Chem 130 – Key For First Exam

Name________________________________

On the following pages you will find questions that cover various topics ranging from nomenclature to periodic properties, and from electromagnetic radiation to the quantum model of the atom. 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. *Of particular importance for this exam: if a question asks you to explain a periodic trend, it is insufficient to write that "the <insert your property> of atoms increases to the right and to the top of the periodic table." Instead, your answer must explain why this trend exists.*

Partial credit is willingly given on all problems so be sure to answer all questions!

Question 1 _____/16	Question 5 _____/13
Question 2 _____/6	Question 6 _____/13
Question 3 _____/13	Question 7 _____/13
Question 4 _____/13	Question 8 _____/13

Total _____/100

Useful equations, constants, Slater's rules, and a periodic table are provided on a separate handout.

Problem 1. For each of the following, provide **one** example of an element that fulfills the stated condition. If no element meets the condition, then write NONE. Limit your elements to those in the first five rows of the periodic table (H through Xe). *Do not use any element more than once!* The possible elements are listed below with some additional comments where appropriate.

(a) is a transition metal: Sc to Zn, and Y to Cd

(b) has a core electron configuration of [Kr]: Rb to Xe

(c) forms a monoatomic ion with a charge of –2: O, S, Se, Te (forms noble gas configuration)

(d) has exactly five electrons in a *d*-orbital: Mn & Tc (s^2d^5); also Cr & Mo (s^1d^5)

(e) forms a monoatomic ion with a charge of +1: H to Rb (Group 1); also Ga, In, Cu

(f) has exactly two unpaired electrons: Ti & Zr (d^2); C, S, Ge, & Sn (s^2p^2); O, S, Se, & Te (s^2p^4)

(g) is in the same period as aluminum: Na to Ar (period = row)

(h) has five peaks in its photoelectron spectroscopy spectrum: Al to Ar ($1s^22s^22p^63s^23p^x$)

(i) has a valence shell that consists of only *s* electrons: H to Rb; Be to Sr; also He

(j) is a halogen with a covalent radius larger than that for chlorine: Br & I (radius increases down a group)

(k) forms an ion with a charge of +2 that has a noble gas electron configuration: Be to Sr

(l) has ten core electrons: Na to Ar (core of Ne)

(m) is an alkali metal: Li to Rb (group 1, but not hydrogen)

(n) is in the *s*-block: H to Rb, He, Be to Sr (valence electrons *s* only)

(o) has a *Z* of 30: Zn (atomic number of 30)

Problem 2. Fill in the missing information for these three compounds.

Formula	Name	Covalent or Ionic?
$CaCl_2$	calcium chloride	ionic
$Fe_3(PO_4)_2$	iron(II) phosphate	ionic
P_4O_{10}	tetraphosphorous decoxide	covalent

Problem 3. Consider two photons, one with a wavelength of 250 nm and one with a frequency of 5.0×10^{14} s^{-1}. Which photon has the greatest energy, in Joules, and what is that energy?

We know that $E = hc/\lambda = h\nu$ where λ is a wavelength and ν is the a frequency. Calculating the energy for each photon we have $E = (6.626 \times 10^{-34}$ Js)$(2.998 \times 10^{8}$ m/s)/(250×10^{-9} m) $= 7.95 \times 10^{-19}$ J for the photon with a wavelength of 250 nm and $E = (6.626 \times 10^{-34}$ Js)$(5.0 \times 10^{14}$ $s^{-1}) = 3.31 \times 10^{-19}$ J for the photon with a frequency of $.0 \times 10^{14}$ s^{-1}; thus, the first photon has the greater energy.

The ionization energy of a 4*s* electron in potassium, K, is 420 kJ/mol. Is the photon you identified above capable of ejecting this electron? Justify your response with a suitable calculation and a one sentence explanation.

Your answer here depends on your answer to the previous question. Converting the energy of the first photon to kJ/mol gives $(7.95 \times 10^{-19}$ J)(1 kJ/1000 J)$(6.022 \times 10^{23}$ $mol^{-1})$ = 479 kJ/mol, which is sufficient to remove the 4*s* electron from potassium. For the second photon, the equivalent energy is 199 kJ/mol, which is not sufficient to remove the 4*s* electron from potassium.

Problem 4. The first three ionization energies for lithium, Li, are 520 kJ/mol, 7,298 kJ/mol, and 11,815 kJ/mol. The photoelectron spectroscopy, PES, spectrum for lithium, however, has just two peaks, one at 6260 kJ/mol and one at 520 kJ/mol. In 3-5 sentences, explain why there are three ionization energies but just two peaks in the PES spectrum. As part of your answer, explain why both sets of data share a common value of 520 kJ/mol.

Lithium has three electrons in a $1s^{2}2s^{1}$ electron configuration. The three ionization energies are for the successive removal of electrons; thus, lithium has three ionization energies. A photoelectron spectrum, on the other hand, has one peak for each different type of electron, which, for lithium are a 1*s* electron and a 2*s* electron. Both sets of data have share a value of 520 kJ/mol because this is the energy of lithium's 2*s* electron, which is the lowest energy peak in the PES spectrum and the easiest electron to remove.

Problem 5. The average valence electron energy, AVEE, for the first three noble gases are

He: 2,370 kJ/mol   Ne: 2,730 kJ/mol   Ar: 1,845 kJ/mol

In general, the AVEE becomes smaller as you go down a group. In 2-4 sentences, propose an explanation for why neon's AVEE is greater than that for helium.

The AVEE is the average ionization energy of an element's valence electrons. In general, we expect AVEE values to decrease down a group because the valence electrons are at a greater distance from the nucleus and, by Coulomb's law, have less of an attraction to the nucleus. The larger AVEE for Ne relative to He, therefore, must be related to the effective nuclear charge, which is the other term in Coulomb's law. Using the simplest approach to estimating the effective nuclear charge, in which core electrons completely screen the valence electrons, the effective nuclear charge for He is 2 – 0, or 2, and the effective nuclear charge for Ne is 10 – 2 = 8; this is the reason for the increase in the average valence electron energy.

Problem 6. Consider the set of elements listed below and identify the element with the smallest covalent radius and the element with the largest covalent radius. Explain the reason(s) for your selections in 2-3 sentences. Be sure to make note of the caution on the exam's first page.

N O F P S Cl

smallest covalent radius is for fluorine (F) largest covalent radius is for phosphorous (P)

As we move across a row the increase in Z_{eff} allows the nucleus to pull electrons closer, resulting in a decrease in the radius. As we move down a column, the distance from the nucleus to the valence electrons increases, resulting in an increase in the radius. We expect that N, O, and F are smaller than P, S, and Cl, and that F is smaller than O and F, and that P is larger than S and Cl.

Problem 7. Consider the set of ion listed below and identify the largest ion and the smallest ion. Explain the reason(s) for your selections in 2-3 sentences. Be sure to make note of the caution on the exam's first page.

Na^+ Mg^{2+} Al^{3+} S^{2-} Cl^- Se^{2-}

smallest ion is Al^{3+} largest ion is Se^{2-}

A cation is always smaller than its neutral element because removing an electron allows the nucleus to pull the remaining electrons in more closely as its charge is spread out among fewer electrons; there also is less electron-electron repulsion. An anion is always smaller than its neutral element both because the nucleus' charge is now spread out over more electrons and because there is more electron-electron repulsion. With the exception of Se^{2-}, which is in the fourth row, the other elements are in the third row and, as noted in Problem 6, neutral elements become smaller across a row; thus, Al^{3+} is the smallest ion. Because Se^{2-} is in the fourth row, it is the largest ion.

Problem 8. The structure of the periodic table suggests that electrons will enter a 4*s* orbital before they will enter a 3*d* orbital. Once electrons are in both orbitals, however, it generally is easier to remove an electron from a 4*s* orbital than to remove it from a 3*d* orbital. Using manganese, Mn, as an example, apply Slater's rules to determine Z_{eff} for a 3*d* electron and for a 4*s* electron. Explain, in 1-3 sentences why your results support the observation that manganese will lose a 4*s* electron before it loses a 3*d* electron.

The electron configuration for Mn is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$. Slater's rules require that we reorganize these electrons into groups, as follows: $(1s^2)$ $(2s^2 2p^6)$ $(3s^2 3p^6)$ $(3d^5)$ $(4s^2)$. For a $4s^2$ electron, the effective nuclear charge is estimated as

$$Z_{eff} = Z - S = 25 - (1)(0.35) - (13)(0.85) - (10)(1.00) = 25 - 21.4 = 3.6$$

and for a $3d^5$ electron, the effective nuclear charge is

$$Z_{eff} = Z - S = 25 - (4)(0.35) - (18)(1.00) = 25 - 19.4 = 5.6$$

Note that the all 13 of the $n = 3$ electrons contribute equally to the screening of the $4s^2$ electrons and that the $4s^2$ electron do not contribute to the screening of the $3d^5$ electrons.

From Coulomb's law we know that the greater Z_{eff}, the greater the ionization energy; thus, the larger Z_{eff} for the 3*d* electrons means that they are more stable and the 4*s* electrons are removed before the 3*d* electrons.