. SO_2 # Valence electrons =	7. OF_2 (O in the middle) # Valence electrons =	12. IF_2^{1-} # Valence electrons =
3. HOOH # Valence electrons =	8. AsF_5 # Valence electrons =	8. $TeCl_4$ # Valence electrons =
4. HCP # Valence electrons =	9. COH_2 (every thing attaches to C) # Valence electrons =	9. H_2O # Valence electrons =
5. SiH_4 # Valence electrons =	10. H_2S # Valence electrons =	10. OH^{1-} # Valence electrons =

5-10 Formal Charge and Tips for Lewis Structures

Formal Charge —

Draw Lewis Structures then assign Formal Charges to all atoms

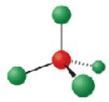
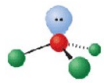
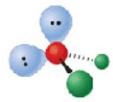
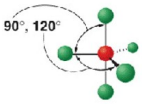
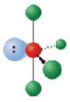
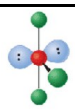
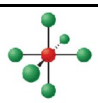
1. PH_3	4. SO_4^{-2}
2. CO_2	5. NO_3^{1-}
3. H_3O^{1+}	6. NH_4^{1+}

Quick and Dirty Rulez for drawing Lewis Structures

5-11 The Shape of Molecules—VSEPR Theory Continued

VSEPR Theory —

5-12 VSEPR Table

Electronic Structure (Family)	Molecular Structure (Geometry)	Areas of e- Density	# of bonds	# of lone pairs	Bond Angles	Polar?	Picture
		2	2	0			
		3	3	0			
		3	2	1			
		4	4	0			
		4	3	1			
		4	2	2			
		5	5	0			
		5	4	1			
		5	3	2			
		5	2	3		*	
		6	6	0			
		6	5	1			
		6	4	2		*	

*Note: Another general rule is that if there any lone pairs on the central atom, the molecule is polar. This is an exception to that rule.

5-13 VSEPR Practice

	PH_3	SF_4	I_3^{1-}	BH_3	O_3
Step 1: Draw the Lewis Structure and verify that it is correct					
Step 2: Count the areas of electron density surrounding the central atom.					
Step 3: Identify the shape.					
Step 4: Determine if the molecule is polar.					
Step 5: Redraw the molecule using wedges and dashes for bonds where appropriate					

5-14 Intermolecular Forces: Attractions Between Molecules

Intermolecular Forces (IMFs)—

Bonds vs. IMFs

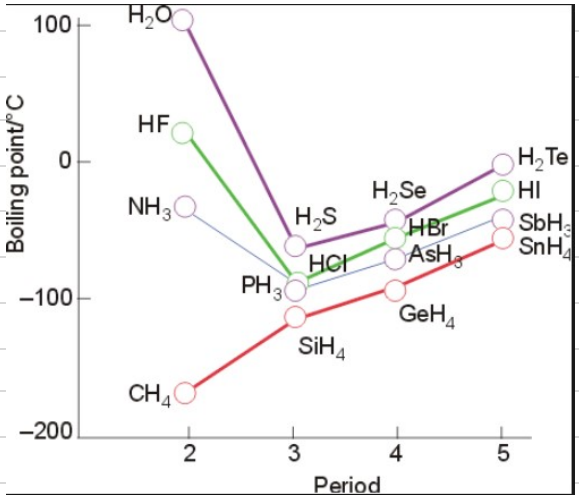
3 Types of IMF's

Dispersion Force

Dipole-Dipole Force

Hydrogen Bond

5-15 Effects of IMFs on Physical Properties



Ex. 1) Compound A has a low molecular weight and exhibits hydrogen bonding and dispersion forces. Compounds B also has a low molecular weight and exhibits only dispersion forces. Which compound is mostly likely a gas at room temperature. Explain your answer.

Ex. 2) Compound W has a low molecular weight and exhibits hydrogen bonding and dispersion forces. Compound B also has a low molecular weight and exhibits only dispersion forces. Which compound is mostly likely a gas at room temperature. Explain your answer.

5-16 Summary