Chem 130 – First Exam Key

Name________________________________

On the following pages you will find questions covering 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 _____/15   Question 4 _____/16

Question 2 _____/10   Question 5 _____/15

Question 3 _____/14   Question 6 _____/15

Question 7 _____/15

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!*

(a) forms a monatomic ion with a charge of –2
any of O, S, Se, Te, Po

(b) is a halogen
any of F, Cl, Br, I, At

(c) is in the third period
any of Na, Mg, Al, Si, P, S, Cl, Ar

(d) has a valence shell electron configuration of ns^2nd^3
none; if ns^2, then $(n\text{–}1)d^3$ is possible, but nd^3 is not possible

(e) forms a +2 ion with a noble gas electron configuration
any of Be, Mg, Ca, Sr, Ba, Ra

(f) is in the *d*-block
any element from Sc → Zn, Y → Cd, La, Hf → Hg, Ac, Rf → Cn

(g) forms common monoatomic ions with charges of +1 and +2
just Cu

(h) has exactly two electrons in a *d*-orbital
any of Ti, Zr, Hf, Rf

(i) has no unpaired electrons
any ns^2 (He, Be → Ra), ns^2np^6 (Ne → Rn), $ns^2(n\text{–}1)d^{10}$ (Zn → Hg)

(j) has a valence electron with the quantum numbers $n = 3$, $l = 1$, $m_l = 0$, and $m_s = +½$
any of Al → Ar

(k) is an alkali metal with a covalent radius larger than that for potassium
any of Rb, Cs, Fr

(l) has exactly 10 core electrons
any of Na → Ar

(m) is a metalloid
any of B, Si, Ge, As, Sb, Te

(n) has a first ionization energy greater than that for fluorine
either Ne or He

(o) is deflected by a magnetic field
any element except ns^2 (He, Be → Ra), ns^2np^6 (Ne → Rn), $ns^2(n\text{–}1)d^{10}$ (Zn → Hg)

Problem 2. Fill in the missing information for these three compounds, which you have seen in lab.

Formula	Name	Covalent or Ionic?
$Fe(OH)_3$	iron(III) hydroxide	ionic
NO_2	nitrogen dioxide	covalent
KSCN	potassium thiocyanate	ionic

Problem 3. The energy needed to remove an electron from a single He atom is 3.94×10^{-18} J. To what wavelength of light, in nanometers, does this correspond?

We know that $E = hc/\lambda$ or, upon rearranging, that $\lambda = hc/E$. Substituting in and solving gives

$$\lambda = \frac{(6.626 \times 10^{-34} \text{Js})(2.998 \times 10^{8} \text{m/s})}{(3.94 \times 10^{-18} \text{J})} = 5.04 \times 10^{-8} \text{m}$$

which is equivalent to a wavelength of 50.4 nm.

The Li^+ cation and the He atom have identical electron configurations of $1s^2$. To remove an electron from Li^+, will you need to use light of a longer wavelength, the same wavelength, or a shorter wavelength than that for He? Explain your reasoning in 1–3 sentences.

The ionization energy for Li^+ is greater than that for He for two reasons: Li^+ has a greater nuclear charge and its positive charge draws in the electrons, decreasing the distance between the nucleus and the electrons. From Coulomb's law, both make the ionization energy larger for Li^+. A larger ionization energy corresponds to a shorter wavelength because they are inversely proportional.

Problem 4. The elements Ar, Mg, and K have, in no particular order, first ionization energies, covalent radii, and average valence electron energies of

first ionization energies: 419 kJ/mol 738 kJ/mol 1521 kJ/mol

covalent radii: 0.130 nm 0.097 nm 0.196 nm

average valence electron energies: 420 kJ/mol 1845 kJ/mol 740 kJ/mol

Using the table below, match each element to its first ionization energy, to its covalent radius, and to its average valence electron energy. In the space below the table, and using no more than 4–8 sentences, define each of these properties of an atom and justify your assignments. *Please note the caution on the first page regarding written explanations.*

element	first ionization energy	covalent radius	average valence electron energy
Ar	1521	0.097	1845
Mg	738	0.130	740
K	419	0.196	420

The first ionization energy is the energy to remove one of element's valence electrons. From Coulomb's law we know that a greater n, and thus a greater distance, and/or a greater Z_{eff}, and thus a greater Q_+, increases ionization energy; thus, we expect K ($4s^1$) has a smaller ionization energy than Mg ($3s^1$) or Ar ($3s^2 3p^6$) because its valence electron has a greater n, and that Ar has a greater ionization energy than Mg because it has a greater nuclear charge and, therefore, a greater Z_{eff}. The covalent radius is one measure of an element's size. The radius increases as we move from shells with lower values of n to shells with larger values of n, and size decreases as Z_{eff} increases due to the increased attraction between the electrons and the nucleus; thus, we expect K ($4s^1$) has a larger covalent radius than Mg ($3s^1$) or Ar ($3s^2 3p^6$) because its valence electron has a greater n, and that Ar has a smaller covalent radius than Mg because it has a greater nuclear charge and, thus, greater Z_{eff}. Finally, the average valence electron energy is the average ionization energy for all of an element's valence electrons. Because the ionization energies for an element's valence electrons are similar in value, we expect the trend in AVEEs to mirror closely the trend in first ionization energies.

Problem 5. The first four ionization energies for an element are as follows:

IE_1 = 801 kJ/mol IE_2 = 2,427 kJ/mol IE_3 = 3,660 kJ/mol IE_4 = 25,026 kJ/mol

The element is in either the first or the second row of the periodic table. Identify the element and explain your reasoning. If you cannot narrow your choice to a single element, then explain which elements you can exclude (and why) and what additional information you need to identify the element. In either case, limit your response to 2–4 sentences. *Please note the caution on the first page regarding written explanations.*

First, we exclude the elements from the first row and Li because they have no more than three electrons and, therefore, cannot have four ionization energies. The large increase in ionization energy from IE_3 to IE_4 tells us that IE_4 corresponds to removing an electron from a shell closer to the nucleus than for IE_3; thus, we know the element has three electrons in the n = 2 shell and, of course, two electrons in the n = 1 shell. With an electron configuration of $1s^2 2s^2 2p^1$, we know the element is boron.

Problem 6. Rank the elements in the second row of the periodic table (Li, Be, B, C, N, O, F, Ne) from the smallest first ionization energy to the largest first ionization energy. Place each element in a space below and, in 2–4 sentences, explain how you arrived at this order. *Please note the caution on the first page regarding written explanations.*

Li < B < Be < C < O < N < F < Ne

Although the general trend in ionization energies is to increase across a row due to the increase in Z_{eff}, there are two important reversals. The first reversal is boron and beryllium where the shielding of boron's $2p$ electron by its $2s$ electrons results in boron's $2p$ electrons experiencing a smaller Z_{eff} than do beryllium's $2s$ electrons. The second reversal is between oxygen and nitrogen, which results from the repulsion between the paired electrons in oxygen's $2p$ orbital ($2p^4$) versus the unpaired electrons in nitrogen's $2p$ orbital ($2p^3$), which decreases IE even though Z_{eff} for oxygen is greater than Z_{eff} for nitrogen.

Problem 7. We have considered two models for determining the effective nuclear charge seen by an electron: a simple model based on Bohr's model of the atom and Slater's rules, which are based on a more complete model of the atom. Using chlorine as an example, explain why these two models yield different results for Z_{eff}, illustrating your answers with appropriate calculations for Z_{eff} using each model. Limit your response to 3–6 sentences.

In the simple Bohr model of the atom we assume that core electrons fully screen valence electrons and that valence electrons do not screen each other. In the equation $Z_{eff} = Z - S$, the value of S is equal to the number of core electrons only; thus, Z_{eff} for chlorine ($1s^2 2s^2 2p^6 3s^2 3p^5$) is 17 – 10 = 7. Slater's rules use are based on our more complete model of the atom, where an electron's radial distribution function shows that a valence electron can be closer to the nucleus that a core electron and, therefore, is not shielded fully by the core electrons. Using chlorine as an example, the calculation for Z_{eff} assumes that the valence electrons in the n = 3 shell partially shield each other by 0.35, that the core electrons in the n = 2 shell partially shield the valence electrons by an amount of 0.85, and that only the core electrons in the n = 1 shell fully shield the valence electrons; thus, for chlorine we have Z_{eff} = 17 – (2)(1.00) – (8)(0.85) – (6)(0.35) = 6.1.