

**Hybrid Orbitals (Sections 2.4, 2.6);  $\pi$  bonding (Section 2.3)**

$1s + 3p \rightarrow 4 sp^3$ hybrids	109°
$1s + 2p$ (+ 1 unhybridized p) $\rightarrow 3 sp^2$ hybrids (+ 1 unhybridized p)	120°
$1s + 1p$ (+ 2 unhybridized p's) $\rightarrow 2 sp$ hybrids (+ 2 unhybridized p's)	180°

Why does hybridization occur?

- Hybrid orbitals are big and point in one direction. Their **directionality** leads to **better overlap** which leads to **strong bonds**.
- Hybrid orbitals leads to nice VSEPR angles

If hybridization is so great, why aren't pure monatomic atoms hybridized?

- For an isolated atom, having 1 s and 3 p atomic orbitals is better than 4  $sp^3$  hybrid orbitals
- However, when covalent bonds can result, the small price of hybridizing is paid off a thousandfold by the payoff of making strong, good VSEPR bonds

If hybridization is so great, why aren't all carbons  $sp^3$  hybridized? Why do some stay  $sp^2$  or  $sp$ , and withhold some p orbitals from hybridization?

- p orbitals are withheld from hybridization for the sole purpose of using them to make  $\pi$  bonds.
- Only when double bonds or triple bonds are involved is the hybridization less than the full  $sp^3$
- Each  $\pi$  bond requires the attached atoms to use p orbitals

2 Kinds of Covalent Bonds

- sigma ( $\sigma$ ) bonds: electron density is along the axis between the nuclei
  - $\sigma$  bonds always involve the overlap of s or s-containing hybrids (s, sp,  $sp^2$ ,  $sp^3$ )
- pi ( $\pi$ ) bonds: electron density is either above/below or before/behind, but not along the internuclear axis
  - $\pi$  bonds involve the overlap of parallel p orbitals

The first bond in any bond (whether single, double, or triple), is a  $\sigma$  bond. The "extra" bonds in a double or triple bond are  $\pi$  bonds.

**Bond     $\sigma$      $\pi$** 

Single	1	0
Double	1	1
Triple	1	2

$\pi$  bonds are weaker and more reactive than  $\sigma$  bonds. Most organic reactions involve  $\pi$  bonds