**Problem Set 2**  (**Due February 20th by 7:00 PM**)

Answers to the problems in **RED** need to be submitted through the course website.

**Challenging Review Questions. (Bonus: Due 2/6)**

The first 8 ionization energies of element “X” are shown below (in aJ). Use this information to answer 1-4.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| IE1 | IE2 | IE3 | IE4 | IE5 | IE6 | IE7 | IE8 |
| 1.68 | 3.17 | 4.84 | 8.24 | 10.42 | 35.32 | 42.23 | 49.60 |

1. How many valence electrons does X have?
2. What group does X belong to?
3. Which element would be most similar to X? Zinc, Iron, Calcium, Tin, Bismuth, or Bromine?
4. Clearly explain why the difference between IE3 and IE4 is greater than the difference between IE2 and IE3.
5. Determine the wavelength of the photon needed to move an electron from the ground state of a hydrogen atom to the 7th energy level. Report your answer in nm.
6. The ionic radius of 58Ni2+ is 63 pm. Calculate its density in g mL-1. Hint: a similar problem is on the 1st exam.
7. Using your understanding of periodic trends, order these elements by increasing density (least dense to most dense). **O, N. B, F**

**Ions and Ionic Compounds (Bonus: 2/6)**

1. Predict common charge on each of the following. If more than one exist, write them all. If a common charge does not exist, select 0.

Hydrogen Strontium **Lead Aluminum** Fluorine Selenium **Argon**

1. Classify each of the following as covalent or ionic bonds:

N and N N and Na Ca and H **Sc and Cl Cl and F** **Na and H**

1. Predict the ionic compound that will form between the following:
	1. Nitrogen and sodium
	2. Magnesium and Nitrogen
	3. **Aluminum and Phosphorus**
	4. **Selenium and Potassium**
2. For each of the following lists, identify all variable charge metals.
	1. Na, Zn, Bi, Re, Mn, In
	2. **Ru, Cd, Sn, Zr, K, Al**
3. Determine the electron configuration for each of the following. You may use shorthand notation.

Fe+3  **Fe+2**In+1 **In+3** **Cu+1** Cu+2  Mn+2 **Mn+7**

1. Name each of the following:

**NaH** BeF2 **(NH4)2O** Na2CO3 CrCl6 **Fe2S3**

1. Determine the molecular formula:

Zinc Chloride **Thallium (I) Phosphate** Thallium (III) Phosphide **Tin (IV) Oxide**

**Sodium nitride** Sodium nitrite Sodium nitrate Iron (III) Selenide

**Molecular Compounds (Bonus: 2/13)**

You will need to draw the Lewis Structure for each compound to answer most of these questions

Answer problem 15-19 for each of these compounds.

**ClO2-**ClO3- **CO2** CO32- PCl3 **NO2-** O3**S3** NO3- **I2F-**

1. What is the central atom of the compound?
2. How many double bonds are present in the compound?
3. How many lone pairs are on each atom?
4. What is the formal charge on each atom?
5. Does this molecule have resonance forms? If yes, how many resonance forms exist?
6. For each pair, determine which molecule is the most stable. Give a brief explanation why you made your selection.
	1. BrO3- vs. FO3-
	2. NO4-3 vs. PO4-3
	3. **I3- vs. F3-**
7. For each of the following situations, determine the molecular formula for a neutral molecule **made from only nitrogen and chlorine** that contains:
	1. A single bond between the nitrogen atoms.
	2. **A double bond between the nitrogen atoms.**
8. Name each of the following compounds.

N2O2 N2O4 **C2H6** **SF6**

1. For each pair, pick the molecule that has the **strongest** bond between the indicated atoms.
	1. O2 vs H2O2 Compare O-O bonds
	2. N2 vs. N2H2 Compare N-N bonds
	3. **C2H6 vs. C2H4 Compare C-C bonds**
2. For each pair, pick the molecule that has the **longest** bond between the indicated atoms.
	1. O2 vs H2O2 Compare O-O bonds
	2. N2 vs. N2H2 Compare N-N bonds
	3. **C2H6 vs. C2H4 Compare C-C bonds**
3. Name each of the following:
	1. HCl
	2. H2SO4
	3. **HNO3**
	4. **HBrO2**
	5. **HBrO4**
	6. **HBrO**
4. Write the chemical formula for each compound.
	1. Nitrous acid
	2. **Sulfurous acid**
	3. **Carbonic acid**
	4. **Hydrobromic acid**
	5. **Acetic acid**

**Molecular Geometry, Polarity, and Hybridization (Bonus: 2/16)**

1. What is the hybridization of each **boldfaced atom** in:

**C**Cl3**F** **N**H3 **Br3-1  Xe**F4 **S**F4 **SeF6** **S**O3-2

1. For each of the following pairs, determine which has smaller bond angles.
	1. CH2O vs. CH4 compare H-C-H bond angles
	2. NH4+ vs. NH3 compare H-N-H bond angles
	3. **SO3 vs. SO2 compare O-S-O bond angles**
	4. **SOF4 vs. SF4 compare F-S-F bond angles**
2. Draw the Lewis structure for each molecule listed below and answer each of following questions about the **central atom**.
	1. Determine the electron geometry.
	2. Determine the molecular geometry.
	3. Determine the hybridization.

BrF3 **PH3** Br3-1  **XeF4** SF4 SeF6 SO3-2

1. Determine how many sigma bonds and pi bonds are present between the indicated atoms.
	1. CH2O C and O
	2. **HCN C and N**
2. Determine if each of the following molecules are polar:

CH3OCH3  (C-O-C connectivity) CBr3F **XeF2 CH2NH**

1. For each of the following molecules, identify ALL intermolecular forces that stabilize condensed phases.

H2O CH3CH2CH2CH3 **PH3 HF** HCl NaCl

1. For each of the following pairs, determine which has the higher melting temperature.

CH4 or CF4 CF4 or CBr3F PH3 or NH3 **HF or HCl** **TeH2 or SeH2**

1. Use Molecular Orbital (MO) Theory to determine the **bond order and number of unpaired electrons** for each of the following diatomic atoms. Recall that the orbital order changes when Oxygen or Fluorine are involved:

MO order for just carbon and/or nitrogen: 2s, 2s\*, 2p, 2p, 2p\*, 2p\*

MO order when oxygen and/or fluorine are part of the molecule: 2s, 2s\*, 2p, 2p, 2p\*, 2p\*

C2 **CO** CN-1 N2  **O2** NO F2 **OF+1** OF-1

**Challenge Questions**

Submit your answers to this question **directly to me for bonus points**. You are strongly encouraged to stop by my office with questions.

1. It is possible to mathematically predict if a compound will be ionic or covalent using measurable values for several physical properties that we’ve discussed in class. To a first approximation, this can be accomplished by considering:
* the amount of energy needed to ionize the cation X 🡪 X+ + *e-*
* the amount of energy gained when the anion forms X + *e-* 🡪 X-
* the charge stabilization gained when the two ions interact $E\_{p}=231 aJ∙pm\left(\frac{q\_{1}q\_{2}}{r}\right)$

Using the data in the table below, determine which of the following ionic compounds forms most favorably. Clearly justify your answer.

NaF, NaCl, NaBr, KF, KCl, or KBr

|  |  |  |  |
| --- | --- | --- | --- |
| Atom | Ionization Energy 1 (aJ) | Electron Affinity 1 (aJ) | Ionic Radius (pm) |
| Sodium | 0.823 |  | 105 |
| Potassium | 0.695 |  | 138 |
| Chlorine |  | 0.5795 | 181 |
| Fluorine |  | 0.5449 | 113 |
| Bromine |  | 0.5688 | 196 |

1. The following are descriptions of two different compounds. Your task is to determine the Lewis structure of the compound.
	1. This monovalent anion (meaning a -1 charge) consists of a neutral central atom from the 4th shell with trigonal pyramidal geometry. It is bonded to two halogens from the 3rd shell and one shell 2 element that carries a -1 formal charge. The central atom has one pi bond.
	2. This monovalent anion consists of a neutral central atom from the 5th shell with square pyramidal geometry. It is covalently bonded to two different types of atoms from the 2nd shell, none of which carry a permanent formal charge of -1. One pi bond exists in this molecule and two resonance forms can be drawn.

**Black Problems:**

|  |  |
| --- | --- |
| 8. H+1 or H-1 Sr+2 F-1 Se -210. Na3N Mg3N212. Fe3+: [Ar] 3d5 In+: [Kr] 5s24d10 Cu+2: [Ar] 3d9 Mn+2: [Ar] 3d5 14. ZnCl2 TlP NaNO2 NaNO316. ClO3- = 2 CO2 = 2 CO32- = 1 PCl3 = 0 O3 = 1 NO3- = 118. ClO3- 🡪 Cl = 0 the oxygen double bond = 0, oxygen with single bond = -1 CO2 🡪 C = 0 O = 0 CO32- 🡪 C = 0 the oxygen double bond = 0, oxygen with single bond = -1 PCl3 🡪 P = 0 Cl = 0 O3 🡪 central O = +1, the oxygen double bond = 0, oxygen with single bond = -1 NO3- 🡪 N = +1 the oxygen double bond = 0, oxygen with single bond = -120. a. BrO3- because expanded the octet allow formal charge to be minimized. F cannot expand the octet.b. PO4-3 because expanded the octet allow formal charge to be minimized. N cannot expand the octet.22 dinitrogen dioxide dinitrogen tetraoxide24. H2O2 🡪 single bonds are longer than double bonds N2H2 🡪 double bonds are longer than triple bonds26. a. HNO228. CH4 Tetrahedral vs. trig. planar  NH3 because of lone pair30. a. Double bond 🡪 one sigma and one pi32 H2O (London Dispersion, dipole-dipole, H-bond) CH3CH2CH2CH3 (London Dispersion) HCl (London Dispersion, dipole-dipole) NaCl (London Dispersion, Ion-Ion)34. C2 – BO=2, 0 unpaired; CN- – BO=3, 0 unpaired N2 – BO=3, 0 unpaired; NO – BO=2.5, 1 unpaired F2 – BO=1, 0 unpaired; OF-1 – BO=1, 0 unpaired | 9. N and N 🡪 covalent N and Na 🡪 ionic Ca and H 🡪 ionic11. Bi, Re, Mn, In13. Beryllium fluoride sodium carbonate chromium (VI) chloride15. ClO3- = Cl CO2 = C CO32- = C PCl3 = P O3 = O NO3- = N17. ClO3- 🡪 Cl = 1 1 oxygen has 3 and two oxygens have 2 CO2 🡪 C = 0 O = 2 CO32- 🡪 C = 0 2 oxygens have 3 and one oxygen has 2 PCl3 🡪 P = 1 Cl = 3 O3 🡪 central O has 1, one outer O has 2 and the other has 3 NO3- 🡪 N = 0 two O have 3, one O has 119. ClO3- = 3 CO2 = 1 CO32- = 3 PCl3 = 1 O3 = 2 NO3- = 321. N2Cl423. a. O2 (double bonds are stronger than single bonds) b. N2 (triple bonds are stronger than double bonds)25. a. Hydrochloric acid b. Sulfuric acid27. **C**Cl3**F** 🡪 C = sp3, F = sp3 **N**H3 🡪 sp3  Br3-1 🡪 terminal bromines are sp3, central is sp3d **S**F4 🡪sp3d **S**O3-2 🡪sp329. BrF3 a. Trig. Bipyramidal b. T-shaped c.sp3d Br3-1 a. Trig. Bipyramidal b. linear c.sp3d SF4 a. Trig. Bipyramidal b. seesaw c.sp3d SeF6 a. Octahedral b. Octahedral c.sp3d2 SO3-2 a. Tetrahedral b. Trig. pyramidal c.sp331. CH3OCH3 🡪 yes CBr3F 🡪 yes33. CF4 because it’s bigger = stronger LDF CBr3F because it’s polar NH3 because it can H-bond |