

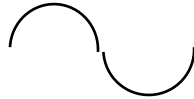
Chap 2

1. Bonding.

a) ψ described as waves, a wave function.

Linear combination of atomic orbitals- LCAO

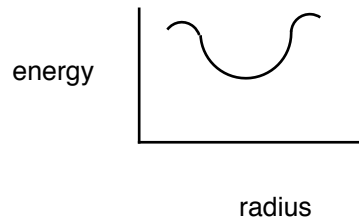
Atomic orbitals combine to and overlap to give standing waves.



Covalent bonding (LCAO)- two non-metals.

Ionic bonding- metal, non-metal

Bond length- stability of covalent bonds dependent on max overlap and least amount of electrostatic repulsion.

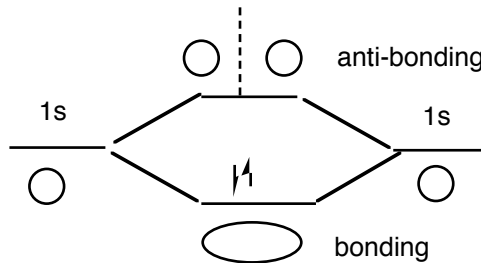


b) Two types of bonds, σ and π .

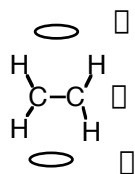
1) sigma bond (single bond)- can add destructively or constructively.

Constructively- bonding molecular orbital (bonding mo) σ

Destructively- anti-bonding molecular orbital. (Anti-bonding mo) σ^*

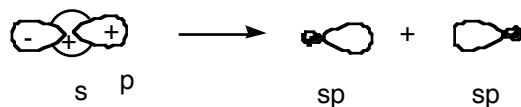


2) \square bond, double and triple bonds. (show triple bonds.)



2. Orbital Theory- explains VSEPR- **Review orbital shapes**

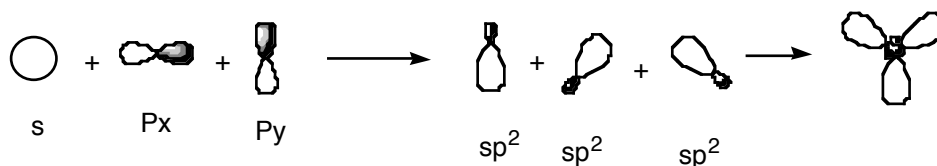
a) sp hybrid- 2 hybrid orbitals form a combined orbital on the same atom.



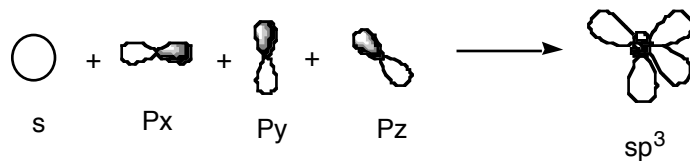
Show overhead.

Start with sp so get two sp hybridized orbits.

b) sp^2 - hybrid- get 3 hybrid orbitals

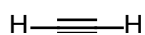


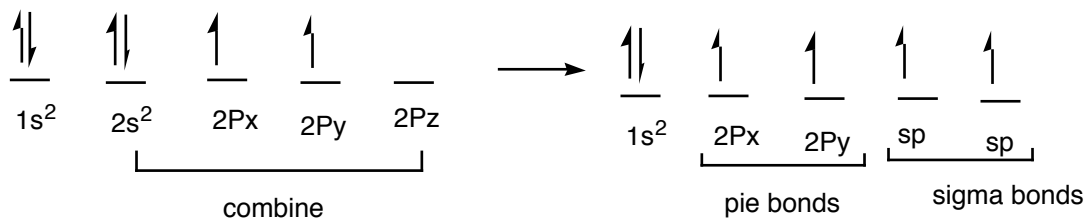
c) sp^3 - get 4 hybrid orbitals.



Relate to VSEPR theory to determine molecular geometry. Sigma bonds determine geometry.

a) sp carbon $1s^2 2s^2 2p^2$

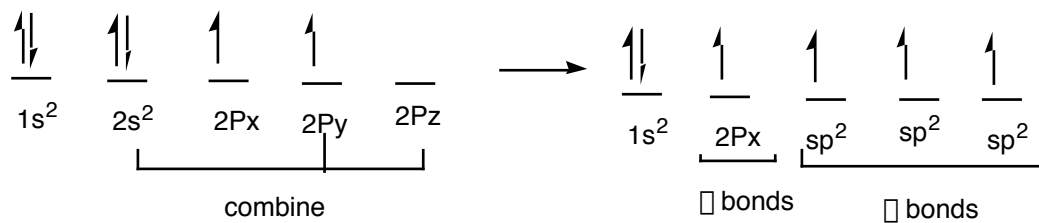
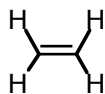




Triple bond has two pie bonds and two sigma bonds that determine geometry.

180 degrees

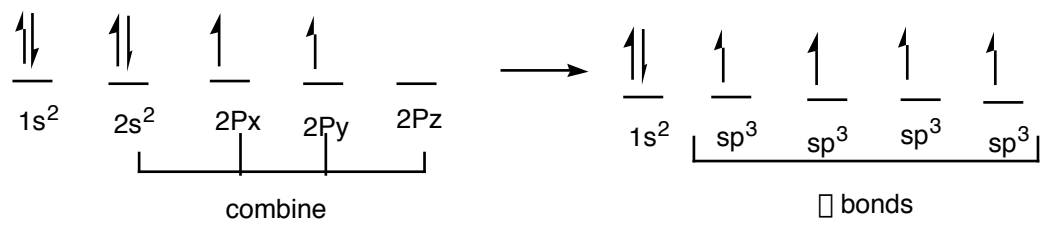
b) sp^2 - trigonal planar geometry- 120 degrees



c) sp^3 - 109 degrees tetrahedral geometry

draw wedges and hash correctly.

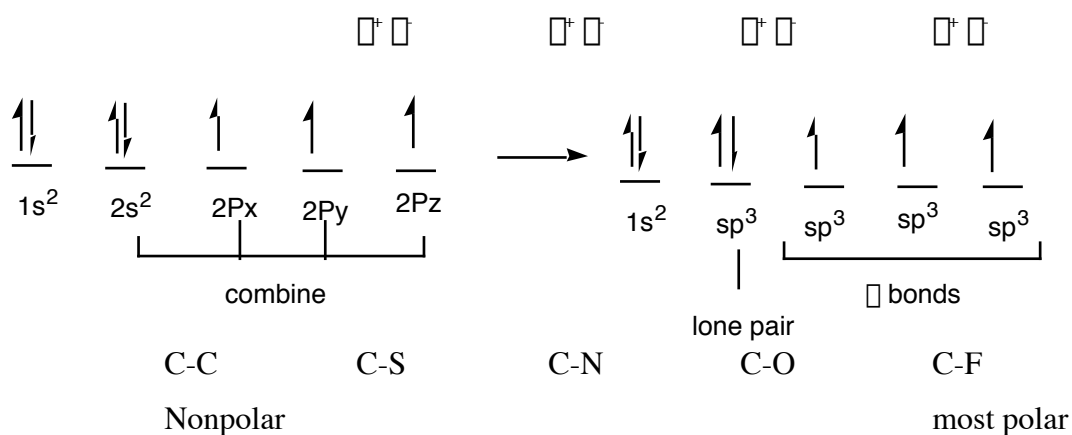




Lone pairs are in hybridized orbitals Ex NH_3 nitrogen $1s^2 2s^2 2p^3$

2) Polarity- bond and molecular

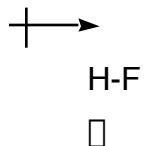
a) bond polarity- dependent on a difference in electronegativity. F is the highest. Unequal sharing of e^- .



Greater difference become ionic compounds C-Li

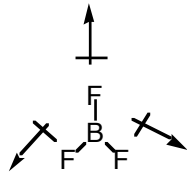
Can measure net dipole moment, μ

$$\mu = q \times d \quad \text{charge} \times \text{distance shown by a vector}$$



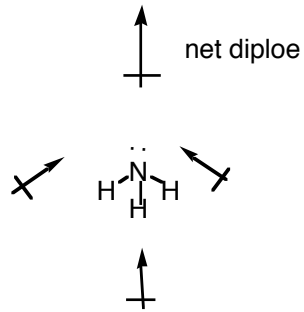
b) Molecular polarity- dependent on molecular geometry





no net dipole moment = 0 symmetrical

NH₃

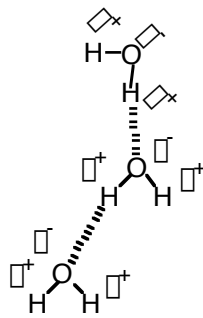


Polar or not polar H₂S and HCN

Intermolecular forces- between two molecules, gives rise to physical properties

3 major forces

a) H-bond, H directly attached to O,N,F HF only one not organic, HCl strong acid.



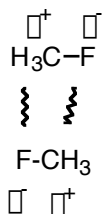
CH₃-CH₂-OH

CH₃-O-CH₃

bp 78 C

-25 C

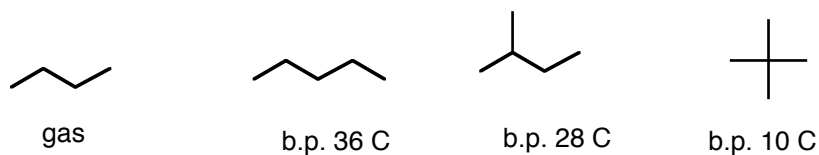
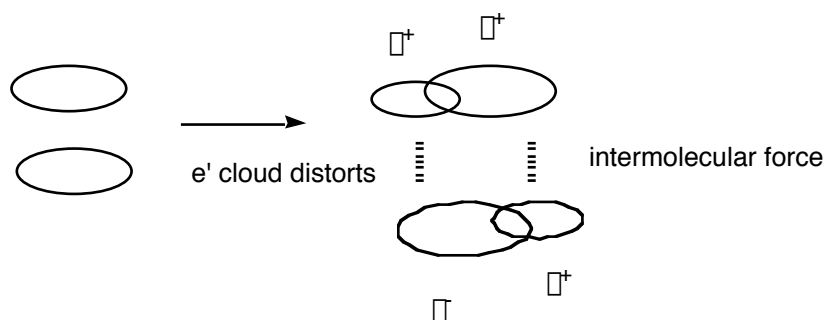
b) dipole-dipole, polar-polar



c) Induced dipole (London forces, dispersion forces, Van der Waals)-
non-polar non-polar interaction.

Directly linked to:

- 1) surface area contact between two molecules. (branching)
- 2) Size of electron cloud , size of molecule or atom.



b.p. $\text{H}_2\text{S} > \text{PH}_3 > \text{SiH}_4$ due to polarity

b.p. $\text{H}_2\text{Te} > \text{H}_2\text{Se} > \text{H}_2\text{S}$ due to size

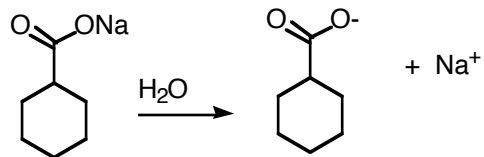
Solubility- depends on intermolecular forces and polarity

	Miscible	Immiscible
$\text{A}-\text{A} + \text{B}-\text{B} \quad \square$	$\text{A}-\text{A}$	$\text{A}-\text{B}$
	$\text{B}-\text{B}$	$\text{B}-\text{A}$

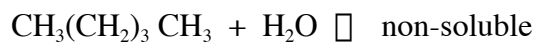
1) Ionic- depends on solubility tables (in water)



Assume ionic bounds with complex organic molecules will dissociate to give ions.



2) Covalent, depends on the number of polar functional groups and molecular geometry, (dipole moment)



Charges make compound more water soluble.