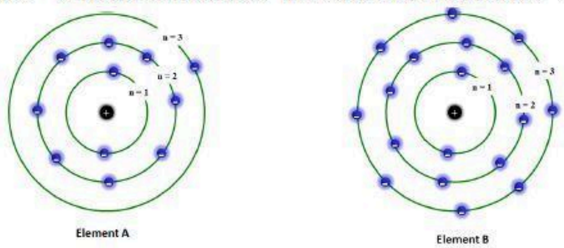
Atomic Structure and Properties
1.7 Periodic Trends
Worksheet Key

1) Which atom in each set has the larger radius?

    a. **Be** or O
    b. **Cu** or Br
    c. F or **I**
    d. O or **As**
    e. Kr or **K**
    f. **Ba** or Li

2) The shell models of two neutral atoms are drawn below. The representations are not drawn to scale. Which atom has the smallest radius? Justify your answer.



Element B has the smaller radius.

The valence electrons of both elements are in the same shell (n = 3). Thus, the shielding effect experienced by the outer electrons of both elements will be similar. Element A has eleven protons in its nucleus, as it is neutral and has eleven electrons. Element B has eighteen protons in its nucleus, as it is neutral and has eighteen electrons. The addition of protons increases the effective nuclear charge on the valence electrons.

$$Z_{\text{eff}} = Z - \sigma$$

According to Coulomb's Law the increase in positive charge increases the force of attraction experienced by each electron, thereby decreasing the radius. This gives element B a smaller atomic radius.

$$F = k\frac{q_1 q_2}{d^2}$$

3) Explain each of the following occurrences by referring to the structure of the atoms in question (energy levels, orbitals, protons, etc.).
    a. The atomic radius of oxygen is smaller than the atomic radius of carbon.

The valence electrons of both elements are in the same shell (n = 2). Thus, the shielding effect experienced by the outer electrons of both elements will be similar. The oxygen atom has eight protons in its nucleus while the carbon atom only has six protons. The addition of protons increases the effective nuclear charge on the valence electrons.

$$Z_{\text{eff}} = Z - \sigma$$

Moving from left to right, more protons are added. According to Coulomb's Law the increase in positive charge increases the force of attraction experienced by each electron, thereby decreasing the radius. This gives oxygen a smaller atomic radius.

$$F = k\frac{q_1 q_2}{d^2}$$

b. The atomic radius of Mg is smaller than the atomic radius of Ca.

The valence electrons in Mg are contained within the 3s orbital and the valence electrons in Ca are contained within the 4s orbital. This means that the outermost electrons in Ca experience a greater shielding effect, possess more energy and are further from the nucleus than the outermost electrons in Mg. According to Coulomb's Law the forces of attraction decrease as the distance between the electrons and the protons increases. Therefore, Mg has a smaller atomic radius.

4) Mg^{2+} and F^- are isoelectronic.

a. Which ion has the smaller radius?

Mg^{2+}

b. Explain why the radii of these two ions are different sizes. Justify your claims.

Both species have the same electron configurations, and thus, the same number of electrons. Mg^{2+} has 12 protons, while F^- only has 9 protons. Because Mg^{2+} has more protons, its nucleus applies a greater force of attraction on its electrons, thereby pulling them closer to the nucleus. According to Coulomb's Law the forces of attraction acting on an electron increases as more protons are added to the nucleus.

$$F = k\frac{q_1 q_2}{d^2}$$

This gives Mg^{2+} a smaller radius.

5) Which species has a smaller radius: Ca or Ca^{2+}? Provide an explanation and justify your claims using your knowledge of atomic structure.

Ca^{2+} has the smaller radius. Both Ca and Ca^{2+} have the same number of protons, but Ca^{2+} has two less electrons. This causes the protons to pull on the electrons in Ca^{2+} with a greater force of attraction. Most importantly, Ca^{2+} has one less shell. Ca has electrons in the n = 4 shell, and Ca^{2+} only has electrons in the n = 3 shell (it lost its two electrons in the n = 4 shell). By losing the n = 4 shell, the radius of Ca^{2+} decreased dramatically. Also, as there are fewer core electrons in Ca^{2+} its valence electrons experience less of a shielding effect.

6) Which species has a smaller radius: F or F^-? Provide an explanation and justify your claims using your knowledge of atomic structure.

F has the smaller radius. Both F and F^- have the same number of protons, but F has one less electron. This causes the protons to pull on the electrons in F with a greater force of attraction. The addition of the extra electron in F^- causes greater electron – electron repulsion which increases the distance between the valence electrons and increases its radius.

7) Explain each of the following occurrences by referencing the structure of the atoms in question (energy levels, orbitals, protons, etc.).
   a. The first ionization energy of Mg is 738.1 kJ/mol, while the first ionization energy of Al is only 577.9 kJ/mol.

This is an exception to the rule. Normally ionization energy increases as we move from left to right across a period in the periodic table, because the shielding effect is similar and the effective nuclear charge increases. The reason for this exception has to do with the electron configurations of these elements. Mg has an electron configuration of $1s^2 2s^2 2p^6 3s^2$, and Al has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^1$. It requires less energy to pull the only electron from the p-orbital in Al, which is a higher energy subshell, than it does to pull an electron from the full s-orbital in Mg, which is a lower energy subshell. The lone electron in the p-orbital of Al is at a much higher energy than those in the full s-orbital of Mg and it is being "shielded" by the other electrons in lower energy subshells.

   b. The first ionization energy of Mg is greater than the first ionization energy of Ca.

Ca has a larger atomic radius than Mg, due to the fact that the valence electrons of Ca are in a 4s orbital and those of Mg are in a 3s orbital. As the radius of Ca is larger, its valence electrons are held with a smaller force of attraction, and thus, it requires less energy to remove them. According to Coulomb's Law the forces of attraction decrease as the distance between electrons and protons increases.

$$F = k\frac{q_1 q_2}{d^2}$$

   c. The first ionization energy of Na is less than the first ionization energy of Cl.

Na has a larger atomic radius than Cl. The valence electrons of both elements are in the same shell (n = 3), so the shielding effect experienced by the valence electrons of both elements will be similar. However, the valence electrons in Cl experience a greater effective nuclear charge, as it has more protons in its nucleus. According to Coulomb's Law the forces of attraction acting on an electron increases as more protons are added to the nucleus.

$$F = k\frac{q_1 q_2}{d^2}$$

Because the force of attraction on the valence electrons in Na is less than that in Cl, it requires less energy to remove an electron from Na.

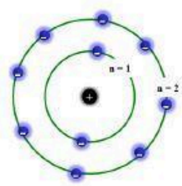
d. The second ionization energy of Na is 4560 kJ/mol, while the second ionization energy of Mg is only 1450 kJ/mol.

The second electron that is removed from Na is taken from the n = 2 shell, whereas the second electron that is removed from Mg is taken from the n = 3 shell. The distance between the nucleus and the n = 2 shell electrons is much less than the distance between the nucleus and the n = 3 electrons. The reduced distance means that the attractive forces between the protons and the n = 2 electrons is greater. According to Coulomb's Law the forces of attraction increase as the distance between electrons and protons decreases.

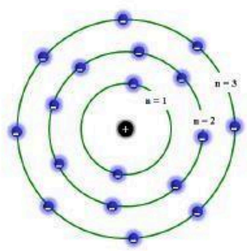
$$F = k\frac{q_1 q_2}{d^2}$$

The greater the attractive force, the more energy that is required to remove an electron.

8) The shell models of two neutral atoms are drawn below. The representations are not drawn to scale. Which atom would have the highest first ionization energy? Justify your answer.



Element A



Element B

Element A has the higher first ionization energy.

Element B has a larger atomic radius than element A, as it has an additional shell. As the radius of element B is larger, its valence electrons are held with a smaller force of attraction, and thus, it requires less energy to remove them. According to Coulomb's Law the forces of attraction increases as the distance between electrons and protons decreases.

$$F = k\frac{q_1 q_2}{d^2}$$

9) Why do the halogens have a negative electron affinity value, while the noble gases have a positive electron affinity value?

The highest energy subshell (*p*-orbital) is full for all of these elements. The added electron would be the only electron in the higher energy subshell. The shielding effect from the inner core electrons is so large that energy must be added in order for them to accept an additional electron. Adding energy is endothermic, which is a positive energy value.

The halogens need one more electron to fill their octets. For this reason they will readily accept an electron in order to fill their highest energy subshells (*p*-orbitals). In doing so, they give off energy. Giving of energy is an exothermic process, which has a negative energy value.

10) Why do the Group 1A elements have negative electron affinity values, while Group 2A elements have positive electron affinity values?

The highest energy subshell (*s*-orbital) is full for all of these elements. The added electron would be the only electron in the highest energy *p*-orbital. The shielding effect from the inner core electrons is so large that energy must be added in order for them to accept an additional electron. Adding energy is endothermic, which is a positive energy value.

Group 1A elements need one more electron to fill their s-orbitals. For this reason they will readily accept an electron. In doing so, they give off energy. Giving of energy is an exothermic process, which has a negative energy value.

11) Which element from each set is most electronegative?

a. **F** or C
b. Al or **Cl**
c. Po or **S**
d. Cs or **I**
e. Ca or **Cl**
f. **O** or Se
g. **Zn** or K
h. **C** or Pb
i. Ga or **O**

12) Sulfur is more electronegative than calcium. Explain why this is and justify your claims using your knowledge of atomic structure.

Electronegativity increases as atomic radius decreases. The valence electrons in larger elements experience a greater amount of shielding from inner core electrons than the valence electrons in smaller elements. When two atoms are involved in a chemical bond, the valence electrons will experience a greater attraction for the positive nucleus that is closest to them. According to Coulomb's Law, the force of attraction increases as the distance between electrons and protons decreases.

$$F = k\frac{q_1q_2}{d^2}$$