

AP Chemistry

Chapter 9 Outline

- a) Molecular Shapes
- i) Lewis structures are two-dimensional models and only show the number and types of bonds in a molecule.
 - (1) Lewis structures do not define the three-dimensional arrangement of atoms in a molecule.
 - ii) Only a few basic shapes are observed for AB_n molecules.
- b) VSEPR Theory—simple but powerful! *Know this well.*
- i) Each lone pair, single bond, or multiple bond produces an electron domain around the central atom.
 - ii) The best arrangement of a given number of electron domains is the one that minimizes the repulsions among them.
 - iii) Molecular geometry is determined by the arrangement of only the atoms in a molecule or ion.
 - (1) Any molecule containing only two atoms is linear.
 - (2) Lone pairs exert greater repulsive forces on adjacent electron domains and so tend to compress the bond angles. (Lone pairs occupy more volume than bonds.)
 - (3) Multiple bonds typically occupy more space than single bonds.
 - iv) For larger molecules, which don't have a single central atom but instead several interior atoms, we typically discuss the geometry around the individual atoms.

# of Electron Domains	Electron-Domain Geometry	# of Bonding Domains	# of Lone Pairs	Molecular Geometry	Ideal Bond Angles
2	Linear	2	0	Linear	180°
3	Trigonal Planar	3	0	Trigonal Planar	120°
		2	1	Bent	
4	Tetrahedral	4	0	Tetrahedral	109.5°
		3	1	Trigonal Pyramidal	107°
		2	2	Bent	104.5°
5	Trigonal Bipyramidal	5	0	Trigonal Bipyramidal	90° or 120°
		4	1	See-Saw	
		3	2	T-shaped	
		2	3	Linear	
6	Octahedral	6	0	Octahedral	90°
		5	1	Square pyramidal	
		4	2	Square planar	

- c) Molecular Shape and Molecular Polarity
- i) **Bond polarity** = measure of how equally the electrons in a bond are shared between the two atoms of the bond
 - (1) Can draw as a bond vector
 - (2) Can use δ^+ and δ^- to indicate uneven charge distribution
 - ii) As the electronegativity difference between the atoms in the bond increases, the dipole moment of the bond increases.
 - iii) For a molecule containing more than two atoms, the dipole moment of the molecule depends on the polarity of the bonds and their geometric arrangement.
 - (1) The overall dipole moment of a molecule is the **vector sum of its bond dipoles**.
 - (2) CO_2 is nonpolar, even though the C-O bond is polar.
 - (a) The two dipoles “cancel” each other out.
 - (b) The sum of the vectors is zero
 - (3) H_2O is polar, because the bond dipoles reinforce each other.
 - (4) For very symmetrical shapes, if all the atoms bonded to the central atom are identical, there will be no dipole moment for the molecule: linear, trigonal planar, tetrahedral, square planar, trigonal bipyramidal, and octahedral.

a) Covalent Bonding and Orbital Overlap

- i) Valence Bond Theory
 - (1) Bonds form when a valence atomic orbital of one atom overlaps with that of another atom.
 - (a) The orbitals share a region of space.
 - (b) Now electrons of opposite spin can share the common space between the nuclei. Electrons are simultaneously attracted to both nuclei.
- ii) The optimum distance between the nuclei puts the system at a potential energy minimum.
 - (1) Too far apart: no attractive forces
 - (2) Too close together: lots of repulsion between the nuclei

a) Hybrid Orbitals

- i) Basically, hybrid orbitals are a way to reconcile valence bond theory and observed molecular geometries.
- ii) Assume that atomic orbitals on an atom mix to form new “**hybrid orbitals**” in a process called “hybridization.” *Know this table!*

Hybridization	Electron Domain Geometry	Bond angles
sp	Linear	180°
sp^2	Trigonal planar	120°
sp^3	Tetrahedral	109.5°
sp^3d	Trigonal bipyramidal*	
sp^3d^2	Octahedral*	

*For the redesigned AP Chemistry course, students are no longer required to know hybridizations involving d orbitals, as the evidence to support this has been questioned.

a) Multiple Bonds

i) [Sigma \(\$\sigma\$ \) bonds](#)

- (1) Electron density is along the bond axis
- (2) End-to-end overlap of orbitals

ii) [Pi \(\$\pi\$ \) bonds](#)

- (1) Electron density is above and below the bond axis
 - (a) Side to side overlap of unhybridized p orbitals
 - (b) Atoms must have sp or sp² hybridization to make π bonds
 - (c) Pi bonds tend to be weaker than sigma bonds, because the total orbital overlap is less.

iii) In general, single bonds are σ bonds. *Know this!*

- (1) [A double bond](#) = 1 σ bond & 1 π bond *Know this!*
- (2) A triple bond = 1 σ bond & 2 π bonds *Know this!*
- (3) Double and triple bonds are more common in small atoms (especially C, N, and O). Larger atoms (S, P, Si) do not form π bonds readily.

iv) Delocalized π

<http://wps.prenhall.com/wps/media/objects/3081/3155729/blb0906/bl09fg30.jpg>

- (1) Resonance structures involving π bonds gives a special stability.
 - (a) The π electrons are in π bonds that extend over more than two bonded atoms.
 - (b) The electrons are “delocalized.”