**Ch 10 Chemical Bonding I: The Lewis Model Name:**

Read the quote at the top of 380 and the paragraph at the bottom of the page.

*10.1 Bonding Models and AIDS Drugs (very interesting read on how protease inhibitors were developed)*

What do bonding theories help to explain and predict?

*10.2 Types of Chemical Bonds*

Why do chemical bonds form?

Summarize the main differences between ionic and covalent bonding. Examine the classic cmpds in *Figure 10.1.*

Metals tend to have few loosely bound valence electrons in their atoms which become “pooled” (shared over the lattice) so they are \_\_\_localized over the entire metal. (electron \_\_\_\_ model)

*10.3 Representing Valence Electrons with Dots*

The ­­­­\_\_\_\_\_\_\_\_\_\_\_ symbols for the period 2 elements have dots placed \_\_\_\_\_\_ly before pairing.

(This pattern is useful for bonding.)

Which group of atoms on the pd table would have octets (except He which has a \_\_\_\_\_\_)? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

In the Lewis model a chemical \_\_\_\_\_\_\_\_\_ is the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ or \_\_\_\_\_\_\_\_\_\_\_\_\_\_ of electrons to attain a

\_\_\_\_\_\_\_\_\_\_\_\_ electron configuration which is the \_\_\_\_\_\_\_\_\_\_\_\_ rule. *CC10.1 \_\_\_\_*

*9.4 Ionic Bonding: Lewis Symbols and Lattice Energies CC10.2 \_\_\_\_*

Notice how both K and Cl have obtained octets. Cl by \_\_\_\_\_\_\_ing an electron and K by \_\_\_\_ing an electron. K+ and Cl− have a econfig which is isoelectronic with which Noble gas? \_\_\_\_

*Note:* I prefer that both the cation and ion be shown with brackets. Example: [ K ]+

*Practice 10.1 (use brackets for both ions)*

*Lattice Energy: The Rest of the Story*

The equation given at the top of pg 399 shows that the formation of solid sodium chloride is highly \_\_\_\_thermic.

Based the IE1 of Na and the EA of Cl why should the reaction be endothermic?

The reaction is exothermic due to \_\_\_\_\_\_\_\_\_\_\_\_energy. According to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ law as cations are

attracted to anions the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ energy between them \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_s. That energy is emitted as \_\_\_\_\_.

Examine *Figure 10.4* and the description of each step in the \_\_\_\_\_-\_\_\_\_\_\_ Cycle. Notice that steps 1-3 are all

\_\_\_\_thermic and only step 4 and 5 are \_\_\_thermic. Since the lattice energy change(5) is so large the overall formation

of NaCl(s) from its elements is \_\_\_\_thermic due to the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ between the ions.

*Note: In AP you should be able to identify what is happening at each step and why it is endo or exo.*

The figure near the right middle of pg 401 shows the bond lengths for the group 1A metal \_\_\_\_\_\_ides. Notice the charges are always +1 and −1. Why does the lattice energy decrease down the group?

The figure below shows the formation of 2 ionic cmpds with similar \_\_\_\_\_\_\_\_\_\_\_\_s between ions.

Why is the lattice energy for CaO so much larger than NaF?

*Practice and More Practice 10.2*

BRIEFLY summarize how our model for ionic bonding accounts for the following properties.

High MPs & BPs –

Not conductive as solids –

Conductive when dissolved in water – *CC 10.3 \_\_\_\_*

Read *Chemistry in Medicine: Ionic Compounds in Medicine.* Good to be familiar with common cmpds in our lives.

*10.5 Covalent Bonding: Lewis Structures*

How do Lewis structures depict covalent bonding in molecular compounds?

Look at the Lewis structure for water on pg 404. Notice that the H atoms have been put at opposite ends of the oxygen atom. (You know that this representation does not show the true SHAPE of a water molecule.)

Why does oxygen form 2 bonding pairs? Why are there 2 lone pairs?

Bonding pairs are often represented by a \_\_\_\_\_\_\_\_\_\_ to emphasize the formation of a chemical \_\_\_\_\_\_\_. (These are called Lewis dot-dash structures. Others texts show all pairs of e’s as dashes to simplify drawing the structure.)

Why do oxygen molecules contain a double bond while nitrogen molecules have a triple bond?

The formation of H3O+ (the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ ion) can be explained by the Lewis model if the oxygen

\_\_\_\_\_\_\_\_\_s an electron to form an \_\_\_\_\_\_\_\_. Note how hydrogen peroxide can be modeled.

The Lewis model accounts for why \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ bonds are highly \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ as the electrons must

be \_\_\_\_\_\_\_\_\_\_\_ed in the space between atoms. In contrast \_\_\_\_\_\_\_\_\_\_ bonds are \_\_\_\_directional hold together an

\_\_\_\_\_\_\_\_\_\_ of atoms (crystal lattice of charged particles tightly held together). The fundamental unit of covalently

bonded compounds are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_s. The interactions (attractions) between molecules

(\_\_\_\_\_molecular forces) are much \_\_\_\_\_\_\_\_er than the bonds within the molecule (\_\_\_\_\_molecular forces). Because

of this how do MPs and BPs of ionic and covalent cmpds compare?

*CC 10.4 \_\_\_\_*

*10.6 Electronegativity and Bond Polarity*

What is the limitation of the Lewis model for hydrogen fluoride? Draw how it can be addressed.

Why is this NOT an ionic bond?

A \_\_\_\_\_\_\_\_\_ bond has a \_\_\_\_\_\_\_\_\_\_\_\_\_\_ pole and a \_\_\_\_\_\_\_\_\_\_\_\_\_ pole. The molecules will align with an

external \_\_\_\_\_\_\_\_\_\_\_\_\_\_ field. MOST (basically ALL for AP) covalent bonds between \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

atoms are actually \_\_\_\_\_\_\_\_\_\_\_\_\_ covalent.

Define electronegativity (EN):

Linus \_\_\_\_\_\_\_\_\_\_\_\_\_ (look up connection to Watson & Crick) assigned a EN value of \_\_\_\_ to \_\_\_\_\_\_\_\_\_\_\_\_\_\_ the

most electronegative element. Two alkali metals \_\_\_\_ & \_\_\_\_ have the lowest value of \_\_\_\_.

EN generally \_\_\_\_\_\_\_\_\_\_\_\_\_ across a period and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ down a group. In general it is

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ related to atomic \_\_\_\_\_\_\_ (larger atom has less ability to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ electrons). Law??

The degree of polarity in a bond depends on the ΔEN (electronegativity \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_). The \_\_\_\_\_\_\_\_\_\_\_\_\_

the difference the more \_\_\_\_\_\_\_\_\_\_\_ the bond. The bond in Cl2 is pure \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ or \_\_\_\_\_ polar.

Typically there is a \_\_\_\_\_\_\_\_\_\_ difference between a \_\_\_\_\_\_\_\_\_\_\_\_\_ and a non\_\_\_\_\_\_\_\_\_ so they form an \_\_\_\_\_\_\_\_

bond (electron is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_). *Table 10.1* shows a difference of 0.4 to be NONPOLAR covalent, however the

only true NP bond would be between 2 of the same atoms (pure). A difference of 2.0 (1.7 in other texts) is considered

to be \_\_\_\_\_\_\_\_\_\_\_\_\_. Therefore any difference between 0 and 2.0(1.7) is considered to be \_\_\_\_\_\_\_\_\_\_\_\_\_ covalent.

*Figure 10.9* makes it appear as if are distinct cut-offs but there is much “gray area”. *Figure 9.10* shows the % \_\_\_\_\_\_\_

character vs EN difference. NO bond is \_\_\_\_% ionic (>\_\_\_\_\_% is used as a guide to be ionic).

Dipole \_\_\_\_\_\_\_\_\_\_\_(µ) can be calculated for bonds (and measured for molecules). AP questions often give data such

as *Table 10.2* in order to examine changes in polarity of bonds and molecules.

*Practice 10.3* a) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ b) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ c) ­­\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

What would you classify the bond between H and S to be? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (CAREFUL!)

*10.7 Writing Lewis Structures for Molecular Compounds*

1) Skeletal structure: H atoms are always \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_(at the \_\_\_\_\_s). Put more \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ atoms

in \_\_\_\_\_\_\_\_\_\_\_\_\_\_ positions and \_\_\_\_\_\_\_ EN atoms (other than H) in \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ position.

2) Calculate \_\_\_\_\_\_\_\_\_\_\_\_ # of \_\_\_\_\_\_\_\_\_\_\_\_\_\_s for molecule by summing \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ electrons of each

atom. For \_\_\_\_\_\_\_atomic ions add electrons for \_\_\_\_\_\_\_\_\_\_\_\_\_\_ charge and subtract for \_\_\_\_\_\_\_\_\_\_\_\_\_.

3) Distribute electrons giving \_\_\_\_\_\_\_\_s (or duets for H). Place 2 electrons \_\_\_\_\_\_\_\_\_\_\_\_ atoms.

(For a minimum of \_\_\_bond.) Then add \_\_\_\_\_ pairs to *terminal* atoms and then to central atom to form octets.

4) If atoms lack an octet, form \_\_\_\_\_\_\_\_\_ or \_\_\_\_\_\_\_\_\_\_\_ bonds as necessary by moving lone pairs from

*\_\_\_\_\_\_\_\_\_\_\_\_* atoms into bonding region with the central atom. *(see example 10.4)*

*Practice 10.4 Practice 10.5 Practice 10.6*

*10.8 Resonance and Formal Charge*

Ozone (O3) has two valid \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ structures. In nature, the molecule exists as an \_\_\_\_\_\_\_\_\_\_\_\_ of the

two structures with equivalent \_\_\_\_\_\_\_s that are \_\_\_\_\_\_\_\_mediate in strength and length between a \_\_\_\_\_\_\_\_\_\_\_

bond and a \_\_\_\_\_\_\_\_\_\_\_\_bond. This can be represented by placing a \_\_\_\_\_\_\_\_\_\_\_-headed arrow between the 2

structures. (This does NOT mean that both structures exist and the molecule constantly switches back and forth. Only

the resonance \_\_\_\_\_\_\_\_\_\_d structure exists.) The image of the hybrid structure in *Figure 9.11* shows that the electrons

are \_\_localized over the two bonds which \_\_\_\_\_\_\_\_\_\_s their \_\_\_\_\_\_\_\_\_\_y and stabilizes the structure.

*Practice 10.7*

In some cases resonance structures are NOT \_\_\_\_\_\_\_\_\_ent, \_\_\_\_\_\_\_\_\_ charge can be used to determine the better one.

Formal charge of an atom is the charge it would have if all \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ electrons were shared \_\_\_\_\_\_\_\_\_\_\_\_\_\_.

***Read how formal charge is calculated and the 4 rules which apply.*** An easier formula to use is as follows:

Formal charge = valence electrons – assigned electrons You know how to determine the number of valence electrons for an atom by looking at its group # on the pd table. To determine “assigned” electrons: count electrons in lone pairs around an atom and add ½ the number of bonding electrons. *In example at bottom of pg 415:A better than B*

*Follow Example 10.8 closely!*

*Practice 10.8 More Practice 10.8 (notice all 3 are “equal”)*

*10.9 Exceptions to the Octet Rule: Odd-Electron Species, Incomplete Octets, and Expanded Octets*

What are free radicals? *(also Google the term to see how they are related to our health)*

NO and NO2 both are common air pollutants. Draw the other possible structure for NO by placing the unpaired e− on oxygen and then show how the other structure is better based on formal charge.

*CC10.5 \_\_\_\_\_\_ Why?*

Read *Chemistry in the Environment: Free Radicals and the Atmospheric Vacuum Cleaner*

How are hydroxyl radicals formed and why are they so important?

Draw a Lewis structure for BeH2.

Why doesn’t B form a double bond in BF3?

B can complete its octet by reaction with NH3 when the \_\_\_\_\_\_ pair in NH3 provides both \_\_\_\_\_\_\_\_\_\_\_\_s for the

bond which is called a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ covalent bond. (middle left of pg 420)

Elements in the \_\_\_ row of the pd table and beyond often exhibit \_\_\_\_\_\_\_\_\_\_ octets of up to \_\_\_(sometimes 14)

electrons. Elements such as \_\_\_\_ and \_\_\_ shown in the examples form expanded octets because the \_\_\_ orbitals are

accessible for bonding. Expanded octets never occur in \_\_\_ period elements. *(Remember this is a bonding THEORY.)*

Why is the 2nd structure of sulfuric acid better than the 1st?

*Practice 10.10 More Practice 10.10*

*10.10 Bond Energies and Bond Lengths*

Define bond energy:

***Bond energies are always \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, because it \_\_\_\_\_\_\_\_\_\_\_ takes \_\_\_\_\_\_\_\_\_\_\_ to \_\_\_\_\_\_\_\_\_\_\_ a bond!***

Why do you think the bond energy for HCl is greater than the energy for Cl−Cl? (Hint: Coulomb’s law)

What makes N2 so stable?

Why are the values in *Table 10.3* “average” bond energies?

Review the example calculation on pg 423 for the given reaction. Note which bonds are broken and which are formed.

The equation: Δ*Hrxn* = Σ(Δ*H’s bonds broken) +* Δ*H’s bonds formed)* works because Δ*H* is \_\_\_\_\_\_\_\_\_\_\_ for

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ bonds and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ for forming them. A reaction is exothermic when the bond

energies formed > bond energies broken and endothermic when bonds formed< bonds broken.

It has been said that “\_\_\_\_\_\_\_\_\_\_\_\_ is stored in \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_bonds” which makes it sound as if \_\_\_\_\_\_\_\_ing

bonds \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ energy. HOWEVER, \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ a chemical bond always \_\_\_\_\_\_\_\_\_\_\_\_ energy!

It is always the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ of chemical bonds that \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ energy!

*CC 10.10 \_\_\_*

*Practice 10.11 More Practice 10.11*

In general \_\_\_\_\_\_\_\_\_\_\_\_ bonds are the \_\_\_\_\_\_\_\_\_est and \_\_\_\_\_\_est bonds, while \_\_\_\_\_\_\_\_\_\_\_\_\_ bonds are the

\_\_\_\_\_\_\_\_\_est and \_\_\_\_\_\_est bonds. Notice that, as the bond gets \_\_\_\_\_\_er, it also becomes \_\_\_\_\_\_er. But does not

always hold as seen by the table of values for nitrogen-\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ single bonds.

*10.11 Bonding in Metals: The Electron Sea Model*

Metals have a tendency to \_\_\_\_\_\_\_\_ electrons due to \_\_\_\_\_ IE values. When metal atoms to form a solid, each atom

\_\_\_\_\_\_\_\_\_\_\_ one or more electrons to an \_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_. EX: sodium metal is an \_\_\_\_\_\_\_\_\_\_ of

\_\_\_\_\_\_\_\_\_\_\_\_ly charged Na+ \_\_\_\_\_\_ immersed in a sea of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ charged e−‘s. \_\_\_\_\_\_\_\_\_\_\_\_\_tion

between them is how the sodium \_\_\_\_ions are held together. What properties are explained by this model?

Read *Chemistry in the Environment: The Lewis Structure of Ozone Click this* [*link*](https://chem.libretexts.org/Bookshelves/Environmental_Chemistry/Supplemental_Modules_(Environmental_Chemistry)/Atmospheric_Chemistry/Ozone#:~:text=The%20visible%20region%20range%20from,bond.)

The diagram does not show O2 being broken by UV to form single O’s which combine with O2 to form O3.

*Exercises* (pgs 431-434) #39, 43, 45, 65, 67, 81, 82, 93, 97, 103, 116 75(a,c,d), 76(a,b,c)