Chapter 8 & 9 blank.

1. (3points) Write the Lewis symbol for: Br–, N, and In.
2. (20 points) Sulfur tetrafluoride (SF_4) reacts slowly with oxygen to form sulfur tetrafluoride monoxide (OSF_4) according to the following unbalanced reaction:

$$SF_{4(g)} + O_{2(g)} \rightarrow OSF_{4(g)}$$

   a. Balance the equation (1 point)
   b. Write a Lewis structure of OSF_4 in which the formal charges of all atoms are zero. (4 points)
   c. Use average bond enthalpies to estimate the enthalpy of the reaction. Is it endothermic or exothermic? (4 points)
   d. Determine the electron domain geometry of OSF_4, (2 points)
   e. Determine the molecular geometry (2 points)
   f. Draw two possible isomers for the molecule based on this geometry. (4 points
   g. Which of these structures is more likely to be observed? Explain (3 points)

yes you write out both (not same as (d))

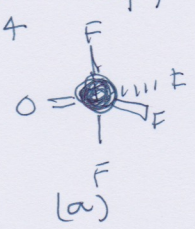
1) $[:\ddot{B}r:]^-$, $:\dot{N}\cdot$, $\cdot\dot{I}n\cdot$

2d) 5 RED = trigonal bipyramidal
AX_5

* 2e) 5 BR, 0 LP = trigonal bipyramidal

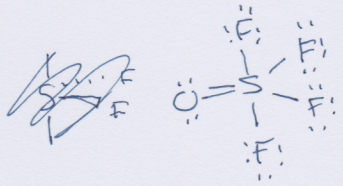
2) a) $2SF_{4(g)} + O_{2(g)} \rightarrow 2OSF_4$

2f)

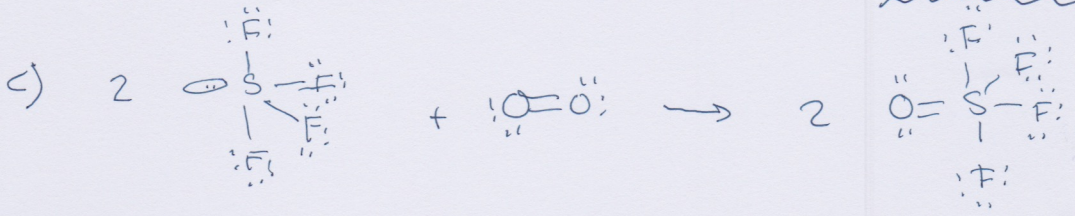


(a) is best. why?
double bonds have more e-density and repulse single bonds so, like lone pairs they go equitorial

b)



c)



8 S-F bonds + 1 O=O bond ⇒ (8 S-F bonds + 2 O=S bonds

$8\text{ moles S-F} \times \frac{327\text{ kJ}}{\text{mol S-F}} + 1\text{ mol } O_2 \times \frac{495\text{ kJ}}{\text{mol } O_2} - 8\text{ mol S-F} \times \frac{327\text{ kJ}}{\text{mol SF}} - 2\text{ mol S=O} \times \frac{523\text{ kJ}}{\text{mol S=O}}$

$\Delta H = -551\text{ kJ}$

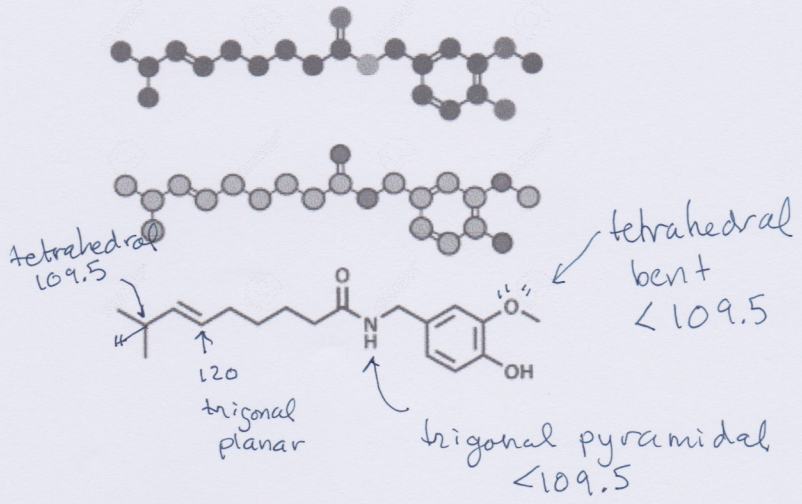
This reaction is exothermic b/c $\Delta H < 0$

3. (4 points) Capsaicin is the molecule responsible for the hot spiciness of chili peppers. A valid Lewis structure and ball and stick model is shown below. [1,2]

not on test
a. What atomic orbitals mix to form the hybrids of N?
b. In what type of hybrid orbital does the lone pairs of N reside?
c. How many π bonds are present?

But what are ∠angles?

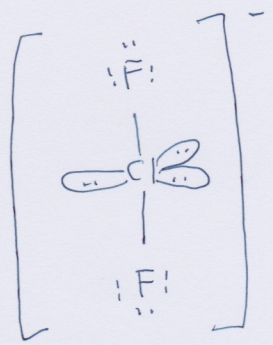
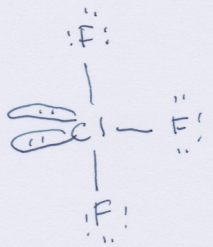
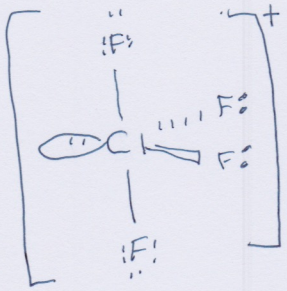
Figure 1 Lewis Structure for Capsaicin



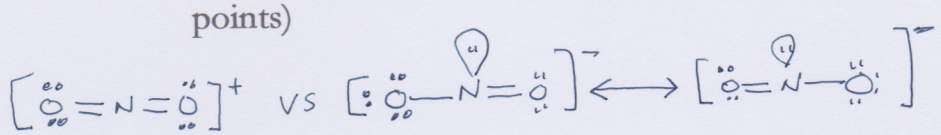
4. (4 points) Silicone tetrafluoride reacts with fluoride to produce the hexafluorosilicate ion, SiF_6^{2-}; GeF_4 behaves similarly, but CF_4 does not. Why doesn't CF_4 react with F^-; to form CF_6^{2-}.

Si is an n = 3 element. Si has the 3s, 3p, 3d subshell available for expanded octets. Carbon is an n = 2 element. Carbon does not have "2d" subshell to expand octets. Carbon must obey the octet rule; Silicon doesn't have to.

5. (6 points) Chlorine can form a variety of molecules and ions when bonded to fluorine. Three formulas are $[ClF_2]^-$, ClF_3, and $[ClF_4]^+$. Draw the Lewis structure for the three species, showing lone pairs when appropriate.

$[ClF_2]^-$	ClF_3	ClF_3
		

6. (8 points) Nitrogen can form several types of compounds when it reacts with oxygen. Two of these compounds are NO_2^+ and NO_2^-. Draw the best Lewis structure for these two ions and based on resonance (or the lack thereof) pick the ion that has the longest N=O bond. (6 points) Explain your answer. (2 points)

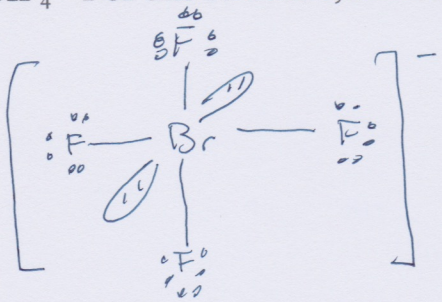


The N=O in NO_2^+ is a true double bond. It has the most e- density between N and O compared to NO_2^-. NO_2^-, due to resonance, shares e- density between two locations on the molecule, this lowers the overall e- density between each N & O. When e- density is lowered the attractive forces are lowered and the bond gets longer. So the double bond in NO_2^- is longer than a ~~regular~~ expected double bond length but shorter than a single bond

7. (15 points) Draw the Lewis structures for the following structures. Follow the instructions. Clearly label the structures.

a. BrF_4^- For this structure, draw a good Lewis structure of the ion.

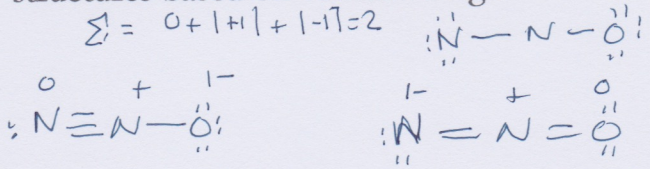
7
28
1
36



b. N_2O, Nitrogen is the central atom. The molecule is linear. There are three resonance structures. Draw them. Determine the two most stable structures based on formal charge.

10
6
16

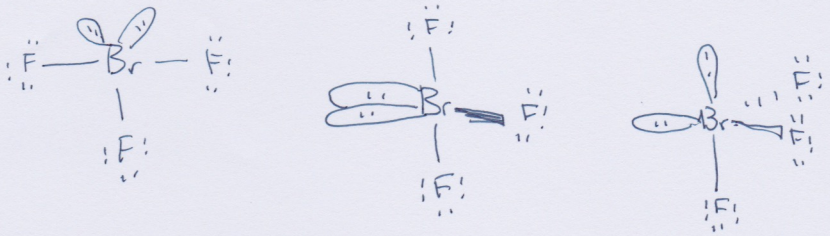
$\sum = 0+1+1+1-1-1 = 2$ $\quad \sum 1-2+1+1+1+0 = 4$



These two are the best b/c the overall number of charges is low, The magnitude of the charges is low, and ⊖ is on the more electronegative oxygen

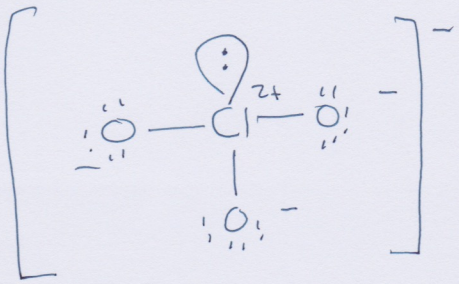
This one is terrible:
⊕ on oxygen
3 charges
$\sum = 4$

c. BrF_3 Use the molecular and electron domain geometry to help you determine where the lone pairs should go for the most stable molecular geometry



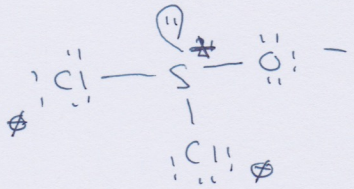
lone pairs go on the equator to minimize 90° lone pair – lone pair and lone-pair – bonding repulsions

d. ClO_3^- Draw the Lewis structure that obeys the octet rule.

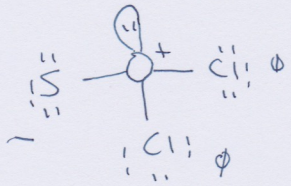


14
6
6
26

e. Cl_2SO Put S in the center and show formal charges for each atom; draw the structure that obeys the octet rule for the central atom



f. Cl_2SO Put O in the center and show formal charges for each atom; draw the structure that obeys the octet rule for the central atom



g. For structures e and f, pick the most stable structure based on formal charge rules.

e is more stable than f.

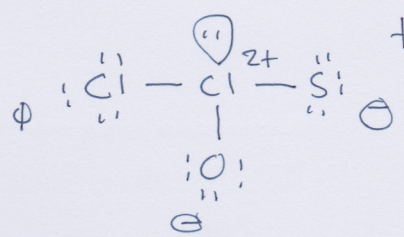
why?

✓ large atom in center

✓ positive charge on more electropositive atom

✓ − charge on more electronegative atom

BTW



this is awful!!