

# CHEMISTRY 121 – AUTUMN 2009

## CHAPTER 8 & 9 PROBLEMS

---



All questions listed below are problems taken from old Chemistry 121 exams given here at The Ohio State University. Read Chapters 8 and 9 and complete the following problems. These problems will not be graded or collected, but will be very similar to what you will see on an exam.

### Lewis Structures

For each of the following molecules complete the following:

- Total # of valence electrons
- Identify the Central Atom
- Identify the Ligands
- Indicate the number of bonding pairs on the central atom
- Indicate the number of lone pairs on the central atom
- Indicate the electron domain geometry of the molecule
- Indicate the molecular geometry of the molecule
- Determine the hybridization of the central atom
- Draw the molecule
- Determine if the molecule is polar or non-polar

1.  $\text{SiCl}_4$
2.  $\text{PF}_3$
3.  $\text{ClF}_3$
4.  $\text{VF}_5$
5.  $\text{SiH}_4$
6.  $\text{SF}_2$
7.  $\text{BeCl}_2$
8.  $\text{NH}_3$
9.  $\text{SF}_4$
10.  $\text{BrF}_3$
11.  $\text{XeF}_2$
12.  $\text{SF}_4$
13.  $\text{ICl}_3$
14.  $\text{XeF}_4$
15.  $\text{BrF}_5$
16.  $\text{BrF}_4^+$
17.  $\text{BrF}_4^-$
18.  $\text{ICl}_2^-$
19.  $\text{ICl}_2^+$

## Lewis Structures and Resonance Structures

For each of the following molecules complete the following:

- Total # of valence electrons
- Identify the Central Atom
- Identify the Ligands
- Indicate the number of bonding pairs on the central atom
- Indicate the number of lone pairs on the central atom
- Indicate the electron domain geometry of the molecule
- Indicate the molecular geometry of the molecule
- Determine the hybridization of the central atom
- Draw the molecule
- Predict the structure according to the octet rule
- Predict the structure according to formal charges
- Draw any appropriate resonance structures

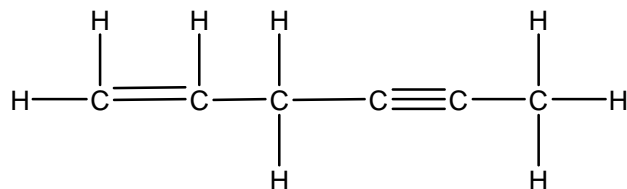
- $\text{NO}_2^-$
- $\text{NO}_3^-$
- $\text{CO}_3^{2-}$
- $\text{SO}_4^{2-}$
- $\text{SO}_3^{2-}$
- $\text{PO}_3^{2-}$
- $\text{PO}_4^{3-}$
- $\text{ClO}_3^-$
- $\text{ClO}_2^-$
- $\text{SOCl}_2$
- $\text{XeOF}_2$
- $\text{XeOF}_4$
- $\text{XeO}_3\text{F}_2$
- $\text{XeF}_4$

## Hybrid Orbitals

- List and sketch the five different electron domain geometries using 3-dimensional wedges and dashed lines when needed. Also indicate the hybridization and bond angles for each model.
- Sketch one s orbital and one p orbital (be sure to include shading when needed) and add the two wave functions together to sketch an "sp" hybrid.
- Sketch one s orbital and two p orbitals (be sure to include shading when needed) and add the three wave functions together to sketch an "sp<sup>2</sup>" hybrid.
- Sketch one s orbital and three p orbitals (be sure to include shading when needed) and add the four wave functions together to sketch an "sp<sup>3</sup>" hybrid.
- (a) Draw a  $\sigma$  bond between two s orbitals. (b) Draw a  $\sigma$  bond between two p orbitals. (c) Draw a  $\pi$  bond between two p orbitals.
- Figure 9.15 listed on page 356 of the textbook shows the potential energy of two hydrogen atoms as a function of the distance between them. (a) To what does a zero energy correspond in the diagram?

(b) According to valence bond theory, why does the energy decrease as the H atoms move from a large separation to a small one? (c) What is the significance of the H - H bond distance at the minimum point in the plot? (d) Why does the energy rise at the H - H distances less than the minimum point in the plot?

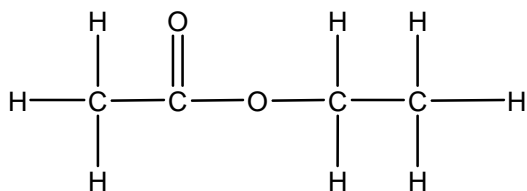
7. The molecule difluoromethane ( $\text{CH}_2\text{F}_2$ ) is used as a refrigerant. Sketch the Lewis structure and determine if the molecule has a non-zero dipole moment. If the molecule is polar, in what direction does the overall dipole moment vector point in the molecule?
8. Consider the hydrocarbon shown below.



(a) What is the hybridization at each carbon atom in the molecule? (b) How many  $\sigma$  and how many  $\pi$  bonds are there in the molecule?

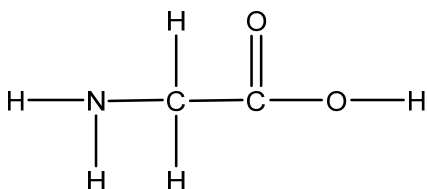
9. Consider the bonding in an  $\text{MgH}_2$  molecule. (a) Draw the Lewis structure and predict its molecular geometry. (b) Before forming the hybrid orbitals on the central atom Mg, the Mg atom must exist in an excited state. Why does this excited state have to occur to allow the Mg to hybridize? (c) What is the hybridization on the Mg central atom? (d) Sketch the two bonds between the Mg hybrids and the two H 1s orbitals. Are these  $\sigma$  or  $\pi$  bonds?
10. (a) Draw a Lewis structure for  $\text{SiH}_4$  and predict its molecular geometry. (b) What is the hybridization of the central atom Si in  $\text{SiH}_4$ ? (c) What must happen to a ground state Si atom in order for hybridization to occur? (d) Sketch the four bonds that occur between the hybrid orbitals on the Si central atom and the four H 1s orbitals. (e) What is the bond angle that forms between each bond?
11. (a) If the valence atomic orbitals on an atom are  $sp$  hybridized, how many unhybridized p orbitals remain in the valence shell? How many  $\pi$  bonds can the atom form? (b) If the valence atomic orbitals on an atom are  $sp^2$  hybridized, how many unhybridized p orbitals remain in the valence shell? How many  $\pi$  bonds can the atom form? (c) If the valence atomic orbitals on an atom are  $sp^3$  hybridized, how many unhybridized p orbitals remain in the valence shell? How many  $\pi$  bonds can the atom form?
12. (a) Draw Lewis structures for methane ( $\text{CH}_4$ ) and formaldehyde ( $\text{H}_2\text{CO}$ ). (b) What is the hybridization at the carbon atom in  $\text{CH}_4$  and  $\text{H}_2\text{CO}$ ? (c) The carbon atom in  $\text{H}_2\text{CO}$  can form multiple bonds, whereas  $\text{CH}_4$  can only form single bonds. Explain this observation using the hybridization at the central carbon atom.
13. The nitrogen atoms in  $\text{N}_2$  form multiple bonds, while the nitrogen atoms in  $\text{N}_2\text{H}_4$  do not. How can you explain this observation?

14. Ethyl acetate,  $C_4H_8O_2$ , is shown below.



(a) What is the hybridization at each of the carbon atoms in this molecule? (b) How many  $\sigma$  and how many  $\pi$  bonds are there in the molecule? (c) What are the bond angles about each carbon atom?

15. Consider the Lewis structure for glycine shown below.



(a) What are the approximate bond angles and hybridizations about each of the carbon and nitrogen atoms in the molecule? (b) How many  $\sigma$  and how many  $\pi$  bonds are there in the molecule?

16. (a) What is the difference between a localized and delocalized  $\pi$  bond in a molecule? (b) How can you determine whether a molecule or ion will exhibit delocalized  $\pi$  bonding? (c) Is the  $\pi$  bond in  $NO_2^-$  localized or delocalized?
17. (a) What are the similarities and differences between atomic orbitals and molecular orbitals? (b) Why are the electrons in a bonding molecular orbital of  $H_2$  at lower energy than the electron in a ground state hydrogen atom? (c) How many electrons can be placed in each molecular orbital of a molecule?
18. (a) Why is the anti-bonding molecular orbital in an  $H_2$  molecule at a higher energy than the 1s orbital in a ground state hydrogen atom? (b) If two p orbitals on one atom combine with two p orbitals on another atom, how many molecular orbitals result?
19. (a) Sketch the  $\sigma$  and  $\sigma^*$  molecular orbitals resulting from the combination of two  $p_z$  atomic orbitals. (b) Sketch the  $\pi$  and  $\pi^*$  MOs resulting from the combination of two  $p_y$  orbitals. (c) Assuming no mixing of the s and p orbitals, arrange the MOs from part (a) and (b) in order of increasing energy.
20. (a) What is the relationship between bond order, bond length, and bond energy? (b) According to molecular orbital theory, would  $Be_2$  or  $Be_2^+$  be more stable? Why? Would you expect either molecule to exist?
21. (a) What is meant by the term diamagnetism? (b) How does a diamagnetic substance respond to a magnetic field? (c) Which of the following would you expect to be diamagnetic:  $N_2^-$ ,  $O_2^{2-}$ ,  $Be_2^+$ ,  $C_2^-$ ?
22. (a) What is meant by the term paramagnetism? (b) How can you determine experimentally whether a substance is paramagnetic? (c) Which of the following would you expect to be paramagnetic:  $O_2^+$ ,  $N_2^{2-}$ ,  $Li_2^+$ ,  $O_2^{2-}$ ?
23. Draw out the molecular orbital diagram for  $B_2^+$ ,  $Li_2^+$ ,  $N_2^+$ , and  $Ne_2^+$ . Would the addition of an electron stabilize or destabilize the resulting molecule?

24. Predict the bond order and magnetism of  $\text{CO}^+$ ,  $\text{NO}^-$ ,  $\text{OF}^+$ , and  $\text{NeF}^+$ .
25. Calculate the bond order and determine the magnetism for  $\text{CN}^+$ ,  $\text{CN}$ , and  $\text{CN}^-$ . Which molecule/ion would molecular orbital theory predict to be the most stable?
26. The nitric oxide molecule readily loses one electron to form the  $\text{NO}^+$  ion. (a) Using molecular orbital theory, explain why this observation makes sense chemically. (b) Arrange the following in order of increasing bond strength:  $\text{NO}$ ,  $\text{NO}^+$ ,  $\text{NO}^-$  and indicate if each molecule/ion is paramagnetic or diamagnetic. (c) Isoelectronic species have the same number of electrons. With what neutral homonuclear diatomic molecule are the  $\text{NO}^+$  and  $\text{NO}^-$  ions isoelectronic?
27. (a) Draw the molecular orbital diagram for  $\text{IBr}$  and determine its bond order. (b) Are the atomic orbitals for I or Br lower in energy? Why?
28. The molecules  $\text{SiF}_4$ ,  $\text{SF}_4$ , and  $\text{XeF}_4$  all have the same molecular formula  $\text{AX}_4$  yet they have different geometries. Predict the shapes of each molecule and explain why the shapes are different.
29. The  $\text{PF}_3$  molecule has a dipole moment of 1.03 D, but the  $\text{BF}_3$  molecule has a dipole moment of zero. How can you explain this difference?
30. It is often observed that  $\sigma$  bonds involving an  $sp$  hybrid are shorter than those involving  $sp^2$  or  $sp^3$  hybrid orbitals. Can you propose a reason for this?
31. The magnetic properties of a compound can be measured using an instrument called a Guoy balance. When the magnetic measurements are made, the sample is generally covered by an atmosphere of pure nitrogen gas rather than in air? Why do you suppose chemists go through the hassle of using a nitrogen atmosphere in these measurements?