9.1 - Hybridization & the Localized Electron Model

Hybridization Overview

So far you have learned how to draw Lewis structures from chemical formulas, and determine shape and polarity of molecules using the VSEPR model.

In this unit, we delve further into the shapes of the electron clouds which form in the valence regions around molecules (Localized Electron (LE) model).

Earlier you learned that different numbers of pairs of electrons in orbitals contributed to the hybridization of that atom - focusing on only the central atom.

Here, we will look at all the atoms in a molecule to get a clearer image as to where the electrons actually are.

Revisiting Sublevel Shape

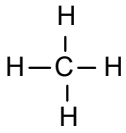
Recall the shapes of the sublevels involved in bonds: S = spherical, P = dumbbell, and D = double dumbbell.

The concept of producing hybridized orbitals relies on merging those different shapes into symmetric shapes which equally fill the space around the atom in question.

Later on, we will look at molecules that rely on hybridization, as well as non-hybridization to display electron bonding locations as we get into double and triple bonds.

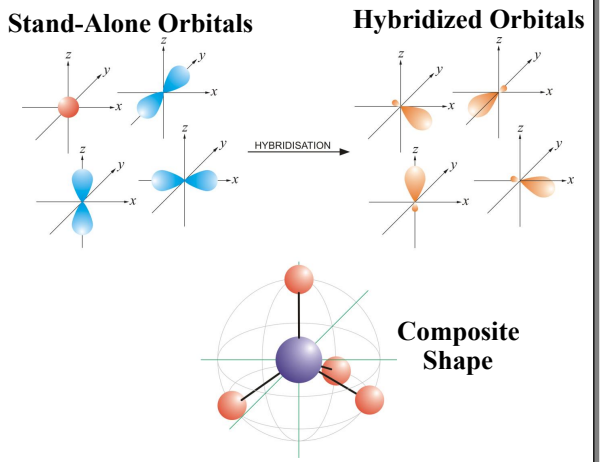
sp^3 Hybridization: Methane Example

We will start with methane, CH_4, as an example. Here, each of carbon's valence electrons is joined with hydrogen, and the Lewis structure is here:



To determine how hybridization merges the single orbital of the 2s shell with the three orbitals of the 2p shell, consider the following geometric example. This shows the progression from stand-alone orbitals to hybridized ones, resulting in the composite tetrahedral shape of the methane molecule.

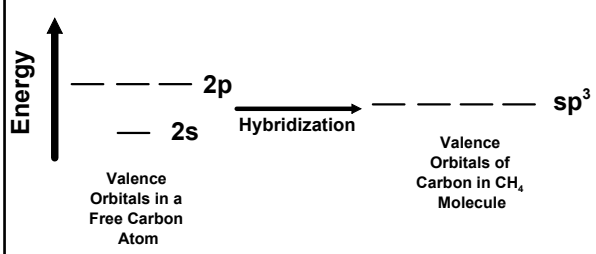
sp^3 Hybridization: Methane Example



Energy Level Approach

It will be increasingly useful to look at hybridization from an energetic perspective as we get deeper into this chapter.

In methane (specifically the carbon atom), the energy level diagram is shown:



1. Ammonia Example

Describe the bonding in the ammonia (NH_3) molecule using the localized electron model.

First, make a Lewis structure:

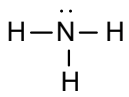


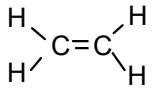
Figure out the hybridization of nitrogen: sp^3 is produced when the 2s and 2p orbitals combine.

How are electrons distributed around nitrogen? Three bonding pairs fill three of the sp^3 orbitals, and the lone pair occupies the fourth sp^3 orbital.

sp^2 Hybridization

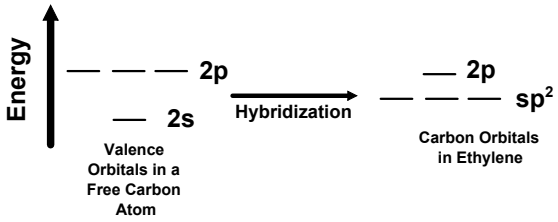
An atom with sp^2 hybridization has a geometry that is trigonal planar, with 120° bond angles.

Consider the ethylene molecule: C_2H_4, with its associated Lewis Structure:



Here, each carbon has sp^2 hybridization due to having only three attached atoms, and no lone pairs.

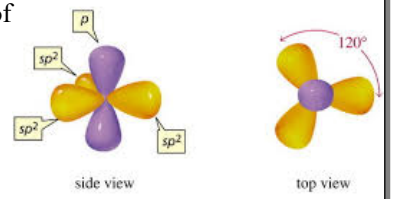
Energetically, it looks like this:



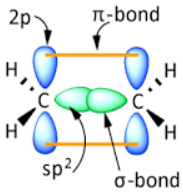
sp^2 Hybridization

The upshot is that the unhybridized 2p sublevel allows for the presence of the double bond.

The combination of sp^2 and 2p orbitals results in the following for each carbon atom of the molecule:



When the two carbon atoms join with BOTH single and double bonds, the sp^2 and 2p orbitals are involved. Note in the picture that only one bond is labeled, although there are two lines show possible bond locations.



Terminology

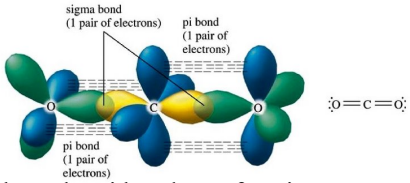
So far we've dealt with bonds by calling them single, double, or triple.

The single bond is a pair of electrons centered on a line between the atoms: these are **sigma (σ) bonds.**

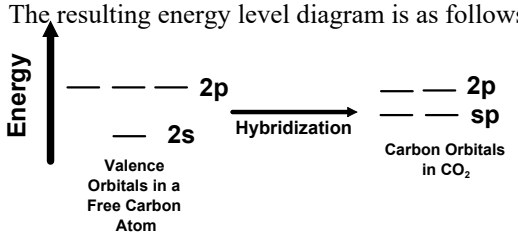
When parallel p orbitals join, as you saw in ethylene, and form double or triple bonds, these multiple bonds are called **pi (π) bonds**.

sp Hybridization

Another hybridization occurs in CO_2. With two atoms connected to it, carbon undergoes sp hybridization, while oxygen undergoes sp^2:

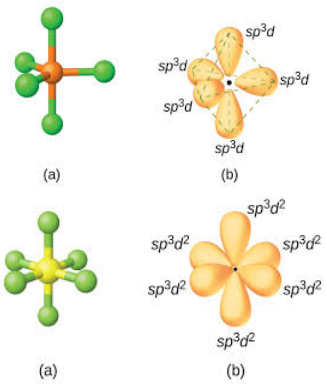


Note where the pi bonds are forming.
The resulting energy level diagram is as follows.



More Hybrids

We will study more possible hybridization configurations, but here are two examples of common ones: sp^3d and sp^3d^2:

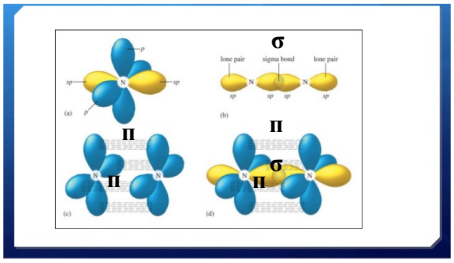


2. Nitrogen Example

Describe the bonding in the N_2 molecule.

Lewis structure: :N≡N:

To account for all three bonds, per each nitrogen atom, two pairs of electrons will be sp hybridized (along a line drawn between the nitrogen centers) - one bonding and one lone pair. An additional two pairs of bonding electrons are located between non-hybridized 2p orbitals.



Homework

Preview 9.2 - 9.3

9.1 - Problems in your Booklet
Due: Next Class