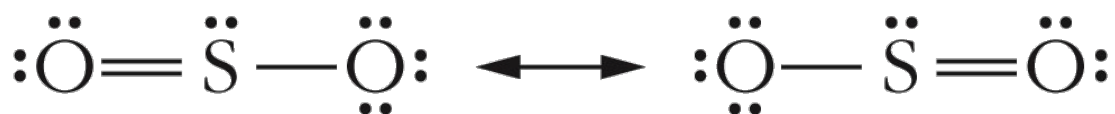


# Drawing Lewis Structures

Some examples

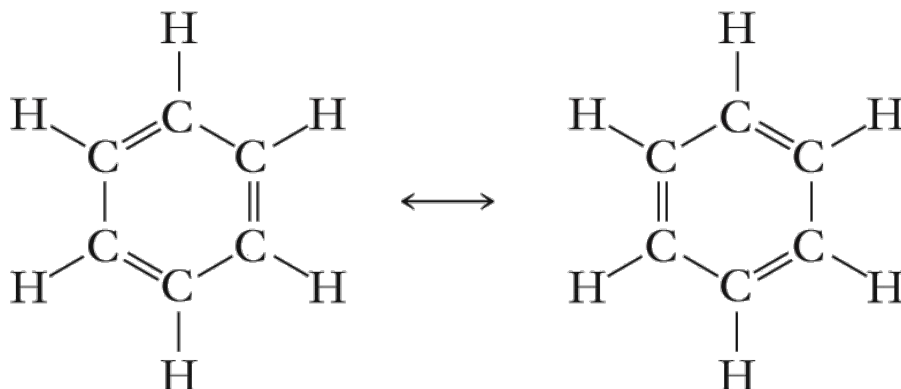


## Resonance



- The two resonance structures for  $\text{SO}_2$ .
- The “real” structure is a hybrid of the two structures.
- The **double arrow** indicates the structures are resonance structures.
- How do we know this is right? **Experimentally both S–O bonds are equal in length.**

## Resonance



- The two resonance structures for benzene.
- Indicates the two ways to distribute 3 double bonds among 6 carbon atoms.
- The true structure is an average of the two resonance structures.

## Formal Charge

If two nonequivalent Lewis Structures are possible, how do we decide?

- Calculate formal charges on each atom
  1. Assign each lone pair electron solely to its atom
  2. Assign half of the bonding electrons to the atom

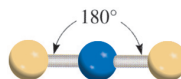
$$\text{Formal Charge} = (\# \text{ of valence } e^-) - (\# \text{ of assigned } e^-)$$

- Choose structure with lowest formal charge
- Examples:  $\text{CO}_2$ ,  $\text{CH}_2\text{O}$

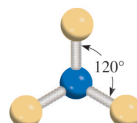
# Bonding orbitals in polyatomics

In polyatomics, the bonding geometry is often incompatible with our  $s$  and  $p$  orbitals on atoms

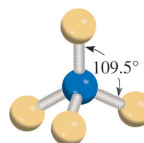
- $\text{BeCl}_2$  is linear      $\text{Cl}-\text{Be}-\text{Cl}$



- $\text{BF}_3$  is triangular planar

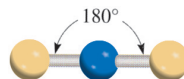


- $\text{CH}_4$  is tetrahedral

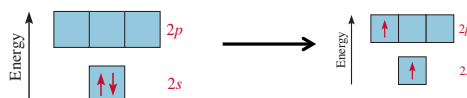


# Bonding orbitals in polyatomics

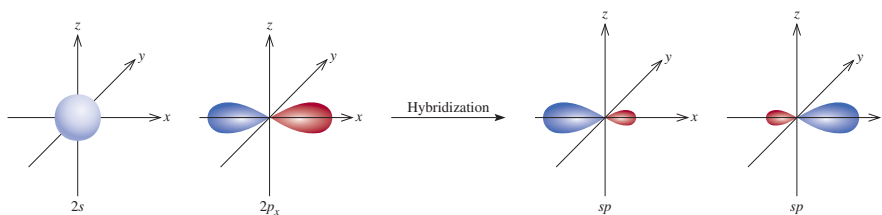
Look first at  $\text{BeCl}_2$



Be does not have two orbitals oriented at  $180^\circ$  that can bond to each of the F atoms. How can we form such orbitals - We construct **hybrid orbitals**



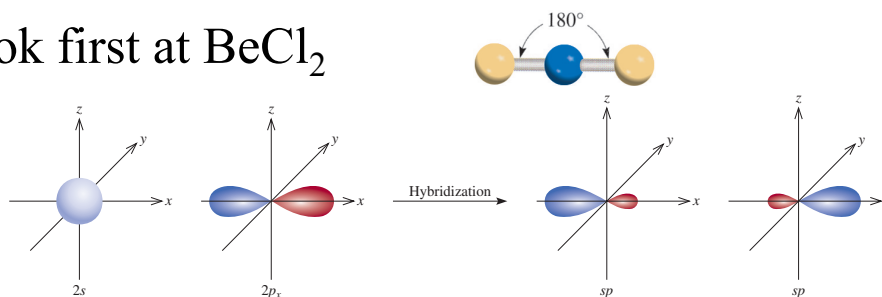
The energy gained by bonding can overcome the  $2s/2p$  energy difference and these 2 orbitals can *hybridize* to form 2  $sp$  hybrid orbitals



See  
Section 7.7

# Bonding using $sp$ hybrid orbitals

Look first at  $\text{BeCl}_2$



$$\text{First } sp \text{ orbital} = \frac{1}{\sqrt{2}}[\phi_{2s} + \phi_{2p_x}]$$

$$\text{Second } sp \text{ orbital} = \frac{1}{\sqrt{2}}[\phi_{2s} - \phi_{2p_x}]$$

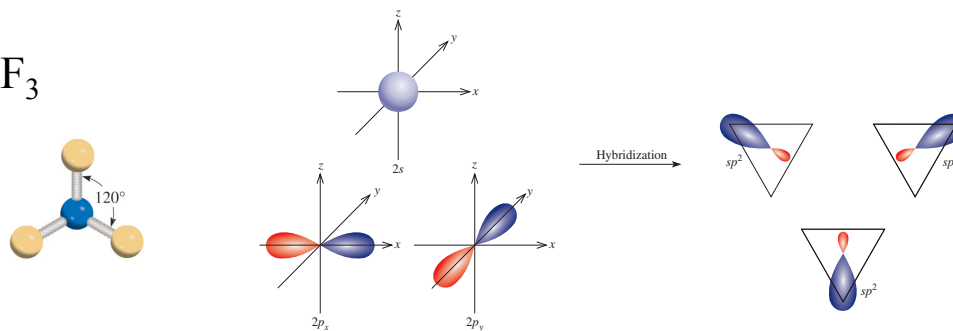
Bonding in  $\text{BeCl}_2$  then comes from the overlap of the H atom  $1s$  orbitals with the  $sp$  hybrid orbitals on Be



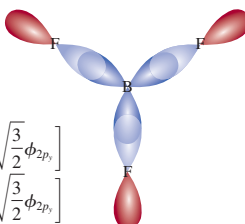
## Other types of hybrid orbitals

- $sp^2$  hybrid orbitals (triangular planar geometry)

$\text{BF}_3$



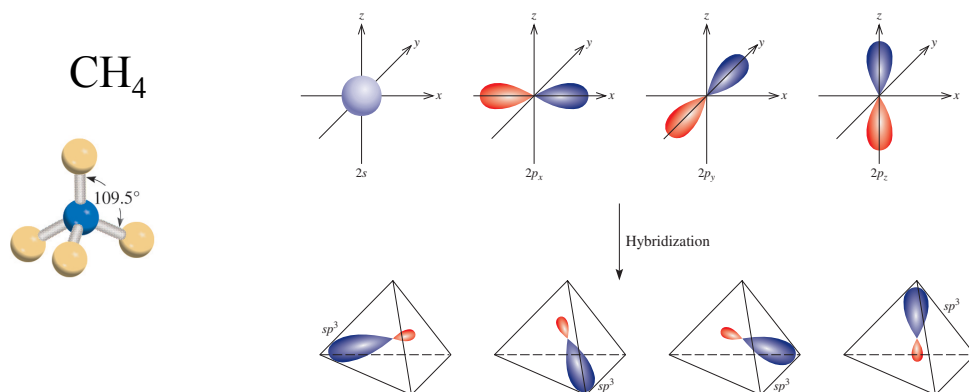
$$\begin{aligned} \text{First } sp^2 \text{ orbital} &= \frac{1}{\sqrt{3}}[\phi_{2s} + \sqrt{2}\phi_{2p_x}] \\ \text{Second } sp^2 \text{ orbital} &= \frac{1}{\sqrt{3}}\left[\phi_{2s} - \frac{1}{\sqrt{2}}\phi_{2p_x} + \sqrt{\frac{3}{2}}\phi_{2p_y}\right] \\ \text{Third } sp^2 \text{ orbital} &= \frac{1}{\sqrt{3}}\left[\phi_{2s} - \frac{1}{\sqrt{2}}\phi_{2p_x} - \sqrt{\frac{3}{2}}\phi_{2p_y}\right] \end{aligned}$$



Here the  $sp^2$  orbitals on B overlap with the  $2p_z$  orbitals on F to form the bonds in  $\text{BF}_3$

# Other types of hybrid orbitals

- $sp^3$  hybrid orbitals (tetrahedral geometry)



$$\text{First } sp^3 \text{ orbital} = \frac{1}{2}[\phi_{2s} + \phi_{2p_x} + \phi_{2p_y} + \phi_{2p_z}]$$

$$\text{Second } sp^3 \text{ orbital} = \frac{1}{2}[\phi_{2s} + \phi_{2p_x} - \phi_{2p_y} - \phi_{2p_z}]$$

$$\text{Third } sp^3 \text{ orbital} = \frac{1}{2}[\phi_{2s} - \phi_{2p_x} + \phi_{2p_y} - \phi_{2p_z}]$$

$$\text{Fourth } sp^3 \text{ orbital} = \frac{1}{2}[\phi_{2s} - \phi_{2p_x} - \phi_{2p_y} + \phi_{2p_z}]$$

Here the  $sp^3$  orbitals on C overlap with the  $1s$  orbitals on H to form the bonds in CH<sub>4</sub> (methane)

