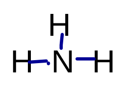
Chapter 4 - Hybridization Problems #1

Click on the pink circles to listen to the written portions and get helpful hints! Use the Drawing and Text box tools to complete the questions!

One model to explain bonding in molecules is called the **valence bond theory**. It is based on the concept of **hybridized molecular orbitals**. Atomic orbitals (*s,p,d*) are mathematically mixed to form molecular orbitals with defined shapes and energy levels. In organic chemistry, we need only look at the hybridization of the one 2*s* and three 2*p* orbitals containing the valence electrons of C,O and N.

When **four** equivalent molecular orbitals are needed, all **four** atomic orbitals are mixed to give ***sp*^{3} hybridization**. When **three** equivalent molecular orbitals are needed, **three** of the atomic orbitals are mixed to give ***sp*^{2} hybridization**. In this case, one unhybridized ***p*** atomic orbital remains, with its shape and energy level unchanged. When **two** equivalent molecular orbitals are needed, only **two** of the four atomic orbitals are mixed, giving ***sp* hybridization**. Two unhybridized atomic *p* orbitals remain.

. Lewis structures help to determine how many equivalent molecular orbitals are present. Complete the Lewis structure of ammonia, NH_3.



a. How many atoms are bonded to N — 3

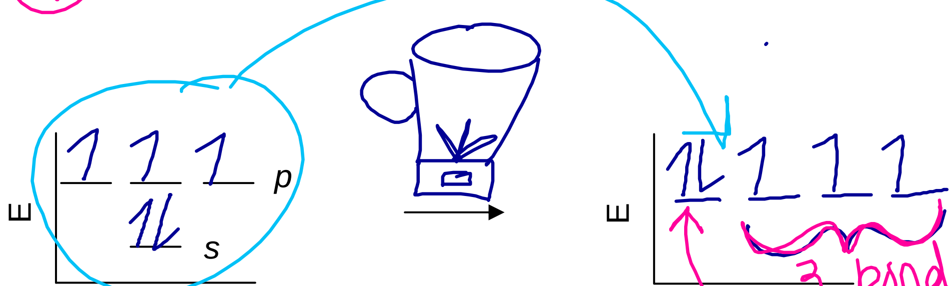
b. How many lone pairs of e^- are on N — + 1

c. How many equivalent molecular orbitals — 4 ← Steric

d. Hybridization — sp^3 — SPPP 1234

e. Fill in the e^- for atomic N ← 5 valence e^-
f. Draw the hybrid orbitals for NH_3
g. Fill in the e^- for N in NH_3
h. Compare the Lewis structure with the molecular orbital diagram

4 total



atomic orbitals — molecular orbitals

3 bonds — Lone pair — sp^3 orbs

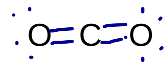
N is in group 15 = 5 val.

$4 : sp^3 = sppp$

$3 = sp^2 = spp$

$2 = sp = sp$

2. Complete the Lewis structure of carbon dioxide, CO_2.



a. How many atoms are bonded to C ___2___

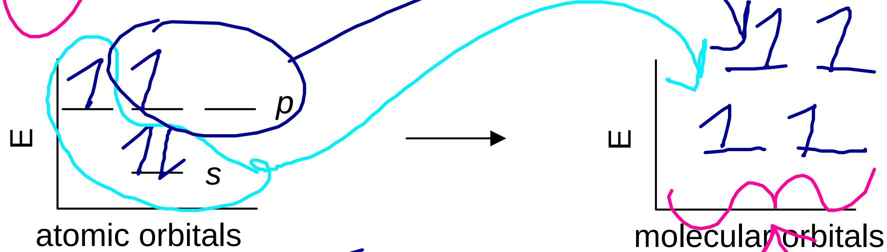
b. How many lone pairs of e^- are on C ___+ 0___

c. How many equivalent molecular orbitals ___2___

d. Hybridization ___sp___

e. Fill in the e^- for atomic C
f. Draw the hybrid orbitals for C in CO_2
g. Fill in the e^- for C in CO_2
h. Compare the Lewis structure with the molecular orbital diagram

π bonds to O



σ bond to O

$C = 4$ val e^-

3. Repeat this for the oxygen in CO_2.

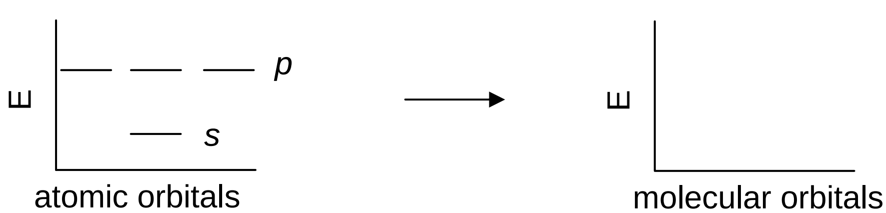
a. How many atoms are bonded to O ______

b. How many lone pairs of e^- are on O ______

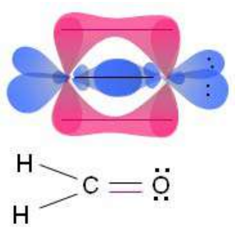
c. How many equivalent molecular orbitals ______

d. Hybridization ______

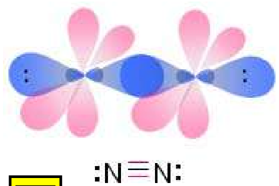
e. Fill in the e^- for atomic O
f. Draw the hybrid orbitals for O in CO_2
g. Fill in the e^- for O in CO_2
h. Compare the Lewis structure with the molecular orbital diagram



Bonds made with hybridized orbitals are called σ (sigma) bonds. All single bonds are σ-bonds. Bonds made with unhybridized *p* orbitals are called π (pi) bonds. In multiple bonds (= or ≡) one bond is a σ-bond and the others are π-bonds. Sigma bonds result from the end-to-end overlap of hybridized orbitals. The bonding electrons are held directly between the two nuclei. Pi bonds result from the side-to-side overlap of *p* orbitals. The bonding electrons are held above and below the two nuclei.



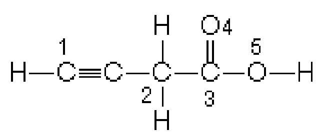
The top figure shows the electronic structure of CH_2O. The C and O are both sp^2 hybridized. Two of the orbitals on O are populated with lone pair e^-. The bond angle between the hybridized orbitals is **120°** making this a **trigonal planar** structure. The unhybridized *p* orbital is at 90° to the sp^2 orbitals



The bottom figure shows the electronic structure of N_2. Both N are *sp* hybridized, with **180°** between these orbitals. One hybridized orbital forms a σ-bond and a lone pair of e^- occupies the other. The two *p* orbitals on each N form two π bonds between the atoms.

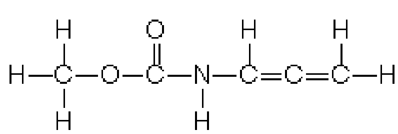
The four sp^3 orbitals lie at **109.5°** to each other and form a **tetrahedron**.

4. For the compound below, determine the hybridization, bond angles and lone pair e^- at each numbered atom.



	hybridization	bond angle	lone pair e^-
1			
2			
3			
4			
5			

5. For the compound below determine



a. How many sp^3, sp^2 and *sp* hybridized atoms are present

b. How many lone pairs of e^- are present

c. How many σ-bonds are present

d. How many π-bonds are present

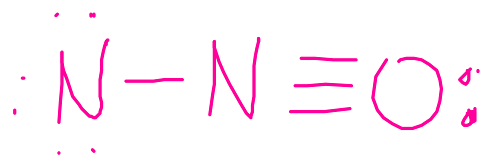
It is often possible to draw more than one Lewis structure for a compound. For example, N_2O can be drawn as:

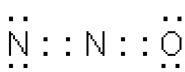
To determine which structure is the most reasonable, the **formal charge** on each of the atoms in the molecules can determined. The formal charge is an indication of the electron density at the atoms in a particular structure. It is determined as:

formal charge = valence e^- - lone pair e^- - ½ shared e^-

Nitrogen is in Group V, and has 5 valence electrons. Oxygen is in Group VI, with 6 valence electrons.

6. Assign formal charges to all of the atoms in the three resonance structures of N_2O.

a.  

b. 

c. 

d. Based on the chemical nature of these elements, which is the worst structure and why?

e. Which is the best structure, and why?

The atomic Lewis structures of C, N and O are 

7. What is the most stable bonding pattern for

	bonds	lone pair e^-
a. C		
b. N		
c. O		