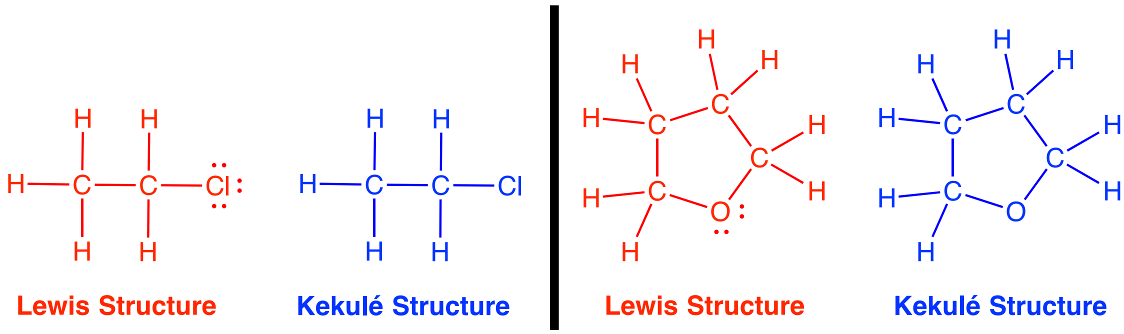
Organic Chemistry Prep Workshop – Day 2

Drawing Organic Structures

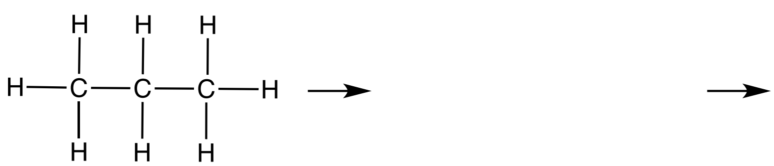
- **Lewis Structures** – Includes all atoms, all bonds, and all lone pair electrons.
- **Kekulé Structures** – Lewis structures with the lone pairs omitted.



Lewis Structure Kekulé Structure | Lewis Structure Kekulé Structure

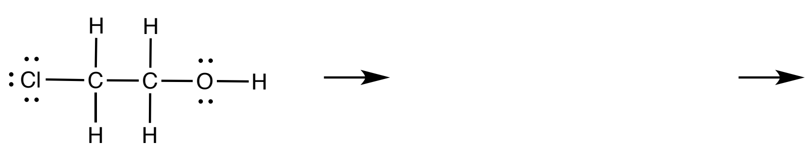
Skeletal Structures

- Also called bond-line structures
- Drawn in a zig-zag fashion
- Each end represents a carbon
- Each vertex represents a carbon
- The hydrogen on carbons are implied

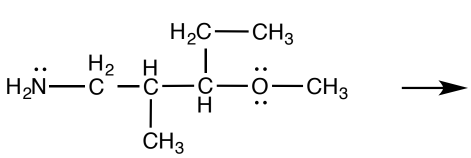


Partially Condensed Skeletal Structure

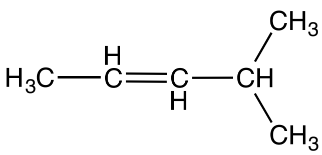
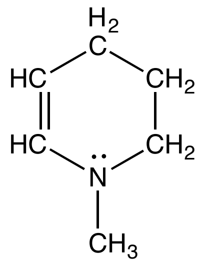
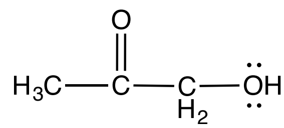
- Heteroatoms (O, N, S, Cl, etc.) must be drawn in.
- Hydrogen bonded to heteroatoms must be drawn in.
- Lone pairs may be drawn in but are often omitted.

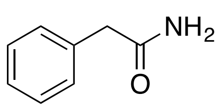
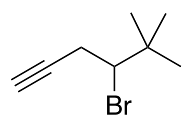
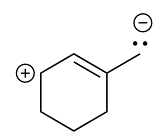


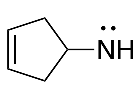
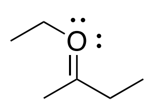
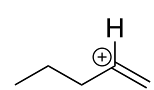
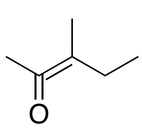
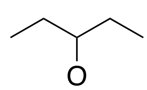
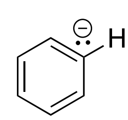
Partially Condensed Skeletal Structure

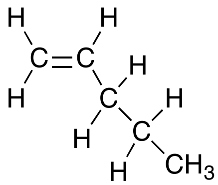
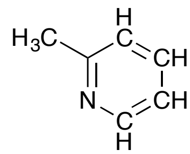
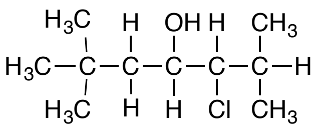
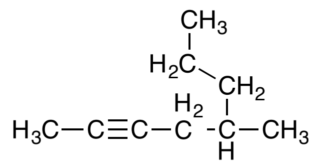


Skeletal Structure

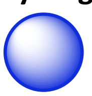
You Try 2-1		
For each structure below, determine the number of hydrogen that are present on every atom.		
		

You Try 2-2		
Every structure below has a problem. Determine what is wrong with each.		
		
		

You Try 2-3	
Draw a valid skeletal structure for each of the following.	
	
	

Atomic Orbitals

Hydrogen: 1s[1]



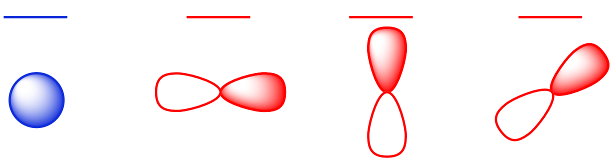
s-orbital

Sigma (σ) bonds are formed by the head-on overlap of orbitals.

Bonding = orbital overlap

Bonding in H_2:

Carbon: $2s^22p^2$

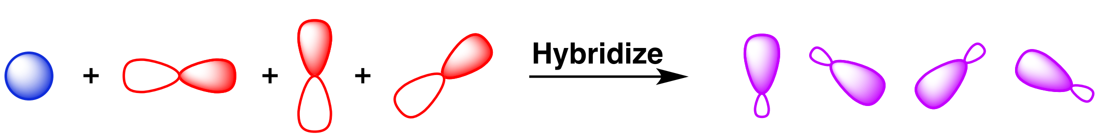


2s-orbital p$_x$-orbital p$_y$-orbital p$_z$-orbital

Why is CH_4 tetrahedral with four equivalent bonds?

Hybridization

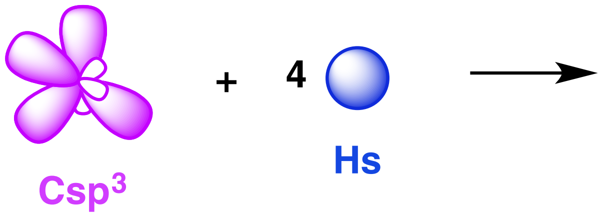
Orbitals can hybridize to become more stable.



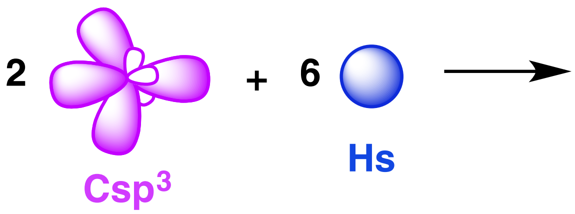
s p$_x$ p$_y$ p$_z$

4 x sp^3 hybrid orbitals

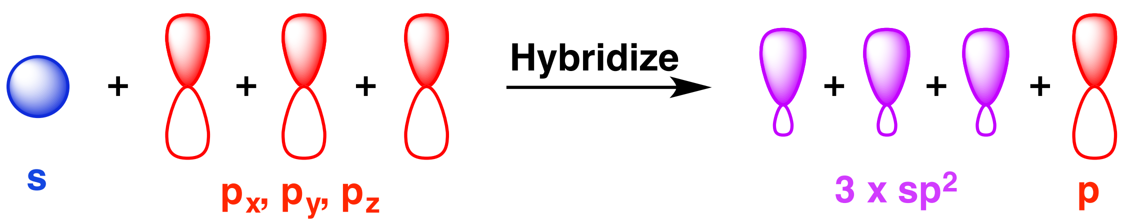
Bonding in Methane (CH_4)

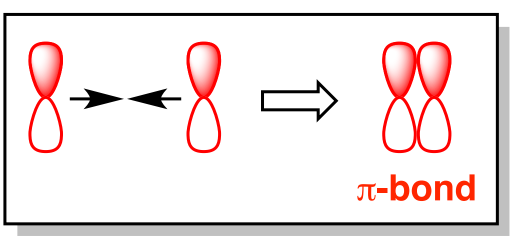


Bonding in Ethane ($H_3C\text{-}CH_3$)

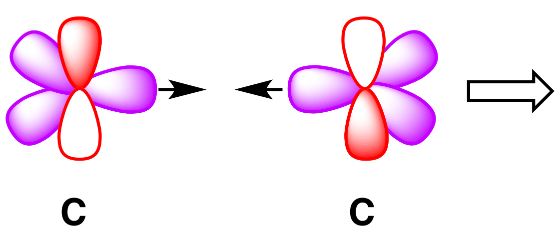


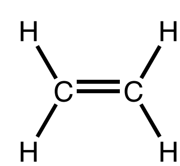
Bonding in Ethene ($H_2C{=}CH_2$)



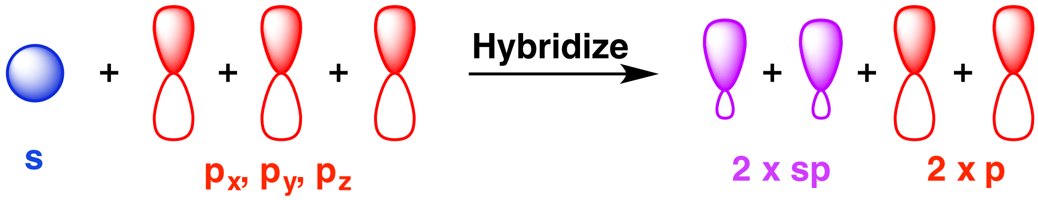


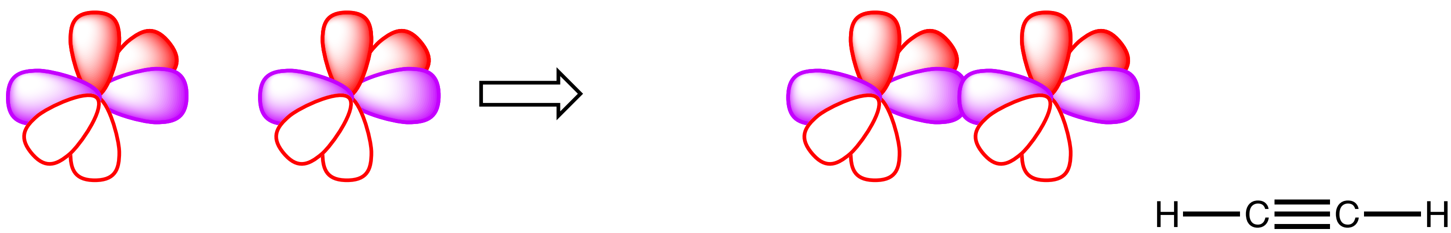
Sigma (σ) bonds are formed by the head-on overlap of orbitals.
Pi (π) bonds are formed by the side-on overlap of p-orbitals.

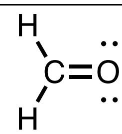




Bonding in Ethyne (HC≡CH)





You Try 2-4	
Construct an orbital picture for each of the following:	
$:NH_3$	

Hybridization Summary

An atom with 4 groups is sp^3 hybridized.

- **Tetrahedral Geometry**
- **109.5° bond angles**

What counts as a group?

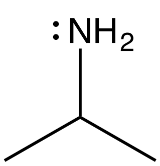
A bond: — = ≡

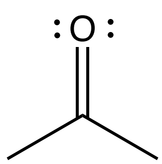
A lone pair of electrons: • •

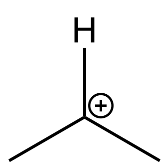
An atom with 3 groups is sp^2 hybridized.
- **Trigonal Planar Geometry**
- **120° bond angles**

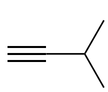
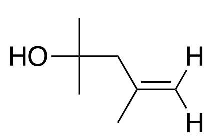
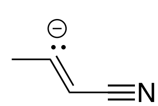
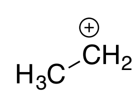
An atom with 2 groups is sp hybridized.
- **Linear Geometry**
- **180° bond angles**

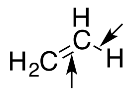
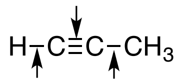
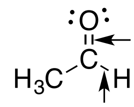
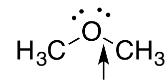
Analyzing an Atom's Orbitals







You Try 2-5			
Determine the hybridization for every non-hydrogen atom in each molecule below.			
			

You Try 2-6			
For each structure, determine the orbitals that make up the indicated bonds.			
			

Bond Strength and Length

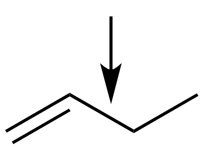
Compare single, double, and triple bonds.

$H_3C—CH_3$ $H_2C=CH_2$ $HC≡CH$

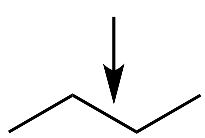
s-Orbitals lie closer to the positively charged nucleus than p-orbitals. Therefore, electrons (-) in s-orbitals are held more tightly.

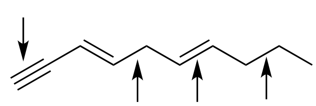
Higher % s-character = more tightly held e- = shorter/stronger bond.

sp^3 = 25% s, 75% p ----- **sp^2** = 33% s, 66% p ----- **sp** = 50% s, 50% p



vs



You Try 2-7
Rank the indicated bonds in the molecule below from strongest (1) to weakest (4).


VSEPR

Valence Shell Electron Pair Repulsion Theory

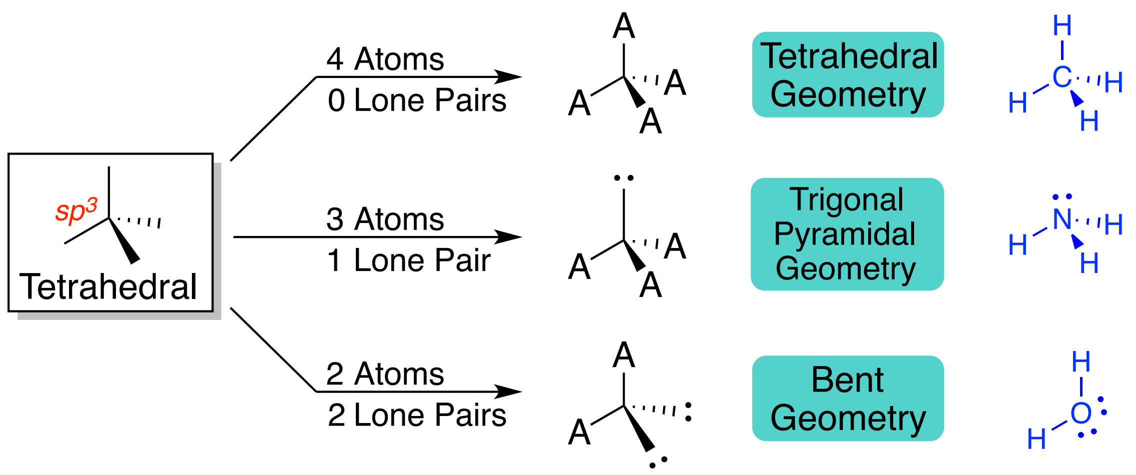
Electron Pair Arrangement = how the groups (bonds & lone pairs) are arranged around the central atom.

Geometry = how the atoms are arranged.

sp^3 Geometries

Electron Pair Arrangement = tetrahedral

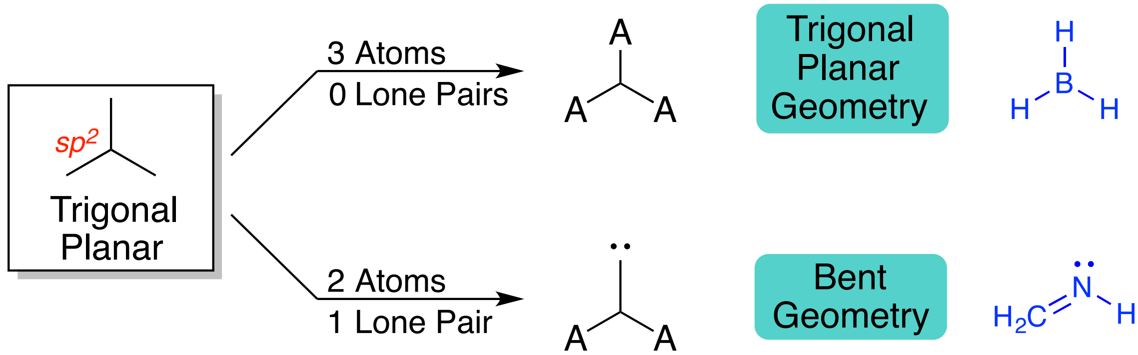
Example



sp^2 Geometries

Electron Pair Arrangement = trigonal planar

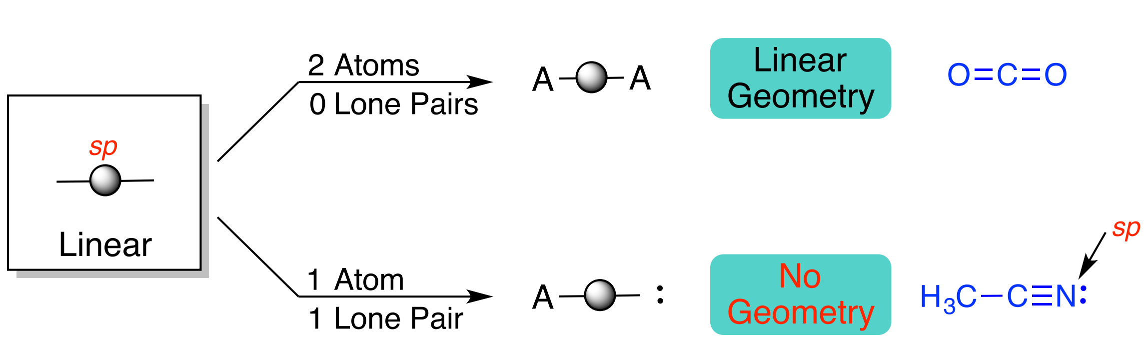
Example

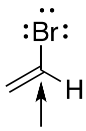
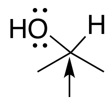
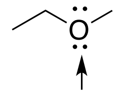
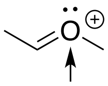
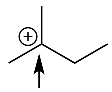
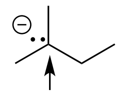


sp Geometry

Electron Pair Arrangement = Linear

Example



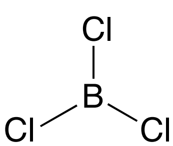
You Try 2-8		
For each molecule, determine the electron pair arrangement and geometry around the indicated atom.		
		
		

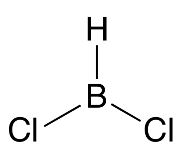
Molecular Polarity

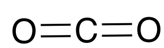
Bond Dipole = The dipole for an individual bond.
Molecular Dipole = Vector sum of the individual bond dipoles in a molecule.
If a molecule has a net molecular dipole moment, the molecule is said to be polar.
In a molecule, there are individual dipoles associated with:

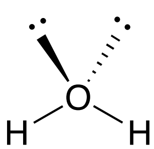
- A–B bond (A → B) where B is more electronegative than A
- A–: where the dipole moment points to the lone pair A→:

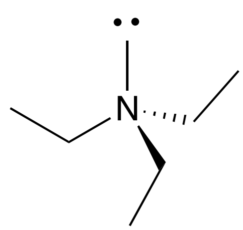
You must first determine the molecular geometry in order to determine whether or not a molecule is polar.

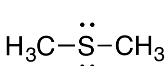
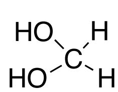
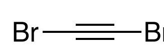




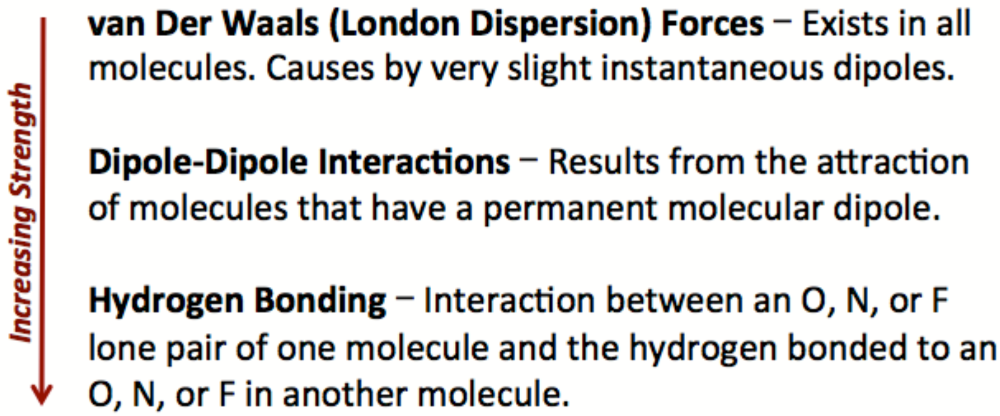






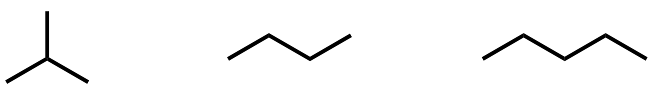
You Try 2-9		
Determine whether or not each molecule is polar. If it is polar, draw the net molecular dipole.		
		

Intermolecular Forces



van Der Waals (London Dispersion) Forces – Exists in all molecules. Causes by very slight instantaneous dipoles. The strength of van Der Waals forces increase with increasing surface area.

Compare:



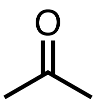
Dipole-Dipole Interactions – Results from the attraction of molecules that have a permanent molecular dipole.

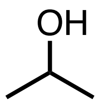
Hydrogen Bonding – Interaction between an O, N, or F lone pair of one molecule and the hydrogen bonded to an O, N, or F in another molecule.

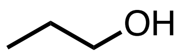
Boiling Point

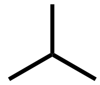
The boiling point for a compound increases as the strength of the intermolecular forces increases.

Rank:

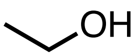




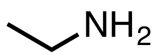




Compare:

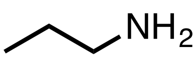


bp 78 °C

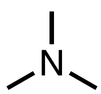


bp 17 °C

Compare:



bp 49 °C



bp 4 °C

You Try 2-10
For each set of compounds, identify all intermolecular forces present and then rank in order of increasing boiling point.