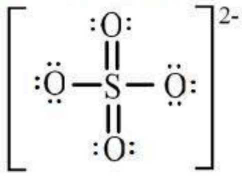
Molecular and Ionic Compound Structure and Properties
2.7 VESPR and Bond Hybridization
Worksheet Key

1) Draw Lewis structures, predict the shape, and give the approximate bond angle for each of the following compounds:

a. Sulfate ion



Tetrahedral, 109.5°

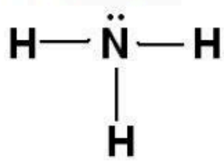
b. Water

Bent, 104.5°

c. Carbon dioxide

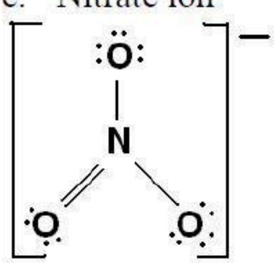
Linear, 180°

d. Ammonia



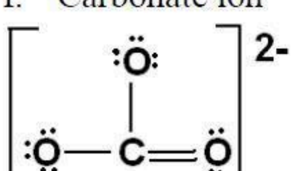
Trigonal Pyramidal, 109.5°

e. Nitrate ion



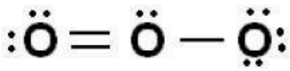
Trigonal Planar, 120°

f. Carbonate ion



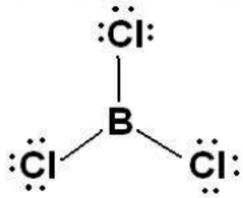
Trigonal Planar, 120°

g. Ozone (O_3)



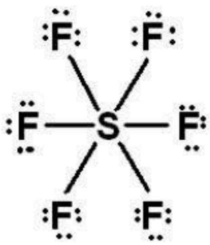
Bent, 120°

h. Boron trichloride



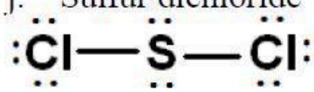
Trigonal Planar, 120°

i. Sulfur hexafluoride



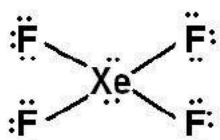
Octahedral, 90°

j. Sulfur dichloride



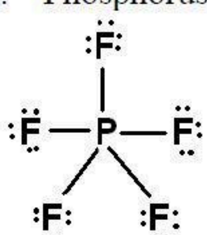
Bent, 104.5°

k. Xenon tetrafluoride



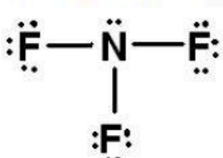
Square Planar, 90°

l. Phosphorus pentafluoride



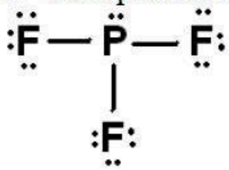
Trigonal Bipyramidal, 90° and 120°

m. Nitrogen trifluoride



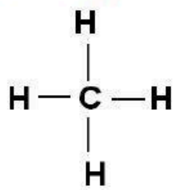
Trigonal Pyramidal, 109.5°

n. Phosphorus trifluoride



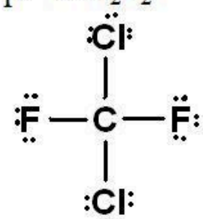
Trigonal Pyramidal, 109.5°

o. CH_4



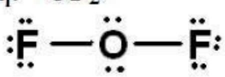
Tetrahedral, 109.5°

p. CCl_2F_2



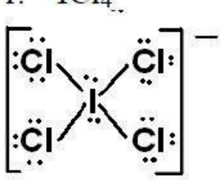
Tetrahedral, 109.5°

q. OF_2



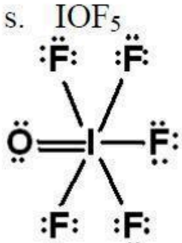
Bent, 104.5°

r. ICl_4^-



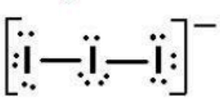
Square Planar, 90°

s. IOF_5



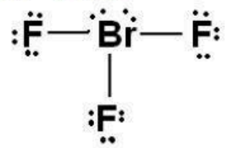
Octahedral, 90°

t. I_3^-



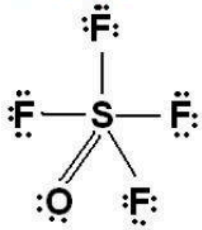
Linear, 180°

u. BrF_3



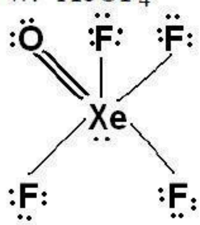
T-Shaped, 90°

v. SOF_4



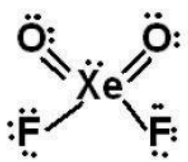
Trigonal Bipyramidal, 90° and 120°

w. $XeOF_4$



Square Pyramidal, 90°

x. XeO_2F_2



Seesaw, 90° and 120°

2) The charge clouds of both methane (CH_4) and ammonia (NH_3) are arranged in a tetrahedral geometry. Explain why the actual bond angles in methane (CH_4) are 109.5°, while the actual bond angles in ammonia (NH_3) are 107.3°.

All four charge-clouds around the carbon atom in methane are bonds, and thus, they are the same distance apart and repel one another with the same force. Ammonia, on the other hand, has three bonds and one lone pair. The two lone electrons on ammonia spend most of their time closer to the nitrogen atom than the bonding pairs. This also puts the lone pair in a position that is closer to the bonding pairs. Because of this, the lone pair exerts a slightly greater repulsive force on the bonding pairs, which causes them to move a little closer together.

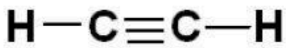
3) Explain why CH_4 has a tetrahedral geometry, while SF_4 has a seesaw shape.

The central atom in methane is surrounded by four charge-clouds. All of these charge clouds are bonding pairs. The four bonding pairs repel one another with the same force, causing them to move as far away from one another as possible. This results in a

tetrahedral geometry with bond angles of 109.5°.
The central atom in SF_4 is surrounded by five charge-clouds. Four of these charge clouds are bonding pairs, and one is a lone pair. A seesaw shaped geometry results from this configuration.

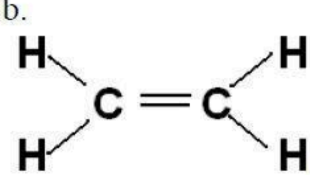
4) In the following structures, identify:
   i. The shape around each carbon atom
   ii. The bond angles emerging from each carbon atom.
   (Hint: Look at each carbon individually. Every atom that is bonded to the carbon you are looking at is considered to be a terminal atom – even other carbons. After finding the shape around the first carbon atom, move to the second carbon and treat it as the central atom.)

a.



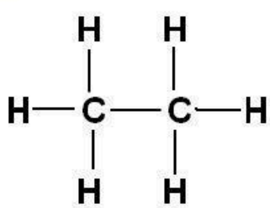
1) Linear, 180°; 2) linear, 180°

b.



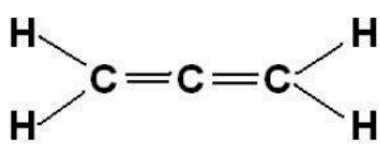
1) Trigonal planar, 120°; 2) Trigonal planar, 120°

c.



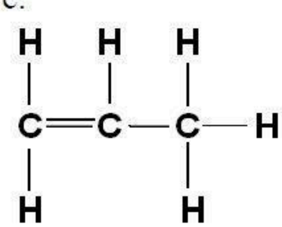
1) Tetrahedral 109.5°; 2) Tetrahedral 109.5°

d.



1) Trigonal planar, 120°; 2) Linear, 180°; 3) Trigonal planar, 120°

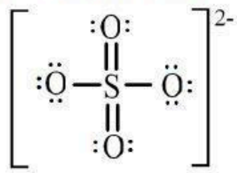
e.



1) Trigonal Planar, 120°; 2) Trigonal Planar, 120°; 3) Tetrahedral, 109.5°

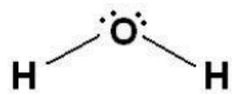
5) Identify the type of hybridization (type of hybrid orbitals) around the central atom in each of the structures listed below. You drew the Lewis diagrams for these molecules in question 1 of this worksheet.

a. Sulfate ion



4 Charge Clouds: sp^3 hybridization

b. Water

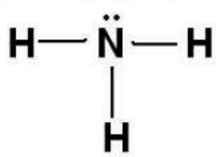


4 Charge Clouds: sp^3 hybridization

c. Carbon dioxide

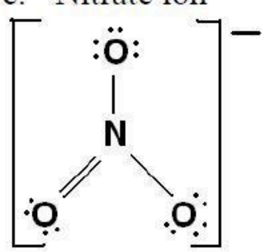
2 Charge Clouds: sp hybridization

d. Ammonia



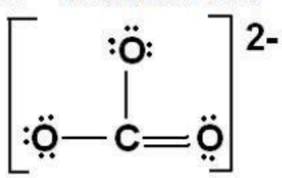
4 Charge Clouds: sp^3 hybridization

e. Nitrate ion



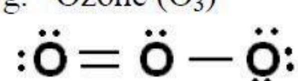
3 Charge Clouds: sp^2 hybridization

f. Carbonate ion



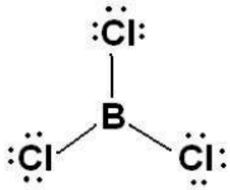
3 Charge Clouds: sp^2 hybridization

g. Ozone (O_3)



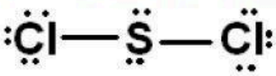
3 Charge Clouds: sp^2 hybridization

h. Boron trichloride



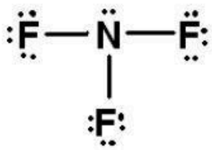
3 Charge Clouds: sp^2 hybridization

i. Sulfur dichloride



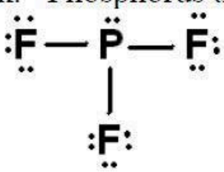
4 Charge Clouds: sp^3 hybridization

j. Nitrogen trifluoride



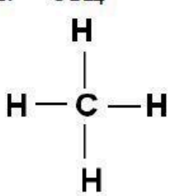
4 Charge Clouds: sp^3 hybridization

k. Phosphorus trifluoride



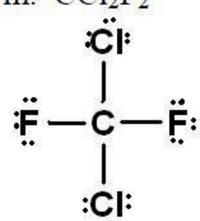
4 Charge Clouds: sp^3 hybridization

l. CH_4



4 Charge Clouds: sp^3 hybridization

m. CCl_2F_2



4 Charge Clouds: sp^3 hybridization

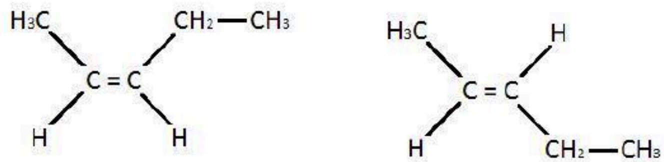
n. OF_2

4 Charge Clouds: sp^3 hybridization

6) The following questions pertain to σ-bonds and a π-bonds.
   a. Which type of bond has a greater amount of bond energy? Justify your answer.

σ-bonds have a greater amount of bond energy. They result from the direct overlap of orbitals, and pi-bonds result from the side to side interaction between unmorphed p-orbitals. Because of this, σ-bonds have a shorter bond length, and thus, a greater amount of bond energy..

   b. Which type of bond causes the following isomers? Justify your answer.



Pi-bonds cause the isomers above as they are rigid and prevent sigma-bonds from spinning.

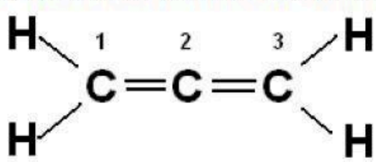
   c. Extended networks of π-bonds can cause a delocalization of electrons. Explain how this occurs.

Electrons are being shared in each pi-bond. If the pi-bonds flip-flop back and forth between adjacent p-orbitals on two sides of an atom, the shared electrons in the p-orbitals can become delocalized.

   d. What is the difference between a σ-bond and a π-bond?

σ-bonds result from the direct overlap of orbitals. Pi-bonds result from the side to side interaction between unmorphed p-orbitals. Pi-bonds are only present in double and triple bonds, whereas σ-bonds are present in single, double, and triple bonds.

7) Use the structure below to answer the following questions.



   a. What type of hybrid orbitals form around the first carbon?

sp^2

   b. What type of hybrid orbitals form around the second carbon?

sp

c. How is the double bond between the first and second carbon formed? Reference the types of orbitals involved in the bonds and the methods of bonding.

The double bond contains a σ-bond and a π-bond. The σ-bond results from the overlap of the sp^2 hybrid orbital on the first carbon and the sp hybrid orbital on the second carbon. The π-bond results from the side-to-side interaction between the unmorphed p-orbital on the first carbon and one of the unmorphed p-orbitals on the second carbon.

8) Use the structure below to answer the following questions.

$$H-C\equiv N:$$

a. What type of hybrid orbitals form around the carbon?

sp

b. How many σ–bonds are there between the carbon and the nitrogen?

1

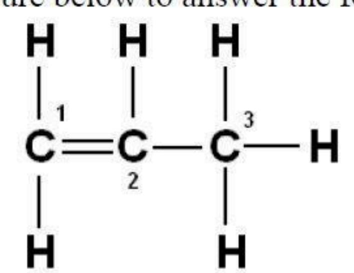
c. How many π-bonds are there between the carbon and the nitrogen?

2

d. How many unmorphed p-orbitals remain on the carbon atom in this compound?

2

9) Use the structure below to answer the following questions.



a. What type of hybrid orbitals form around the second carbon?

sp^2

b. What type of hydrid orbitals form around the third carbon?

sp^3

c. How many unmorphed p-orbitals remain around the first carbon?

1

d. How many σ-bonds exist in the entire structure?

There are 8-σ bonds in this molecule. Every bond contains one σ bond. Double bonds have a σ-bonds and a π-bond.

e. How many π-bonds exist in the entire structure?

There is one π-bond between the first and second carbon.

10) Is the sulfate ion [question 1) a.] polar or non-polar? Explain by discussing shape, polar bonds, and whether or not dipoles cancel.

The sulfate ion has a tetrahedral geometry where sulfur is bonded to four oxygen atoms and has no lone pairs. The individual bonds are polar, as oxygen is more electronegative than sulfur. However, the overall compound is **non-polar**, as the dipoles cancel due to the symmetrical arrangement of the four S – O bonds.

11) Is water [question 1) b.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The water molecule has a bent geometry where the oxygen atom is bonded to two hydrogen atoms and has two lone pairs. The individual bonds are polar, as oxygen is more electronegative than hydrogen. The overall compound is also **polar**, as the dipoles do not cancel due to the asymmetrical arrangement of the two O – H bonds. The oxygen end is slightly negative and the hydrogen ends are slightly positive.

12) Is ammonia [question 1) d.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The ammonia molecule has a trigonal pyramidal geometry where the nitrogen atom is bonded to three hydrogen atoms and has one lone pairs. The individual bonds are polar, as nitrogen is more electronegative than hydrogen. The overall compound is also **polar**, as the dipoles do not cancel due to the asymmetrical arrangement of the three N – H bonds. The nitrogen end is slightly negative and the hydrogen ends are slightly positive.

13) Is O_3 [question 1) g.] polar or non-polar? Explain by discussing shape, dipolar bonds, and whether or not dipoles cancel.

The ozone molecule has a bent geometry where the central atom is bonded to two other atoms and has one lone pairs. The individual bonds are non-polar, as the electronegativity difference in an O – O bond is zero. Therefore, the overall compound is also **non-polar**, as there are no dipoles.

14) Is boron trichloride [question 1) h.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The boron trichloride molecule has a trigonal planar geometry where the boron atom is bonded to three chlorine atoms and has no lone pairs. The individual bonds are polar, as chlorine is more electronegative than boron. However, the overall compound is **non-polar**, as the dipoles cancel due to the symmetrical arrangement of the three B – Cl bonds.

15) Is sulfur hexafluoride [question 1) i.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The sulfur hexafluoride molecule has an octahedral geometry where the sulfur atom is bonded to six hydrogen atoms and has no lone pairs. The individual bonds are polar, as fluorine is more electronegative than sulfur. The overall compound is **non-polar**, as the dipoles cancel due to the symmetrical arrangement of the six S – F bonds.

16) Is xenon tetrafluoride [question 1) k.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The xenon tetrafluoride molecule has a square planar geometry where the xenon atom is bonded to four fluorine atoms and has two lone pairs. The individual bonds are polar, as fluorine is more electronegative than xenon. The overall compound is **non-polar**, as the dipoles cancel. The dipoles are equal and opposite to one another in the same plane.

17) Is phosphorus pentafluoride [question 1) l.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The phosphorus pentafluoride molecule has a trigonal bipyramidal geometry where the phosphorus atom is bonded to five fluorine atoms and has no lone pairs. The individual bonds are polar, as fluorine is more electronegative than phosphorus. The overall compound is **non-polar**, as the dipoles cancel due to the symmetrical arrangement of the five P – F bonds.

18) Is bromine trifluoride [question 1) u.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The bromine trifluoride molecule has a T-shaped geometry where the bromine atom is bonded to three fluorine atoms and has two lone pairs. The individual bonds are polar, as fluorine is more electronegative than bromine. The overall compound is also **polar**, as the dipoles do not cancel due to the asymmetrical arrangement of the three B – F bonds.

19) Is XeO_2F_2 [question 1) x.] polar or non-polar? Explain by discussing shape, polarity of bonds, and whether or not dipoles cancel.

The XeO_2F_2 molecule has a sea-saw geometry where the xenon atom is bonded to two fluorine atoms and two oxygen atoms. There is also two lone pairs around the xenon. The individual bonds are polar, as fluorine and oxygen are more electronegative than bromine. The overall compound is also **polar**, as the dipoles do not cancel.

20) CCl_4 is non-polar, but CH_2Cl_2 is polar? Explain why this is by discussing shape, polarity of bonds, and whether or not dipoles cancel.

Both CCl_4 and CH_2Cl_2 have tetrahedral geometries, and all of the bonds in both structures are polar. The dipoles in CCl_4 cancel, making the structure non-polar. The dipoles in CH_2Cl_2 do not cancel. Chlorine is more electronegative than carbon, so the two chlorine atoms acquire partial negative charges. Carbon is more electronegative than hydrogen, so the two hydrogen atoms acquire partial positive charges. The overall structure is polar, as it acquires a positive and a negative end.