Key for Chem 130 – Second Exam

Name________________________________

On the following pages you will find questions that cover the structure of molecules, ions, and solids, and the different models we use to explain the nature of chemical bonding. Read each question carefully and consider how you will approach it before you put pen or pencil to paper. If you are unsure how to answer one question, then move on to another question; working on a new question may suggest an approach to the one that is more troublesome. If a question requires a written response, be sure that you answer in complete sentences and that you directly and clearly address the question.

Question 1 _____/24	Question 5 _____/12
Question 2 _____/8	Question 6 _____/12
Question 3 _____/12	Question 7 _____/8
Question 4 _____/12	Question 8 _____/12

Total _____/100

Some potentially useful equations and constants are provided here. A periodic table and other potentially useful data are provided on a separate handout.

$c = \lambda\nu$ $\qquad$ $E = h\nu$ $\qquad$ $KE = h\nu - W$

$$\frac{1}{\lambda} = 1.09737\times10^{-2}\ \text{nm}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$V \propto \frac{Q_+Q_-}{d}$$

$$\text{AVEE} = \frac{xIE_s + yIE_p + zIE_d}{x+y+z}$$

(valence shell electrons only)

$$\text{FC}_a = \text{V}_a - \text{N}_a - \frac{\text{B}_a}{2}$$

$$\delta_a = \text{V}_a - \text{N}_a - \text{B}_a\left(\frac{\text{EN}_a}{\text{EN}_a + \text{EN}_b}\right)$$

$c = 2.998\times10^{8}$ m/s $\qquad$ $h = 6.626\times10^{-34}$ Js $\qquad$ $N_A = 6.022\times10^{23}$ mol^{-1}

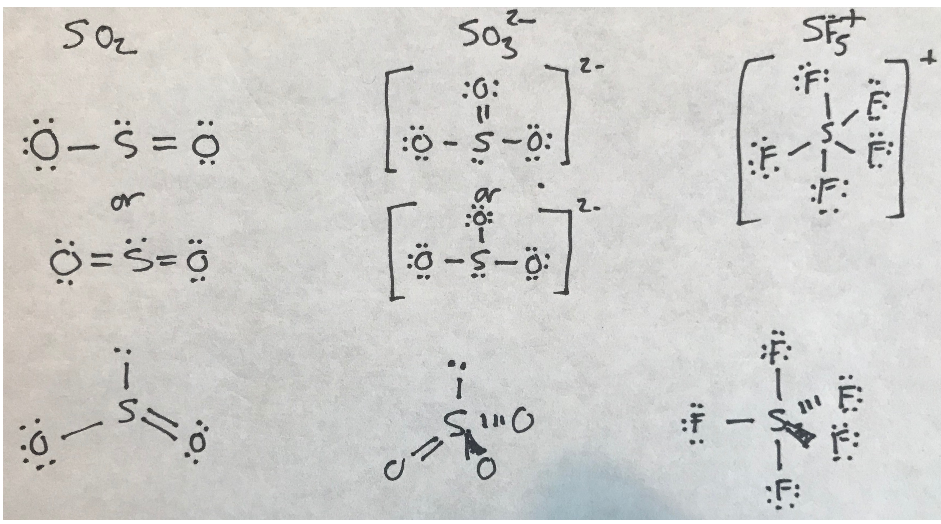
Problem 1. For each of the following, draw **one** valid Lewis structure, which need not be the "best" structure. Next, indicate the three-dimensional geometry by both drawing its VSEPR structure and by providing the name for the geometry around the underlined central atom. Then, predict whether the molecule or ion is polar or is non-polar. Finally, give the idealized bond angles for the stated bonds based on your VSEPR structure; if there is more than one possible bond angle, then list each unique bond angle and annotate your Lewis structure to indicate which is which. *Your answers for these last two items must be consistent with the bonding geometry you identify*.

See the figure below for the different ways to drawing the Lewis structures and the corresponding VSEPR structures. The number of valence electrons, the bonding geometry, the polarity, and the ideal bond angle(s) for each are listed here:

SO_2: 18 electrons; bonding geometry is bent (from a trigonal planar electron domain); polar; ideal bond angle of 120°

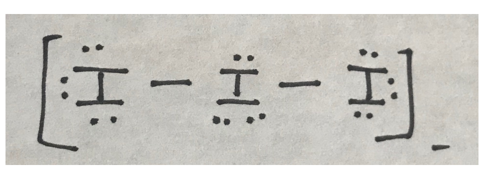
SO_3^{2-}: 26 electrons; bonding geometry is trigonal pyramidal (from a tetrahedral electron domain); polar; ideal bond angle of 109.5°

SF_5^+: 40 electrons; bonding geometry is trigonal pyramidal; non-polar; ideal bond angles of 90°, 120°, and 180°



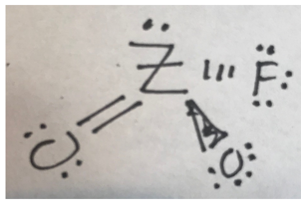
Problem 2. The triiodide anion, I_3^-, is linear. The corresponding anion of fluorine, F_3^-, however, does not exist. In 1–3 sentences, explain why it is impossible for an anion of F_3^- to form.

A Lewis structure for I_3^- shows that it has five electron domains (see figure below), which requires that the central atom, iodine in this case, accommodate more than an octet of electrons. Fluorine cannot do this; thus, the F_3^- ion is not possible.



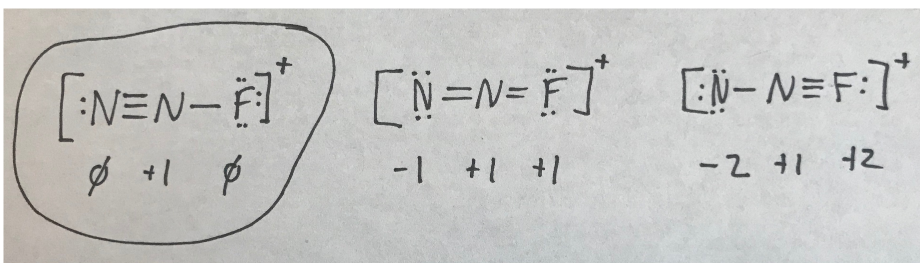
Problem 3. The element *Z*, which is an element in the first three rows of the periodic table, forms the molecule ZFO_2 with a trigonal pyramidal bonding geometry around *Z*. In addition, the length of the Z–O bond suggests it has a bond order greater than 1 but less than 2. Identify *Z* and, in 2–3 sentences, explain how you arrived at your identification.

A trigonal pyramidal structure requires three bonding domains and one non-bonding domain for a total of four electron domains. A Lewis structure that is consistent with this (see below) and that shows a single bond between *Z* and O and a double bond between *Z*O requires a total of 26 electrons. Subtracting 12 electrons for the two oxygens and seven electrons for the one fluorine leaves *Z* with seven electrons, which places it in group 18. As *Z* has 10 electrons around it (two single bonds, one double bond, and a lone-pair; *Z* cannot be F; thus, it must be Cl.



Problem 4. Consider the ion N_2F^+, which has a skeletal structure of N–N–F. Draw all possible resonance structures for this ion and annotate each of your structures with the formal charge for each atom.

See the figure below for the three possible resonance structures and formal charges. These structures must account for the 16 valence electrons in N_2F^+.



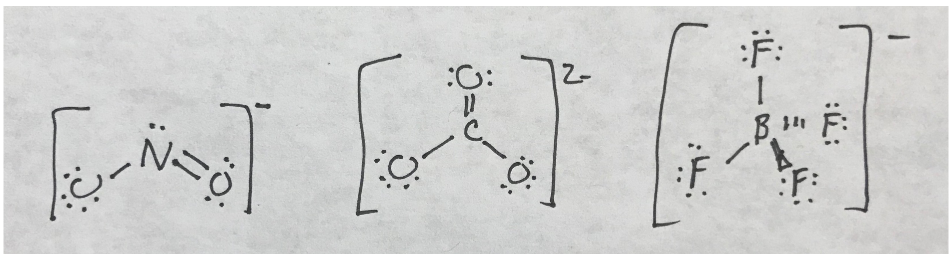
Circle the resonance structure above that provides the best overall picture of the bonding for N_2F^+ and in no more than two sentences, explain your reason for selecting this resonance structure.

Of the three resonance structures, the one that is circled has the fewest atoms with formal charges, which makes it the resonance structure that best represents the ion's bonding.

Problem 5. The underlined central atom for two of the anions below use identical hybrid orbitals to form bonds with the anion's other atoms; the remaining anion uses a different set of hybrid orbitals. Circle the anion that is different and, in 2–3 sentences explain the reason for your choice. As part of your answer, identify the specific hybrid orbitals used by each anion's central atom.

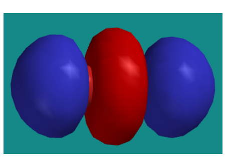
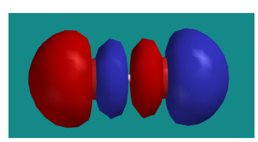
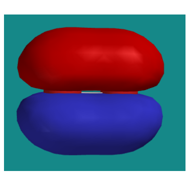
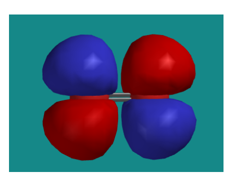
$\underline{N}O_2^-$ $\underline{C}O_3^{2-}$ $\underline{B}F_4^-$

The BF_4^- ion is the one that is different. See the figure below for the Lewis structures for these three ions, which show that the two ions on the left both contain a double bond that requires a sigma bond and a pi bond; the ion on the right, however, has only single, sigma bonds. As a result, the two ions on the left are trigonal planar and have central atoms that are sp^2 hybridized; the ion on the right, however, is tetrahedral and has a central ion that is sp^3 hybridized.



Problem 6. Molecular orbitals form when atomic orbitals interact with each other. Consider the three 2*p*-orbitals on nitrogen. When these orbitals interact with the three 2*p*-orbitals on another atom of nitrogen to form N_2, the resulting molecular orbitals are of four types: sigma bonding, sigma antibonding, pi bonding, and pi antibonding. For each of these four types of molecular orbitals, draw a picture in the table below that shows how the electrons are distributed in space relative to the two nitrogen atoms.

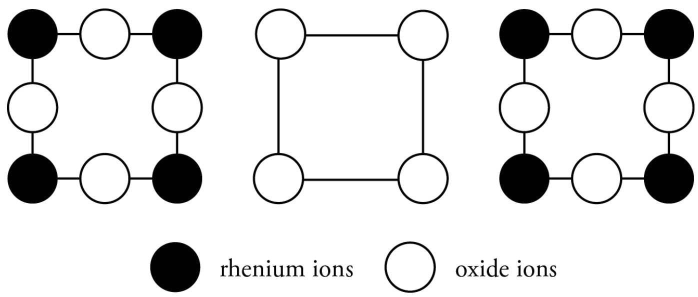
The pictures below show the different molecular orbitals. The key features your pictures needed to make clear are (1) that the electron density for sigma bond (both bonding and antibonding) is aligned along the axis that connects the two nitrogens, but that the electron density for pi bonds (both bonding and antibonding) is located above and below the axis that connects the two nitrogens; and (2) that the antibonding orbitals (both bonding and antibonding) have a node that cuts across the axis that connects the two nitrogens. Precise drawings were not necessary.

sigma bonding 	sigma antibonding 
pi bonding 	pi antibonding 

Problem 7. Consider the lead halides PbF_2, $PbCl_2$, and $PbBr_2$. Which of these compounds likely has the highest melting point? Explain your reasoning in 1–2 sentences.

Answers here needed to be clear that Coulomb's law is important, which could be made directly or inferred by noting that these are ionic, not covalent, compounds. As the cation is the same in all three cases, we focus on the anions only. The anions all have the same charge, so the distance between the ions is the key. The fluoride anion is the smallest of the three; thus, as distance is inversely proportional to the energy of attraction, PbF_2 will have the highest melting point.

Problem 8. The figure below shows three cross-sections through the lattice structure for a solid-state compound that consists of rhenium ions and oxide ions. From left-to-right, these cross-sections are at z = 0, 0.5, and 1.



What is the empirical formula for rhenium oxide. Be sure that how you arrived at your formula is clear by annotating the figure above to show the contribution of each ion. The chemical symbol for rhenium is Re.

For rhenium, there are eight ions, all on corners at one-eighth each; thus, there is one full rhenium ion. For oxide, there are 12 ions, all on edges at one-fourth each; thus, there are three full oxide ions. The empirical formula is ReO_3.

Based on your empirical formula, what is the charge on the rhenium ion? In one sentence, explain how you arrived at your answer.

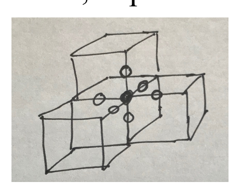
Each oxide has a charge of –2 for a total charge of –6; thus, rhenium has a charge of +6.

What kind of lattice structure do the rhenium ions exhibit? In one sentence, explain how you arrived at your answer.

The rhenium ions are on the corners of the unit cell, which means they have a lattice structure that is simple cubic.

In what kind of holes within the oxide lattice are rhenium ions found? In one sentence, explain how you arrived at your answer.

The coordination number for rhenium within the oxide lattice is six, which means that the rhenium ions are in octahedral holes. See the picture to the right for a more detailed view.



What is coordination number for the oxide ions relative to the rhenium ions? In one sentence, explain how you arrived at your answer.

Each oxide is between two rhenium ions, which gives them a coordination number of two.