Chem E-1a Friday Review Answers Chapter 8: The Periodic Table

- 1. a) $[Ar]4s^23d^{10}4p^6$
 - b) $[Xe]4f^{14}5d^8$
 - c) [Xe] $4f^{14}5d^{10}$
 - d) $[Ar]3d^2$

2.

a) Look at the electron configurations:

| 1 st IE of Al: | Al [Ne] $3s^2 3p^1$ | \rightarrow | Al ⁺ [Ne] $3s^2$ |
|---------------------------|---------------------|---------------|-----------------------------|
| 1 st IE of Mg: | Mg [Ne] $3s^2$ | \rightarrow | Mg^+ [Ne] $3s^1$ |

 1^{st} IE of Al < 1^{st} IE of Mg : This is an exception to the periodic trend, because the 3p orbital in Al is higher in energy than the 3s orbital in Mg, so it costs more energy to remove the outermost electron from Mg.

| 2 nd IE of Al: | Al ⁺ [Ne] $3s^2$ | \rightarrow | Al^{2+} [Ne] $3s^{1}$ |
|---------------------------|-----------------------------|---------------|-----------------------------------|
| 2 nd IE of Mg: | Mg ⁺ [Ne] $3s^1$ | \rightarrow | Mg^{2+} [Ne] = [He] $2s^2 2p^6$ |

 2^{nd} IE of Al > 2^{nd} IE of Mg: This follows the periodic trend, and is explained because the *n* of both is the same, but the Z_{eff} of Al (or in this case Al⁺) is greater, so the ionization energy is greater.

 $3^{rd} \text{ IE of Al:} \quad \text{Al}^{2+} [\text{Ne}] \ 3s^1 \qquad \rightarrow \qquad \text{Al}^{3+} [\text{Ne}] = [\text{He}] \ 2s^2 2p^6$ $3^{rd} \text{ IE of Mg:} \quad \text{Mg}^{2+} [\text{He}] \ 2s^2 2p^6 \qquad \rightarrow \qquad \text{Mg}^{3+} [\text{He}] \ 2s^2 2p^5$

 3^{rd} IE of Al < 3^{rd} IE of Mg: This is an exception to the trend. It is much harder to remove the outermost electron from Mg²⁺ because it has a stable noble gas electron configuration, and the outermost electrons are in a lower-energy 2p orbital.

b) EA of N: N [He] $2s^2 2p^3 \rightarrow N^-$ [He] $2s^2 2p^4$ EA of C: C [He] $2s^2 2p^2 \rightarrow C^-$ [He] $2s^2 2p^3$

Adding an electron to carbon makes a half-filled 2p subshell. This is a favorable process, so the electron affinity will be very negative (exothermic). Nitrogen already has a half-filled 2p subshell, so adding an electron requires pairing in an orbital with an electron already and disrupting the half-filled subshell. Spin-pairing energy makes this process unfavorable.

c) The first electron affinity of oxygen is negative, like the first electron affinities of most elements, because it is a favorable process to add an electron to a neutral oxygen atom. When you add a second electron to a -1 oxygen ion, however, you are adding a negatively charged particle (an electron) to an already negatively charged species (a -1 ion) so the charge repulsion causes this to be an unfavorable process, and the electron affinity (the 2nd EA of oxygen) is thus positive. (Note that it does not matter that this 2nd EA forms a full p subshell and noble-gas configuration. The charge repulsion is enough to make this an overall unfavorable process. Because of these charge repulsions, all 2nd electron affinities are positive. Similarly, all 2nd ionization energies are greater than 1st ionization energies for the same element.)

3.

- a) The radius of Rb is greater than the radius of Sr. The outermost electron in each has the same value of n, the principal quantum number. However, Sr has a greater nuclear charge, but the same number of core electrons as Rb, so the outermost electron in Sr experiences a greater Z_{eff} . Thus Sr is smaller than Rb.
- b) Rb is larger than Rb⁺. Positive ions are always smaller than their corresponding neutral atoms, because the removal of an electron results in a decrease in size. Fewer electrons held by the same nuclear charge results in a smaller raidus.
- c) Rb⁺ is smaller than Kr because they are isoelectronic (i.e. they have the same electron configurations) but Rb⁺ has a greater nuclear charge, so it is smaller.
- 4. a) F. F has the lowest value of n and the greatest Z_{eff} of all atoms in this group, so it has the smallest radius.
 - b) Mg²⁺. All of these species are isolelectronic (they all have the same electron configuration as Ne.) Thus the one with the greatest nuclear charge, which is Mg²⁺, will be the smallest.
- 5) a) Mg: The outer electron experiences a greater effective nuclear charge in Mg than in Na, so Mg is a smaller atom, the outermost electrons are held stronger to the nucleus, and the ionization energy of Mg is thus larger. This follows the standard periodic trend. (In addition, it also has a full s-subshell, so it's extra "stable" the way it is.)
 - b) P: an exception! P has a half-filled p subshell, making it particularly stable. S will have a half-filled subshell after removal of the outermost electron. Thus the first ionization energy of P is greater. Note that this is an exception to the normal periodic trend.