Periodic Trend Practice Questions key

1. Arrange each of the following sets of atoms or ions in order of increasing radius (i.e. from smallest to largest):

a) Ba, Ca, Cs, Sc

Atomic radius increases as you move from right to left and from top to bottom on the periodic table. Therefore, we expect the following order:

Sc < Ca < Ba < Cs

The reasons for these trends are complex, but they basically boil down to two factors: adding protons makes an atom smaller by decreasing the size of the orbitals (so atoms get smaller as you go across a row), and increasing the n value for the outermost orbitals makes an atom much larger (so atoms get larger as you go down a column).

b) Br–, Rb+, Se2–, Sr2+

These ions are isoelectronic – they all have the same electron configuration ([Kr]). For isoelectronic ions, the size decreases as the number of protons increases, because additional protons pull the electrons inward, making all of the orbitals smaller. Therefore, we expect the following order:

Sr2+< Rb+ < Br–< Se2

c) Zr2+, Zr3+, Zr4+

These ions have the same nucleus (because they’re the same element!), but they have different numbers of electrons. As you remove electrons from an atom, the atom gets smaller, because the remaining electrons do not repel each other as strongly. Of these three ions, Zr2+ has the most electrons (38) and Zr4+ has the fewest (36), so we expect the following order:

Zr4+ < Zr3+< Zr2+

1. An element in period 4 (elements 19 through 36) has the following ionization energies. Identify the element and explain logic.

IE1 = 633 kJ/mol IE2 = 1235 kJ/mol IE3 = 2389 kJ/mol

IE4 = 7090 kJ/mol IE5 = 8843 kJ/mol IE6 = 10679 kJ/mol

The first three ionization energies are fairly small, but starting with the fourth, we see a big jump in the amount of energy required to remove each electron. This tells us that the first three electrons are outer-shell electrons, which are fairly easy to remove, but the subsequent electrons are coming from the inert-gas core. Since three electrons are easy to remove, the element must have exactly three electrons outside of its inert-gas core. The only element that fits this description in period 4 is **scandium (Sc)**, which has the configuration [Ar]4s23d1.

1. Both vanadium and its 3+ ion are paramagnetic. Use electron configurations to explain why this is so.

V [Ar] 4s2 3d3 $\frac{\uparrow \downright }{4s} \frac{\uparrow }{} \frac{ \uparrow }{} \frac{ \uparrow }{3d} \frac{ }{} \frac{ }{} $

Original atom is paramagnetic because there are unpaired d electrons. When a +3 ion is formed the ion is still paramagnetic because the outermost s electrons are lost before the inner d electrons so there are still unpaired electrons.

V+3 [Ar] 3d2 $\frac{ }{4s} \frac{\uparrow }{} \frac{ \uparrow }{} \frac{ }{3d} \frac{ }{} \frac{ }{}$

1. Explain how effective nuclear charge and ionization energy are related.

The larger the effective nuclear charge, the more tightly the outer electrons are held and the higher the ionization energy

1. Why is the first ionization energy of oxygen less that the ionization energy for nitrogen

Nitrogen -Electron configuration [He]2s22p3 has three electrons in individual 2p orbitals with aligned spins because electrons try and stay as far apart as possible. Oxygen -Electron Configuration - [He]2s22p4 must put 2 electrons in a 2p orbital with opposite spins. This causes a slight repulsion between these electrons which is greater than electrons in separate orbitals making this electron slightly easier to remove.

In nitrogen, there are 3 electrons in outer shell i.e., it is half filled orbital and it required more energy to remove one electron because it is in stable state. But it in oxygen an electron can be easily removed.

1. Use the concepts of effective nuclear charge, shielding, and *n* value of the valence orbital to explain the trend in atomic radius as you move across a period in the periodic table from left to right.

As you move across a row in the periodic table, the n level stays the same. However, the nuclear charge increases and the amount of shielding stays about the same since the number of inner electrons stays about the same. So, the effective nuclear charge experienced by the electrons in the outermost principal energy level increases, resulting in a stronger attraction between the outermost electrons and the nucleus and therefore, a smaller atomic radii.

1. Each of the following energy terms is equal to ΔH for a particular chemical process. Write a specific, ***balanced chemical equation*** for each process.
2. the electron affinity of nitrogen

N(g) + e− →N− (g)

1. the first ionization energy of aluminum

Al(*g*) Al+(*g*) + e-