Checklist for Exam 3

Chapter 5. Quantum Chemistry and Periodicity

General ideas.

- □ I know how energy, wavelength and frequency are related to one another (e.g. as wavelength increases, the frequency decreases and so does the energy.) I can do this qualitatively and quantitatively using equations that will be provided. I can use the equations and know to be careful with the units!
- □ I know the general regions of the electromagnetic spectrum and where the visible region fits into the scheme of things. I know that violet corresponds to 380 nm and red to 780 nm
- I know what the lines in the hydrogen discharge spectrum represent. I can use the Balmer-Rydberg equation (provided) to calculate wavelength or frequency given two quantum numbers, n_f and n_j.
- \Box I know that removing an electron corresponds to taking the electron to n_f = infinity.
- □ When considering an electron moving from one energy level to another, I know which transitions correspond to absorptions/gaining and which correspond to releases/discharges of energy.
- □ I know that the wavelength and frequency refer to single photon/electron events and that if I use these values to calculate energy, the energy I get (typically ~10⁻¹⁹ J) refers to a single photon/electron event. If I want the answer in kJ/mol, I must multiply by Avogadro's number (and convert J to kJ, too.)
- □ Conversely, if I am given energy in kJ/mol, and I want wavelength or frequency, I must first divide by Avogadro's number and then convert to joules.

Quantum numbers.

- □ I am conversant in the four quantum numbers, *n*, *l*, m_l , and m_s . I know that the *n* quantum number is the energy quantum number and that *n* determines the allowed values for *l*
- □ I can relate the *n* quantum number to the rows on the periodic table.
- □ I know that the *I* quantum number is the orbital shape quantum number and that *I* determines the allowed values for m_{I} .
- I know that when / = 0, this is called the s-orbital and when / = 1, it is the p-orbitals and so on. The sequence for / is 0 (s-orbital), 1 (p-orbitals), 2 (d-orbitals), 3 (f-orbitals), 4 (g-orbitals), 5 (h-orbitals), 6 (i-orbitals), and so on through the alphabet.
- □ I can relate the *I* quantum number to the columns on the periodic table.
- □ I know how the notation "1s" "3p" etc. relate to the quantum numbers *n* and *l*.
- □ I can use quantum numbers to identify an orbital and I can assign quantum numbers to an orbital.
- □ I can write the general valence-shell electron configuration for each group of the periodic table, and identify the blocks in which the elements are located.
- □ I know that the m_l quantum number refers to specific orbitals within a value for *l*. Specifically, m_l ranges from -l to +l. so if l = 3, m_l ranges from -3 to +3 (7 values)
- □ I can sketch and name each of the *s*, *p*, and *d* orbitals.
- □ I know that orbitals are designated by three quantum numbers, *n*, *l*, and m_l . Electrons are designated by these same three quantum numbers plus one more, m_s .
- □ I know the Pauli Exclusion Principle, Hund's Rule and the Aufbau Principle.

Chapter 6. Ionic Bonding and Periodic Trends

Periodic trends.

- □ I know the periodic trend across the periodic table for effective nuclear charge, Z_{eff}.
- □ I can write the ground-state electron configurations for any element.
- □ I can use electron configurations or orbital-filling diagrams to determine the number of unpaired electrons in these species.
- □ I know the periodic trend both across and down for atomic radius (size). I know that $Z_{\text{eff.}}$ explains the trend in size across the periodic table and the vertical trend is explained by the fact that increasingly larger shells (bigger n values) are being filled as one descends the periodic table.
- □ I can predict the ground-state electron configuration for ions.
- □ I know that cations are smaller than their neutral counterparts due to the increase in Z_{eff} . Furthermore, +2 cations would be expected to be smaller than +1 cations, for the same reason.
- □ In a similar way, I know that anions are larger than their neutral atoms. Again, the lowered Z_{eff} for anions accounts for this fact.
- □ I know the definition for first ionization energy: $E \rightarrow E^+ + e^-$. I know the periodic trend for first ionization energy and that ionization energies are always positive no element wants to lose an electron. I know that the left to right trend is explained largely by increasing Z_{eff} . (Notable "modifications" to the rule occur for electron configurations that possess one electron in an orbital (e.g. $2s^1$ or $4p^1$) or one more than half-filled configurations (e.g. $2p^4$ or $4d^6$). I know that the first ionization energy decreases as one goes down a column because the outermost electron is in increasingly bigger n shells and is easier to remove out to n = infinity (the definition of ionization.)
- □ I know the definitions for second and third ionization energies $(2^{nd} \text{ ionization energy: } E^+ \rightarrow E^{+2} + e^- \text{ and } 3^{rd} \text{ ionization energy: } E^{+2} \rightarrow E^{+3} + e^-)$. I know that subsequent ionization energies increase (e.g., the 2^{nd} ionization energy is always takes more energy than the first.)
- I know that subsequent ionization energies increase, but then jump substantially when an electron is being removed from a filled shell. For example, magnesium can lose two electrons, but losing the third is very difficult.
- □ I know the definition of electron affinity: E + e⁻ → E⁻. I know left-to-right periodic trend for electron affinity and that it is largely explained by Z_{eff}. It is important to know that filled shells have zero or close to zero electron affinities because they have filled shells and it takes energy to promote an electron to a higher shell.
- □ I know the definition of lattice energy: $m A^{-n}(g) + n B^{-m}(g) \rightarrow A_m B_n(s)$. I know lattice energy is always exothermic and it increases with the charges on the cation and anion. I know that one tangible result of large lattice energy is a decrease in solubility.