Chem 130 – First Exam Key

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.

Please write neatly!

Problem 1. For each of the following, provide <u>one</u> example of an element that fulfills the stated condition. If no element meets the condition, then write NONE. *Do not include lanthanides and actinides in your answers, and do not use any element more than once!* Possible answer shown below

(a) is an alkaline earth elements in Group 2

(b) has a core electron configuration of [Ar] elements in the fourth period

(c) has common ions with charges of –1 and of +1 H is the only choice

(d) has exactly three electrons in a *p*-orbital elements in Group 15

(e) forms a monoatomic ion with a charge of +1 elements in Groups 1 and 11, plus Ga, In, Tl

(f) has exactly one unpaired electron elements in Groups 1, 3, 11, 13, and 16

(g) is in the same period as chlorine elements in the third row

(h) has a valence electron with q.n. $n=3$, $l=2$, $m_l=0$, $m_s= -\frac{1}{2}$ 3d: period four (but not K and Ca)

(i) has a valence shell that consists of only *s* electrons elements in Groups 1 and 2

(j) is a halogen with a covalent radius smaller than that for chlorine F is the only choice

(k) forms ion with a charge of –1 that has a noble gas electron configuration elements in Group 17

(l) has exactly two core electrons elements in the second period

(m) is a metalloid (also called a semi-metal) B, Si, Ge, As, Sb, Bi, Po

(n) is in the *p*-block elements in Groups 13, 14, 15, 16, 17, and 18

(o) has a half-filled orbital elements in Groups 1, 7, and 15

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

Formula	Name	Covalent or Ionic?
NH_4NO_3	ammonium nitrate	ionic
$Al_2(SO_4)_3$	aluminum sulfate	ionic
S_4N_3	tetrasulfur trinitride	covalent

Problem 3. The figure to the right shows three possible transitions in which the sole electron in He$^+$ moves from a shell of higher n to a shell of lower n. Of these transitions, which will emit light with the longest wavelength? Without doing any calculations, give your answer as A, B, or C, and explain the reason for your choice in 1-2 sentences.
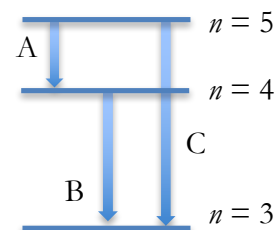
There is an inverse relationship between energy and wavelength, so longer wavelengths correspond to smaller energies; thus, A, which has the smallest difference between energy levels is the transition that has the shortest wavelength.

Given your answer above, what wavelength of light (in nm) is emitted. Note that a correct answer to this question does not depend on whether you correctly identified A, B, or C as the transition with the longest wavelength.

The relationship between the energy levels and the wavelength of light is

$$\frac{1}{\lambda} = 1.09737 \times 10^{-2}\ \text{nm} \times (+2)^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

Note that Z has a value of +2 for helium. For A, n_1 is 4 and n_2 is 5, giving a wavelength of 1013 nm. For B, n_1 is 3 and n_2 is 4, giving a wavelength of 469 nm. For C, n_1 is 3 and n_2 is 5, giving a wavelength of 320 nm.

Problem 4. The figure to the right shows the photoelectron spectrum for element X; note that the spectrum has six peaks, with the two peaks at the far right being very close in energy. (a) What is the electron configuration for X?

We read the spectrum from left-to-right in terms of orbitals, using the height along the y-axis to give the number of electrons; thus, $1s^2\,2s^2\,2p^6\,3s^2\,3p^4$.
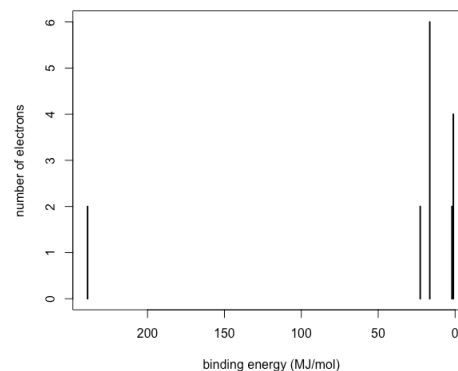
(b) Draw a circle around each peak in the photoelectron spectrum that corresponds to a core electron. In one sentence, explain your reason for choosing these peaks.

The core electrons are the 1s, 2s, and 2p orbitals, which have the greater ionization energies and are the first three peaks from left-to-right. Note that you need to explain that core electrons have greater ionization energies.

(c) The first ionization energy for element Y, which lies immediately above or immediately below X in the periodic table, is greater than the first ionization energy for X. Identify element Y by name and explain the reason for your choice, limiting your response to more than two sentences.

First, the electron configuration tells us that element X is S, which means that element Y is either O or Se. We know that ionization energies decrease down a column due to the increased distance between the valence electrons and the nucleus; thus, element Y is O. Note that immediately above (up) and immediately below (down) describes movement within a group (column); we use right and left to describe movement across a period (row).

Problem 5. Copper forms two common cations, one with a charge of +1 and one with a charge of +2. The elements on either side of copper, nickel and zinc, form just one common cation with a charge of +2. Write the electron configurations for each element and then explain why the differences in the charges of these ions make sense, limiting your response to no more than 2–3 sentences.

The electron configurations are $[Ar]4s^23d^8$ for Ni, $[Ar]4s^13d^{10}$ for Cu, and $[Ar]4s^23d^{10}$ for Zn where the unusual electron configuration for Cu reflects the stability of a full $3d$ subshell. When copper forms a +1 cation, it loses its single 4s electron giving an ion with both an empty $4s$ subshell and a filled $3d$ subshell; for Ni and for Zn, the +2 ion gives an empty $4s$ subshell and, for Zn, a full $3d$ subshell as well. There is a stability to ions when valence subshells are either empty or full.

Problem 6. The following three species have identical electron configurations: Ar, S^{2-}, and K^+. Which of these species has the largest ionization energy and which has the smallest ionization energy? Or, do all three have identical ionization energies? Explain the basis of your answer, limiting your response to 2–4 sentences.

The species with the largest ionization energy is K^+ and the species with the smallest ionization energy is S^{2-}.

As the three species are isoelectronic, they have identical electron configurations and the same distribution of core and valence electrons when we consider Z_{eff}. The differ in two significant ways: the magnitude of the charge on the nucleus, Z, and the charge. Both work to increase the ionization energy for K^+ and to decrease the ionization energy for S^{2-} relative to Ar; potassium has the largest Z and its positive charge makes it more difficult to remove additional electrons, and sulfur has the smallest Z and its negative charge makes it easier to remove an electron.

Problem 7. Consider the set of five ions listed below and rearrange them so that they are listed from the smallest radius to the largest radius.

$$Na^+ \quad Mg^{2+} \quad S^{2-} \quad Cl^- \quad Se^{2-}$$
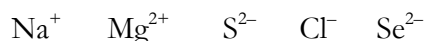
smallest radius ⟶ Mg^{2+} ⟶ Na^+ ⟶ Cl^- ⟶ S^{2-} ⟶ Se^{2-} ⟶ largest radius

Explain the basis of your ranking, limiting your response to 2–4 sentences.

For elements in the same period (all but Se), cations are smaller than anions because the loss of an electron in a cation means that the remaining electrons are pulled in closer to the nucleus, and the addition of an electron in an anion pushes electrons away from the nucleus due to electron-electron repulsion. Within a period, cations and anions decrease in size across the row due to the increase in the charge on the nucleus, which pulls electrons closer to the nucleus. The largest radius is for Se^{2-}, which is in the fourth period instead of the third period.

Problem 8. The electron configuration for arsenic often is written as $[Ar]4s^23d^{10}4p^3$ to indicate the order in which the element's orbitals fill. This configuration seems to suggest that the $3d$ electrons are in the valence shell; however, the $3d$ electrons generally are considered to be core electrons, leaving the $4s$ and the $4p$ electrons as valence electrons. Use Slater's rules to estimate the effective nuclear charge for one of arsenic's $3d$ electrons and for one of its $4s$ or $4p$ electrons. Next, using these values, explain why it is reasonable to consider arsenic's $3d$ electrons as core electrons. Limit your explanation to 2–3 sentences.

The grouping of electrons is $(1s^2)$ $(2s^22p^6)$ $(3s^23p^6)$ $(3d^{10})$ $(4s^24p^3)$. For a $3d$ electrons, Slater's rules give

$$Z_{eff} = 33 - (9)(0.35) - (18)(1.00) = 11.85$$

and for a $4s4p$ electron, Slater's rules give

$$Z_{eff} = 33 - (4)(0.35) - (18)(0.85) - (10)(1.00) = 6.3$$

Core electrons have significantly greater ionization energies than do valence electrons. From Coulomb's law, we know that a larger Z_{eff} equates to a greater ionization energy; thus, the much larger Z_{eff} for a $3d$ electron relative to a $4s4p$ electron provides strong evidence that the 3d electrons are core electrons.