Chemistry Name: _____

Section _____ CHEMICAL BOND KEY Date: _____

Chapters 7 and 8: Chemical Bonding Answer Key

Chemical bond overview

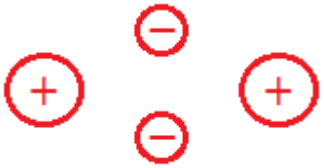
A chemical bond is the force that holds atoms together

Bond formation is always exothermic

Bond breaking is always endothermic

Make a sketch showing how all chemical bond are similar. Show 2 nuclei and 2 electrons.

Explain the sketch: the nuclei of two atoms are mutually attracted to electron density between the two nuclei



Longer bonds tend to be weaker bonds

Shorter bonds tend to be stronger bonds

There are four main types of chemical bonds (in order from strongest to weakest)

Bond Type	Main Characteristics
Metallic	sea of freely moving electrons
Ionic	electrons transferred from one atom to another
Polar covalent	electrons are shared unevenly
Nonpolar covalent	electrons are shared evenly

Metallic bonds are the strongest type because It is the only bond with two positive ions on either side of the electron density. Because positive ions are smaller than atoms, these bonds are the shortest possible bonds and shorter bonds are stronger bonds.

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Ionic bonds are very strong because they have all the strength of any other chemical bond plus the extra strength of a positive ion attracted to a negative ion

Polar covalent bonds are slightly stronger because they have all the strength of any other chemical bond plus the extra attraction of slightly positive and negative atoms

Nonpolar covalent bonds have no extra added strength

Molecules are groups of atoms held together by covalent bonds

The bonds that form molecules can be polar, nonpolar, or a mix of polar and nonpolar covalent bonds

Molecules can be polar or nonpolar based symmetry and electron distribution (Bond polarity may help accentuate polar bonds but are not a requirement.)

There are three main types of intermolecular forces (in order from strongest to weakest)

Force Type	Main Characteristics
Hydrogen bonds	electron pair and a nearly naked proton
Dipole-dipole attractions	polar molecules (one end +, the other –)
van der Waals forces	momentary uneven electron distribution

Hydrogen bonds only occur when hydrogen atoms are covalently bonded to N, O or F atoms. Only N, O, and F have a high enough electronegativity to nearly strip the proton (H atom) of electron density and are small enough to keep the density away from the H

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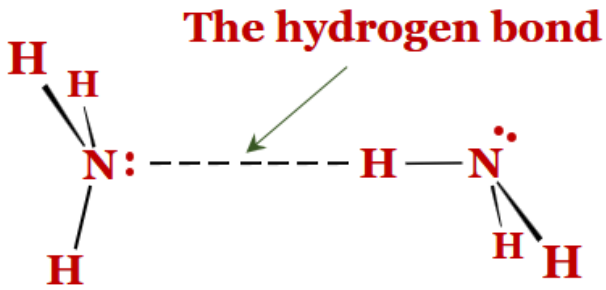
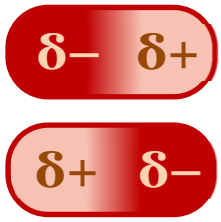
A hydrogen bond is the fairly strong attraction between the electron pair on the N, O, or F on one molecule and the nearly stripped proton of a molecule next door

Make a sketch showing the hydrogen bonding in liquid ammonia, $NH_{3\,(l)}$.



Dipole-dipole interactions only occur when both molecules are polar

Make a sketch showing dipole-dipole interactions between two molecules.



van der Waals forces (also known as London dispersion forces)

only occur between nonpolar molecules

They are the weakest IMF because they only last for a brief time due to constantly shifting electrons

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Ion formation

Positive ions form when atoms lose electrons

Metals tend to lose electrons and form positive ions or cations

because metals have low effective nuclear charges and many electron shells

Metals tend to lose all of their valence shell electrons

Metal cations tend to be much smaller than their atoms because fewer electrons tend to repel less and loss of an electron shell causes atoms to become much smaller

The result is usually a stable octet in the electron shell one closer to the nucleus

Metals usually do not lose electrons from the electron shell below the valence shell because the electrons are much closer to the nucleus than the old electron shell and the effective nuclear charge is much stronger in the new shell (increases by 8)

Negative ions form when atoms gain electrons

Nonmetals tend to gain electrons and form negative ions or anions

because nonmetals have high effective nuclear charges and fewer protons in the nucleus

Nonmetals form a stable octet by completely filling the s and p sublevel in the valence electron shell

Nonmetal anions tend to be larger than their atoms because adding more electrons in an electron shell increases the amount of repulsion

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Nonmetals do not gain electrons past the stable octet because more electrons would have to go in the electron shell one farther from the nucleus where the effective nuclear charge is 8 lower (and usually negative)

The most likely metal ions can be predicted by looking at their electron configuration

and removing all the valence shell electrons

The electron configuration for $_{50}Sn$ is 2 – 8 – 18 – 18 – 4

and its most likely ion is Sn^{4+}

The most likely nonmetal ions can be predicted by looking at their electron configuration

and adding enough electrons to bring the valence shell up to 8 (or up to 2 for hydrogen)

The electron configuration for $_{34}Se$ is 2 – 8 – 18 – 6

and its most likely ion is Se^{2-}

Using electronegativity to determine bond type

Electronegativity Difference and Bond Type	
Bond Type	Δ EN
Hydrogen bond	≥ 1.7
Polar covalent bond	0.4 to 1.6
Nonpolar covalent bond	≤ 0.3

Determine the electronegativity difference (Δ EN) and bond type for:

NaCl		3.2 – 0.9 = 2.3	ionic bond
HCl		3.2 – 2.2 = 1.0	polar covalent bond
PH_3		2.2 – 2.2 = 0	nonpolar covalent bond
HF	(an exception)	4.0 – 2.2 = 1.8	very polar covalent bond

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Factors that affect bonds and determine the state of matter (solid, liquid, or gas)

Temperature – high temperatures favor gases

Bond strength (or type) – chemical bonds tend to form solids at room temperature, hydrogen bonds usually form liquids (H_2O), dipole-dipole and van der Waals tend to form gases

Molar mass – high mass molecules move slowly and tend to form liquids and solids while low mass molecules tend to form gases

Shape – smaller spherical molecules tend to be gases (neopentane) while long molecules tend to be liquids or solids (n-pentane)

Polarizability – high temperatures favor gases

The four types of solids and their properties

Metallic solids – characterized by a sea of freely moving electrons

Conductivity: good for both the solid and the liquid state

Malleability: easily hammered into thin sheets (also ductile or drawn into wires)

Solubility: not soluble in either polar or nonpolar solvents

Hardness: Tend to be very hard (W = 7.5 to Na = 1.5)

Melting Point: Tend to be very high (W = 3422ºC but Hg = −38.8 ºC)

Boiling Point: Tend to be high (W = 5930 ºC but Hg = 356.7 ºC)

Luster: Metallic luster (bright, shiny, and reflective)

Ionic solids – characterized by a lattice network of + and − ions

Conductivity: poor for the solid but good in the liquid state or in water solution

Malleability: Not malleable – brittle with tendency to cleave

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Solubility: soluble in polar solvents (like water) not soluble in nonpolar solvents

Hardness: Tend to be very hard (Al_2O_3 = 9 and NaCl = 2.5)

Melting Point: Tend to be high (Al_2O_3 = 2072 ºC and NaCl = 801 ºC)

Boiling Point: Tend to be high (Al_2O_3 = 2977 ºC and NaCl = 1413 ºC)

Luster: Vitreous luster (glassy, not metallic)

Network solids – characterized by a lattice of covalent bonds (forms a single 'molecule')

Conductivity: good insulators

Malleability: not malleable, brittle and will cleave

Solubility: not soluble in either polar or nonpolar solvents

Hardness: Tend to be very hard (diamond = 10 to quartz = 7)

Melting Point: Tend to be high (quartz = 1670 ºC BP = 2230 ºC)

Boiling Point: Tend to be high and decompose (diamond = 3550 ºC decomposition)

Luster: Vitreous luster (glassy, not metallic)

Molecular solids – characterized by a lattice of intermolecular forces (IMF)

Conductivity: good insulators

Malleability: range from tendency to clear to soft (waxy)

Solubility: soluble in like solvents (polar for ionic and polar, nonpolar for nonpolar)

Hardness: Varies from hard (ice) to soft (wax)

Melting Point: Tend to be low (water = 0 ºC and N_2 = −210 ºC)

Boiling Point: Tend to be high (water = 100 ºC but Hg = −195.8 ºC)

Luster: Vitreous luster (glassy, not metallic)

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Bonding and Solubility

Metals and network solids are generally: insoluble

Ionic and molecular solids are generally: insoluble

Nonpolar solvents can dissolve nonpolar substances because only other nonpolar molecules can form the momentary dipoles required for van der Waals forces

Nonpolar solvents cannot dissolve polar or ionic substances because permanent dipoles cannot form the momentary and shifting van der Waals forces

Polar solvents can dissolve polar and ionic substances because opposing forces can attract

The solubility rule is: like dissolves like

Solubility in water can be determined by referring to: Table F in the 2011 reference tables

Show the solubility of the following substances in the balanced equation below by filling in either (s) or (aq) for each reactant and product.

$$Na_2SO_{4\,(aq)} + BaCl_{2\,(aq)} \rightarrow 2\ NaCl_{(aq)} + BaSO_{4\,(s)}$$

Lewis Electron Dot (LED) notation, formulas, and structures

Write LED notations for the following:

Write LED formulas for the following:

Write LED structures for the following:

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Guidelines for LED formulas and structures:

Hydrogen will almost never be the central atom

Carbon will almost always be the central atom

Most elements (except H which can only have 2 e^- in its valence shell) will form a

stable octet

A few elements tend to remain electron deficient, including: Be, B, and Al

Other electron deficient elements tend to form multiple bonds

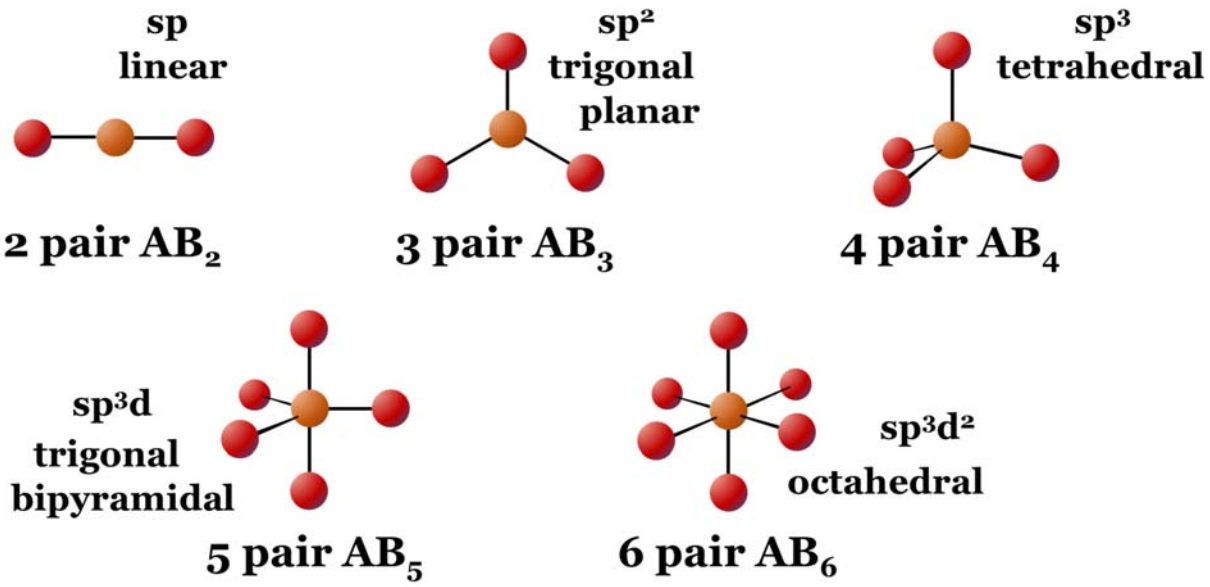
Show the LED structure for CO_2:

Elements that can form extended octets are found in Period 3 or higher

Show the LED structure for PF_5:

Valence Shell Electron Pair Repulsion Theory (VSEPR)

In the space below, make sketches of the molecular geometries VSEPR predicts. Include the number of electron pairs, the name of each geometry, and the hybridization for each sketch



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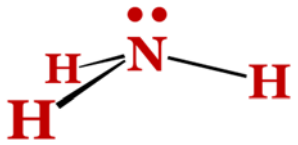
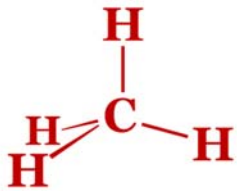
Draw a 3-D rendering of an NH_3 molecule in the space below



Molecular polarity from VSEPR

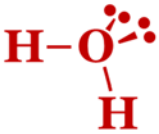
Symmetrical molecules will be nonpolar

Draw three LED structures that exemplify nonpolar molecules.



Asymmetrical molecules will be polar

Draw three LED structures that exemplify polar molecules.



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