

Chem E-1a
Friday Review Notes
Chapter 10: Chemical Bonding II

VSEPR, Geometry, Hybridization, and Bond Angles:

- First draw a complete Lewis Structure, and count the number of Electron Domains on the central atom(s).
- Then see the “Complete Table of Hybridization and Geometry” in this packet.
- Make sure you know all of the information in this table.

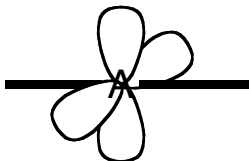
Drawing Molecules to Indicate Molecular Geometry:

- Lewis Structures are not intended to represent the actual geometry of a molecule.
 - Bonds in Lewis Structures should only be straight lines.
 - In this class, do not use “dashes” or “wedges” in your standard Lewis Structures.
- If you are asked to draw a “Perspective Sketch” of the geometry of a molecule, then you should use “straight lines”, “dashes”, and “wedges” to clearly indicate the 3-dimensional geometry of the molecule.
 - See the “Complete Table of Hybridization and Geometry” for line, dash, and wedge depictions of the standard molecular geometries.

Drawing the σ and π Bonding in a molecule:

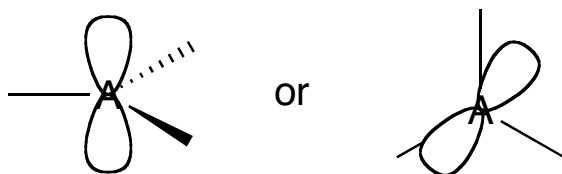
- First draw a complete Lewis Structure, and determine the hybridization, electron geometry, and molecular geometry around each central atom.
- Use hybrid orbitals to form σ bonds and to hold lone pairs of electrons.
- Use unhybridized p orbitals to form π bonds.
 - Any π bond (or delocalized π bond which exists due to resonance) must be formed by p orbitals that are aligned side-by-side and parallel to each other.
 - Familiarize yourself with the orientation of unhybridized p orbitals in relation to the hybrid orbitals for both sp and sp^2 hybridizations:

On an sp hybridized atom:



- The bold lines represent the geometry of the sp hybrid orbitals (linear, 180° apart.)
- For simplicity in this drawing, the hybrid orbitals are not shown, although their geometry and orientation are depicted.
- The 2 unhybridized p orbitals are shown. Note that the unhybridized p orbitals are at right angles to each other (as we know p orbitals are!) and at right angles to the sp hybrid orbitals.

On an sp^2 hybridized atom:



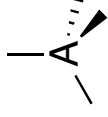
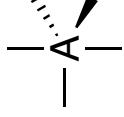
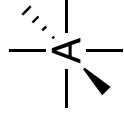
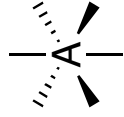


- The geometry and orientation of the sp^2 hybrid orbitals are represented with the lines, dashes, and wedges. Note that the geometry of the hybrid orbitals is trigonal planar, with 120° bond angles. For clarity, the actual hybrid orbitals are not drawn.
- The 1 unhybridized p orbital is shown. Note that the unhybridized p orbital is oriented perpendicular to the trigonal plane of the hybrid orbitals.
- Two different perspectives are shown above:
 - In the first version, we are looking “side on” at the trigonal plane of the hybrid orbitals, and the p orbital is flat in the page.
 - In the second version, we are looking straight down at the trigonal plane of the hybrid orbitals (which are thus all flat in the page) and the unhybridized p orbital is sticking out of the page and going into the page.
 - The perspective in the first drawing is typically preferred because it clearly shows the relative orientations of the hybrid orbitals and the unhybridized p orbitals. In general, you should probably only use the second perspective if absolutely necessary. (It is not incorrect, but it is not as clear as possible.)

Molecular Orbital Diagrams:

- You need to be able to draw the full molecular orbital diagram for any diatomic molecules:
 - Make sure you know the “BCN” and “OFNe” ordering of molecular orbitals.
 - Set up your molecular orbital diagram by drawing and labeling all of the molecular orbitals and the atomic orbitals from which they are derived.
 - Use a light or dotted line to connect the molecular orbitals to the atomic orbitals that they come from.
 - Bonding molecular orbitals should be lower in energy than the atomic orbitals from which they are derived.
 - Antibonding molecular orbitals should be higher in energy than the atomic orbitals from which they are derived.
 - If the diatomic molecule contains two different atoms, then the atomic orbitals on the more electronegative atom should be lower in energy than the atomic orbitals on the less electronegative atom.
 - Fill the atomic and molecular orbitals with the appropriate number of electrons.
 - If the structure contains unpaired electrons, it is paramagnetic. If all electrons are paired, it is diamagnetic.
 - Bond order = $\frac{1}{2}$ (# of bonding electrons – # of antibonding electrons)
 - The higher the bond order, the stronger and shorter the bond.
- You also need to be able to draw depictions of the molecular orbitals indicating phase, and showing how atomic orbitals can combine in phase to form bonding molecular orbitals and out of phase to form antibonding molecular orbitals.

The Complete Table of Hybridization and Geometry

Number of Electron Domains	Hybridization	Electron-Pair Geometry	Line, Dash, and Wedge Perspective Representation of the Electron-Pair Geometry	Number of Lone Pairs	Molecular Geometry	Bond Angles	Polar? (***)
2	sp (two sp hybrid orbitals)	Linear	—A—	0	Linear	180°	No
(2 unhybridized p orbitals are available for π bonding)							
3	sp ² (three sp ² hybrid orbitals)	Trigonal Planar	 or 	0	Trigonal Planar	120°	No
	(1 unhybridized p orbital is available for π bonding)			1	Bent	< 120°	Yes
4	sp ³ (four sp ³ hybrid orbitals)	Tetrahedral		0	Tetrahedral	109.5°	No
	(0 unhybridized p orbitals are leftover; no π bonding)			1	Trigonal Pyramidal	< 109.5°	Yes
				2	Bent	< 109.5°	Yes
5	sp ³ d (five sp ³ d hybrid orbitals)	Trigonal Bipyramidal		0	Trigonal Bipyramidal	90° and 120°	No
	(0 unhybridized p orbitals are leftover; no π bonding)			1	See-Saw	< 90° and < 120°	Yes
				2	T-Shaped	< 90°	Yes
Note: Contains Axial and Equatorial Positions.				3	Linear	180°	No
Lone pairs (if any) go in Equatorial Positions							
6	sp ³ d ² (six sp ³ d ² hybrid orbitals)	Octahedral	 or 	0	Octahedral	90°	No
	(0 unhybridized p orbitals are leftover; no π bonding)			1	Square Pyramidal	< 90°	Yes
				2	Square Planar	90°	No

Number of Electron Domains (or "Number of electron pairs") = (Number of Other Atoms Something is Bonded To) + (Number of Lone Pairs)

Hybrid orbitals are used to form σ bonds and to hold lone pairs of electrons.

Unhybridized p orbitals are used to form π bonds.

***Note: "**Polar?**" refers to a molecule in which all terminal atoms are the same. In general, if terminal atoms are different, the molecule will be polar.