

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 ____/24

Question 2 ____/6

Question 6 ____/13

Question 3 ____/16

Question 7 ____/13

Question 4 ____/12

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 four rows of the periodic table (H through Kr). *Do not use any element more than once!*

- (a) is a noble gas: He, Ne, Ar, Kr
- (b) has a core electron configuration of [Ar]: any element in the fourth row
- (c) forms a monoatomic ion with a charge of +3: Al, Cr, Fe, Sc, Ti, Mn, V, Co, Ge
- (d) has exactly two electrons in a *p*-orbital: C, Si, Ge
- (e) forms a monoatomic ion with a charge of -2: O, S, Se
- (f) has exactly three unpaired electrons: N, P, As, V, Co
- (g) is in the same period as sulfur: Na, Mg, Al, Si, P, Cl, Ar
- (h) has six peaks in its photoelectron spectroscopy spectrum: K, Ca
- (i) has a valence shell that consists of only *s* and *p* electrons: B to Ne, Al to Ar, Ga to Kr
- (j) is an alkaline metal with a covalent radius smaller than that for potassium: Na, Li
- (k) forms +1 ion with noble gas electron configuration: Li, Na, K
- (l) has six core electrons: NONE
- (m) is an alkaline earth: Be, Mg, Ca
- (n) is in the *p*-block: B to Ne, Al to Ar, Ga to Kr
- (o) has a *Z* of 24: Cr

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

Formula	Name	Covalent or Ionic?
Al ₂ (SO ₄) ₃	aluminum sulfate	ionic
Cl ₂ O ₇	dichlorine heptoxide	covalent
Co(OH) ₂	cobalt (II) hydroxide	ionic

Problem 3. Consider two photons, one with a wavelength of 200 nm and one with a frequency of $8.0 \times 10^{14} \text{ s}^{-1}$. Which photon has the smaller energy and what is that energy in Joules? Explain your reasoning with both a calculation and a brief written response of no more than two sentences.

$$E = hc/\lambda = (6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s}) / (200 \times 10^{-9} \text{ m}) = 9.93 \times 10^{-19} \text{ J}$$

$$E = h\nu = (6.626 \times 10^{-34} \text{ Js})(8.0 \times 10^{14} \text{ s}^{-1}) = 5.30 \times 10^{-19} \text{ J}$$

The smaller of the two energies is $5.30 \times 10^{-19} \text{ J}$, or the photon with a frequency of $8.0 \times 10^{14} \text{ s}^{-1}$.

Is either photon capable of ejecting a 4s electron from calcium, which has an ionization energy of 590 kJ/mol? Explain your reasoning with both a calculation and a brief written response of no more than two sentences.

Only need to test the photon with more energy. First convert to kJ/mol

$$E = 9.93 \times 10^{-19} \text{ J} \times (1 \text{ kJ}/1000 \text{ J}) \times 6.022 \times 10^{23} \text{ mol}^{-1} = 598 \text{ kJ/mol}$$

This is greater than 590 kJ/mole; thus, yes, one of the photons has sufficient energy.

Problem 4. The figure on the right shows two views of the photoelectron spectrum for the element X. The view on the left shows the full spectrum and the figure on the right shows only those peaks with energies less than 50 MJ/mol.

What is the electron configuration for X?



Draw a circle around the peak or peaks in the photoelectron spectrum that corresponds to X's valence electron(s). In one sentence, explain the reason for your choice.

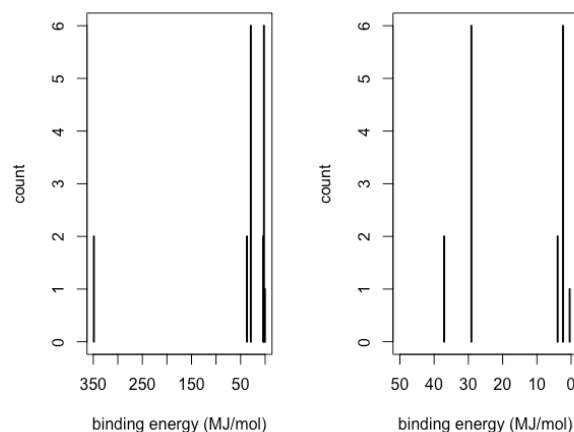
The valence shell is the 4s orbital, which is the rightmost peak in the photoelectron spectrum.

For the peak with a binding energy of approximately 30 MJ/mol, what are the values for the quantum numbers n , l , and m_l .

This is a 2p electron; thus $n = 2$, $l = 1$, and $m_l = -1, 0, \text{ or } +1$. Note that we list all three values for m_l because all are possibilities.

Suppose you pass a beam of gas-phase atoms of X through the poles of a magnet. Will the beam of atoms pass through the magnet without being deflected? Explain your reasoning in 1-2 sentences.

No. The beam of atoms will be deflected because the single electron in the 4s orbital is unpaired.



Problem 5. The table below shows several properties for the elements lithium, beryllium, and boron, where IE_1 , IE_2 , and IE_3 are the first, second, and third ionization energies, AVEE is the average valence electron energy, and EA is the electron affinity.

property	Li	Be	B
IE_1 (kJ/mol)	520	900	800
IE_2 (kJ/mol)	7300	1760	2430
IE_3 (kJ/mol)	11800	14850	3660
AVEE (kJ/mol)	520	900	1170
EA (kJ/mol)	60	0	27

Use the values in this table to answer the following, limiting yourself to 2-4 sentence each.

Explain the trend in the first ionization energies for these elements.

Ionization energies generally increase across a period because of the increase in effective nuclear charge, which, through Coulomb's law is directly proportional to IE. This is sufficient to explain the first ionization energy for Be relative to Li, but not the first ionization energy for B. For the latter, the valence shell includes both $2s$ and $2p$ electrons (the others have just $2s$ electrons). Because s electrons screen p electrons better than p electrons screen s electrons, the effective nuclear charge for the valence electron in B is smaller than expected, decreasing its IE.

Compared to lithium and beryllium, the third ionization energy for boron is rather small. Why?

The third ionization energy for Li and for Be removes a $1s$ electron; however, for B, the third ionization energy is for a $2s$ electron. For B, the third ionization energy is from its valence shell instead of its core electrons, as is the case for Li and for Be; thus, the much smaller ionization energy.

Interestingly, the AVEE for Li and for Be are identical to their first ionization energies. This is not the case for boron. Explain why.

The AVEE is the average of an element's first ionization energies for its valence electrons. For Li and for Be, the valence shell consists only of $2s$ electrons; thus, their IE and AVEE values are the same. Boron has both $2s$ and $2p$ valence electrons, so its AVEE value is an average of two different IEs and, therefore, different from the first ionization energy.

Explain why lithium has a more positive electron affinity than does boron and why beryllium has an electron affinity of zero.

There is a particular stability when an element has filled orbitals in its valence shell. As beryllium has an electron configuration in which the $2s$ orbitals is filled, it has a negligible electron affinity. Lithium has a single electron in the $2s$ orbital and boron has a filled $2s$ orbital and one electron in its $2p$ orbital. Lithium fills its $2s$ orbital by adding a single electron, but boron does not; thus, the difference in their electron affinities.

Problem 6. Consider the set of ions listed below, each of which has a noble gas configuration, and rearrange them so that they are listed from the smallest radius to the largest radius.



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

Cations are smaller than their neutral atoms because the effective nuclear charge is spread over fewer valence electrons, pulling each of them in closer to the nucleus. Anions are larger than their neutral atoms because the effective nuclear charge is spread over more valence electrons, which allows their natural repulsion from each other to increase the radius. In general, the greater the charge, the more pronounced the effect. This easily gives the relative rankings of Mg^{2+} , Na^+ , Cl^- , and S^{2-} , which all are in the second row. We can reasonably expect that Se^{2-} is bigger than S^{2-} as it is in the third row. Ranking Mg^{2+} and Li^+ is challenging because we have to consider both the difference in charge and the difference in row; either order is acceptable given what we know, although the correct order is given above.

Problem 7. The electron configuration for arsenic often is written as $[\text{Ar}]4s^23d^104p^3$ 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 groupings from Slater's rules are $(1s^2) (2s^2 2p^6) (3s^2 3p^6) (3d^{10}) (4s^2 4p^3)$. For a $3d$ electron we have

$$Z_{\text{eff}} = Z - S = 33 - (18)(1) - (9)(0.35) = 11.85$$

and for a $4s$ or $4p$ electron we have

$$Z_{\text{eff}} = Z - S = 33 - (10)(1) - (18)(0.85) - (4)(0.35) = 6.3$$

Because Z_{eff} is so much greater for a $3d$ electron than for a $4s$ or $4p$ electron, its ionization energy will be significantly greater as well. As core electrons are much harder to remove than valence electrons, this is consistent with viewing the $3d$ electrons as being core electrons.