**Ch 11 Chemical Bonding II: Molecular Shapes, VBT, and MOT Name:**

*11.1 Morphine: A Molecular Imposter*

How does morphine work in humans? Why is it an imposter?

*11.2 VSEPR Theory: The Five Basic Shapes*

What is VSEPR theory based on?

The preferred geo\_\_\_\_\_\_ of a molecule is the one in which the electron \_\_\_\_\_\_\_\_\_\_ have the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

separation (minimum \_\_\_\_\_\_\_\_\_\_\_) possible.

2 Electron Groups: \_\_\_\_\_\_\_\_ and \_\_\_\_\_ both have 2 electron groups about the \_\_\_\_\_\_\_\_\_\_ atom and both have

\_\_\_\_\_\_\_\_\_\_\_ geometry.

3 Electron Groups: \_\_\_\_ and CH2O (\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_) both have \_\_\_\_\_\_\_\_\_\_\_\_\_\_ geometry which has \_\_\_\_o

bond angles (the H−C−H angle is slightly \_\_\_\_\_\_\_er due to the higher electron \_\_\_\_\_\_\_\_\_\_\_\_ of the double bond

and greater \_\_\_\_\_\_\_\_\_\_\_\_\_\_ on the single bonds. *CC 11.1 \_\_\_\_*

4 Electron Groups: 4 balloons tied together assume a 3-d \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ geometry with \_\_\_\_o bond angles

\_\_\_\_\_\_\_ (methane) is an example. *CC 11.2 \_\_\_\_*

5 Electron Groups: 5 groups assume a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ geometry where \_\_ groups lie in a single

\_\_\_\_\_\_\_\_\_ (\_\_\_\_\_\_\_\_\_\_ial positions at \_\_\_\_o) while the other 2 are \_\_\_\_\_\_\_\_ and below the plane (\_\_\_\_\_\_\_

positions at \_\_\_o to the equatorial atoms). \_\_\_\_\_\_ (phosphorus \_\_\_\_\_\_chloride) is an example.

6 Electron Groups: 6 groups assume a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ geometry \_\_ groups lie in a single \_\_\_\_\_\_\_\_ with

the other 2 above and below with all angles at \_\_\_o. \_\_\_\_\_ (sulfur \_\_\_\_\_\_\_\_\_\_\_ide) is an example.

*Practice 11.1 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_*

*11.3 VSEPR Theory: The Effect of Lone Pairs*

4 Electron Groups with Lone Pairs: NH3 (\_\_\_\_\_\_\_\_\_\_\_\_\_) has 4 e− groups (\_\_lone \_\_\_\_\_ & \_\_ bonding pairs). The

electron \_\_\_\_\_\_\_\_\_\_\_\_ is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ but the geometry of the atoms is \_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_.

The lone pair exerts \_\_\_\_\_\_\_\_\_\_\_\_\_ repulsion than the \_\_\_\_\_\_\_\_\_\_\_ electrons causing the H−N−H bond angle to

to compress from 109.5o to \_\_\_\_o. (A lone pair occupies more \_\_\_\_\_\_\_\_ than a \_\_\_\_\_\_\_\_\_\_\_ pair. *CC 11.3 \_\_\_\_*

Water also has \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ electron domain geometry and \_\_\_\_\_ molecular geometry (\_\_lone pairs) with a

bond angle of \_\_\_\_\_\_o. *Repulsions: most lp-lp > \_\_\_-\_\_\_ > \_\_\_-\_\_\_ least Note the bond angles in Figure 11.4*

5 Electron Groups with Lone Pairs: SF4 has \_\_lp & \_\_bp with \_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_ e− geometry. The lp

occupies an \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ position to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ repulsions resulting in a molecular

geometry called \_\_\_\_\_\_\_\_\_\_\_. In BF3 the 2 lps occupy 2/3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_ positions resulting in

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ molecular geometry. XeF2 has all \_\_ lps occupying \_\_\_\_\_\_\_\_\_\_\_\_\_\_ positions resulting in

\_\_\_\_\_\_\_\_\_\_\_\_\_ geometry.

6 Electron Groups with Lone Pairs: BrF5 has \_\_lp & \_\_bp with \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ e− geometry and

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ pyramidal molecular geometry. XeF4 has \_\_lp & \_\_bp. The lps occupy occupy opposite

positions (since all positions are equal) resulting in \_\_\_\_\_\_\_\_\_\_\_\_ planar molecular geometry.

*Table 11.1* is a nice summary of the electron and molecular geometries for each electron group.

*CC11.4 \_\_\_\_ CC11.5 \_\_\_\_\_*

*11.4 VSEPR Theory: Predicting Molecular Geometries*

Review Examples 11.2 & 11.3 Complete *Practice 11.2 & 11.3* (Include drawings of the Lewis Structures)

Examine the “Representing Molecular Geometries on Paper” at the top of pg 449. (I will give you a handout that uses this technique but it also includes showing the lone pairs.

Note how the Lewis diagram for Glycine at the bottom of 449 makes it appear that all bond angles are 90o or 180o.

Compare this to the “ball and stick” diagram below. You must be able to determine correct angles for these large molecules with multiple “center atoms” by looking at Lewis diagrams. *(up to 4 e−groups for AP)*

*11.5 Molecular Shape and Polarity*

If a diatomic molecule has a \_\_\_\_\_\_\_\_ bond the entire \_\_\_\_\_\_\_\_\_\_\_\_\_\_ will be polar. (see diagrams)

What determines if a polyatomic molecule will be polar?

CO2 and H2O both contain polar bonds. Why is CO2 a non-polar molecule while H2O is a polar molecule?

*Summarizing: Basically, if a molecule contains polar bonds and it is not symmetric the molecule will be polar.*

On pg 452 it is shown how to add vectors (you will learn more in Physics). This is useful for finding the direction of the overall dipole moment in a molecule as shown in *Table 11.2.* Note how dipoles can CANCEL.

*Practice 11.5 (sketch the molecule)*

Read the top of pg 454. This is a good comparison polar and nonpolar. We will learn more on this in future chapters.

Read *Chemistry in Your Day: How Soap Works*

*11.6 Valence Bond Theory: Orbital Overlap as a Chemical Bond*

According to VBT, \_\_\_\_\_\_\_\_\_\_\_\_ reside in \_\_\_\_\_\_\_\_\_\_\_\_\_s (s,p,d,f) localized on atoms or the orbitals are

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ atomic orbitals (a \_\_\_\_\_\_\_\_\_\_\_ of 2 or more standard orbitals).

Why do bonds form (or not form) according to this theory?

Closely examine and understand the changes shown in *Figure 11.6* when 2 H atoms approach each other.

What determines the shape of the molecule in this theory?

Review the example for the formation of H2S.

Why is VBT a better explanation for this molecule than VSEPR? *CC 11.8 \_\_\_\_*

*11.7 VBT: Hybridization of Atomic Orbitals*

What does VBT predict for bonding between hydrogen and carbon? What do we know from experimental data?

VBT treats e−‘s as if they occupied \_\_\_\_\_\_\_\_\_\_\_\_\_ atomic \_\_\_\_\_\_\_\_\_\_\_\_s which is an over\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

The concept of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_recognizes that orbitals in a \_\_\_\_\_\_\_\_\_\_\_\_\_\_ are not necessarily the same

as in an \_\_\_\_\_\_\_\_\_\_\_. Atomic orbitals are mathematically combined to form \_\_\_\_\_\_\_\_\_\_\_ orbitals which

correspond more closely to the actual \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ of e−‘s in bonded atoms. They are still \_\_\_\_\_\_\_\_\_\_\_\_

on individual atoms but have different \_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_. e− density is concentrated in a directional

lobe and hybrid orbitals \_\_\_\_\_\_\_\_\_\_\_\_ energy of the molecule by \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ orbital \_\_\_\_\_\_\_\_\_\_\_\_\_.

In general, the more \_\_\_\_\_\_\_\_\_ an atom forms, the greater tendency to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. Terminal atoms

have the \_\_\_\_\_\_\_\_\_\_ tendency to hybridize. Hybridization is important in \_\_\_\_\_\_\_\_\_\_\_ which forms \_\_ bonds.

How can we account for the tetrahedral geometry in methane by hybridization?

*Figure 11.7 - note how the sp3 orbitals are formed from one s and 3 p orbitals.* *CC11.9 \_\_\_\_*

Does the hybridization model of methane agree with VSEPR?

How is the sp3 hybridization of nitrogen in ammonia different?

In the molecule H2CO (formaldehyde), \_\_\_\_\_\_\_\_\_\_\_\_ is the central atom. The carbon forms \_\_\_\_ hybrid orbitals to

and one \_\_ orbital remains \_\_\_hybridized. This orbital overlaps with a p orbital from \_\_\_\_\_\_\_\_\_\_\_\_ to form a

\_\_\_\_\_\_\_\_\_\_\_\_\_ bond (this bond contains one σ bond and one π bond). A σ bond forms when orbitals overlap

\_\_\_\_ to \_\_\_\_ while the π bond forms when unhybridized \_\_ orbitals overlap \_\_\_\_\_ by \_\_\_\_\_. A \_\_\_\_\_\_\_\_\_\_ bond

in a \_\_\_\_\_\_\_\_\_model always corresponds to one \_\_ and one \_\_ bond in VBT. How do the bonds compare

in terms of strength?

How do 1,2-dichloroeth***a***ne and 1,2-dichloroeth***e***ne differ in terms of rotation around the central bond?

What does this restriction lead to for 1,2-dichloroeth***e***ne? *(bottom of pg463)* What do the prefix designations mean?

Read *Chemistry in Your Day: The Chemistry of Vision* pg 464 Answer the question: \_\_\_\_\_ *CC 11.10 \_\_\_\_*

For C2H2 (\_\_\_\_\_\_\_\_\_\_\_\_ene) the carbon atoms form \_\_\_\_\_\_ hybrid orbitals. The triple bond contains one \_\_\_bond

and two \_\_\_bonds *(bottom of pg 465)*

*CC11.11 \_\_\_\_*

***Note:*** You can easily determine the hybridization of an atom by counting the number of bonds (single, double, or triple) that it forms. If an atom forms 4 bonds it has sp3 hybridization. 3 bonds is \_\_\_\_. 2 bonds is \_\_\_\_.

Since arsenic pentafluoride forms 5 bonds, VBT involves the hybridization of \_\_\_ orbitals. The arsenic atom

forms \_\_\_\_\_\_\_ hybridization. The sulfur atom in SF6 forms \_\_\_\_\_\_\_\_\_ hybridization.

***Note:*** Evidence exists that hybridization of the d orbitals does NOT occur. It will not be tested on the AP exam.

*Table 11.3* shows how the shapes determined by VSEPR correspond to hybridization.

*Practice 11.6 Practice 11.7*

*Practice 11.8 More Practice 11.8*

*11.8 Molecular Orbital Theory*

How does MOT differ from VBT? *Note: AP does not test MOT but it helps explain limitations of VBT*

What is an LCAO? How are they different than hybridized orbitals?

When H2 bonds, the 1s orbitals can combine constructively to form the σ1s (sigma 1s) molecular \_\_\_\_\_\_\_\_\_\_\_\_\_

which is \_\_\_\_\_\_\_\_\_\_ in energy than the 1s \_\_\_\_\_\_\_\_\_\_ orbitals so it is a \_\_\_\_\_ing orbital. A σ\*1s orbital can also

form when one of the 1s orbitals is in the opposite \_\_\_\_\_\_\_\_\_\_. The different phases results in \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

interference . The \* indicates this is an \_\_\_\_\_bonding orbital with \_\_\_\_\_\_\_\_\_\_\_\_\_ energies than the 1s orbitals.

Examine the MO energy diagram for H2 at the top of pg 473. The diagram shows why H2 forms since the electrons can move into a lower energy MO bonding orbital. *Note how bond order is calculated.*  What does a positive bond order mean compared to a zero or negative bond order? In general, what does higher bond order indicate?

Does it make sense that He does not form a diatomic molecule? Notice He2+ can be stable.

*Note:* I will give you a worksheet to complete which goes with the rest of this section.

Why is MOT a better way to represent the bonding in O2 than Lewis of VBT?

*OPTIONAL: Exercises* (pgs 486-7) 33, 35, 37, 39, 41, 43, 67