7.11 (a) What is meant by the term effective nuclear charge?

*Z* \text{eff} \text{ is the proton pull actually experienced by an electron in a given atom. Typically } Z \text{eff} < Z \text{ (the proton count).}

(b) How does the effective nuclear charge experienced by the valence electrons of an atom vary going from left to right across a period of the periodic table?

**Traveling from left-to-right in a period, } Z \text{ eff approaches } Z. \text{ That is, electrons experience more and more proton pull since no new "shielding" core electrons are being added, yet the # of protons IS increasing.}

7.23 How do the sizes of atoms change as we move
(a) from left to right across a row in the periodic table?

**DECREASE.** Since shielding is becoming less effective, the e- are pulled inward more tightly.

(b) from top to bottom in a group in the periodic table?

**INCREASE.** More shells of electrons.

(c) Arrange the following atoms in order of increasing atomic radius: F, P, S, As.

**F < S < P < As** (largest)

**LOGIC:** F has only two levels of e- and is at extreme right of its period (radius decreases left-to-right)

P > S since both in same period, but S is on right so has decreased more.

As > P since arsenic farther down same group as P.

7.25 Using only the periodic table, arrange each set of atoms in order of increasing radius:

(a) Ca, Mg, Be  
Be < Mg < Ca  
Radius increases down a group.

(b) Ga, Br, Ge  
Br < Ge < Ga  
Radius decreases left-to-right in period.

(c) Al, Tl, Si  
Si < Al < Tl  
Tl larger than Al (group tend) 
Al > Si (period trend)
7.27 (a) Why are monatomic cations smaller than their corresponding neutral atoms?

Cations are smaller than their parent atom because 1) often lose outer level of e-; 2) net positive charge pulls e- in tighter; and 3) less electron-electron repulsion.

(b) Why are monatomic anions larger than their corresponding neutral atoms?

Anions are larger than their parent atom because 1) net negative charge allows electrons to wander farther out; and 2) more electron-electron repulsion.

(c) Why does the size of ions increase as one proceeds down a column in the periodic table?
In comparing ions from the same group, as you travel down a group, more levels of electrons exist.

7.29 Consider a reaction represented by the following spheres:

Which sphere represents a metal and which a nonmetal? Explain.

Metals form cations, and since cations are smaller than their parent atom, the first (red) sphere will be the metal atom. Nonmetals form anions, and since anions are larger than their parent atom, the second (blue) sphere will be the nonmetal atom.

7.30 Consider the following spheres:

Which one represents Ca, which Ca\(^{2+}\), and which Mg\(^{2+}\)?

Largest is Ca atom. Intermediate size is Ca\(^{2+}\). Smallest is Mg\(^{2+}\).

Ca atom will be larger than Ca\(^{2+}\) because cations are smaller than their parent atom. The Ca\(^{2+}\) cation is larger than the Mg\(^{2+}\) cation since the calcium cation has one more complete level than the magnesium cation.
7.31 (a) What is an isoelectronic series?
A group of atoms or ions with the same electron count.

(b) Which neutral atom is isoelectronic with each of the following ions:

\[ \text{Al}^{3+} \quad \text{neon (Ne)} \]
\[ \text{Ti}^{4+} \quad \text{argon (Ar)} \]
\[ \text{Br}^{-} \quad \text{krypton (Kr)} \]
\[ \text{Sn}^{2+} \quad \text{cadmium (Cd)} \]

7.35 Consider S, Cl, and K and their most common ions.
(a) List the atoms in order of increasing size.

\[
\text{ATOM RADIUS} \quad \text{Cl} < \text{S} < \text{K}
\]
S is on left of Cl in same period, so is larger.
K has four shells, S only has three.

(b) List the ions in order of increasing size.

\[
\text{ION RADIUS} \quad \text{K}^+ < \text{Cl}^- < \text{S}^{2-}
\]
In this isoelectronic series, more protons means smaller radius.

7.37 For each of the following statements, provide an explanation:
(a) \( \text{O}^{2-} \) is larger than \( \text{O} \)
Anions larger than parent atom.

(b) \( \text{S}^{2-} \) is larger than \( \text{O}^{2-} \)
\( \text{S}^{2-} \) has 3 complete shells, \( \text{O}^{2-} \) has 2 complete shells.

(c) \( \text{S}^{2-} \) is larger than \( \text{K}^+ \)
Isoelectronic; \( \text{K}^+ \) has more protons to pull e- inward.

(d) \( \text{K}^+ \) is larger than \( \text{Ca}^{2+} \)
Isoelectronic; \( \text{Ca}^{2+} \) has more protons to pull e- inward.

7.41 (a) Why are ionization energies always positive quantities?
IE is the energy REQUIRED to remove an electron.

(b) Why does F have a larger first ionization energy than O?
The F atom is smaller than O, having a higher \( Z_{\text{eff}} \), and so its electrons are held more tightly.

(c) Why is the second ionization energy of an atom always greater than its first ionization energy?
After removing an electron, the \( Z_{\text{eff}} \) is larger and the remaining electrons are held more tightly. The net positive charge must be overcome.
7.43 (a) What is the general relationship between the size of an atom and its first ionization energy?

The smaller the atom, the larger the first IE.

7.45 Based on their positions in the periodic table, predict which atom of the following pairs will have the larger first ionization energy:

(a) Cl, Ar  
**Ar** (smaller atom, period trend)

(b) Be, Ca  
**Be** (smaller atom, group trend)

(c) K, Co  
**Co** (smaller atom, period trend)

(d) S, Ge  
**S** (smaller atom, S has less levels than 4th period Ge)

(e) Sn, Te  
**Te** (smaller atom, period trend)

7.47 Write the electron configurations for the following ions:

(a) **In**

\[ [\text{Kr}] 4d^{10} \text{ (lost 5p}^1 \text{ and 5s}^2 \text{) } \]

(b) **Sb**

\[ [\text{Kr}] 5s^24d^{10} \text{ (lost all 5p}^3 \text{) } \]

(c) **Te**

\[ [\text{Kr}] 5s^24d^{10}5p^6 \text{ or [Xe] gained 2 e-} \]

(d) **Te**

\[ [\text{Kr}] 4d^{10} \text{ (lost 5p}^4 \text{ and 5s}^2 \text{) } \]

(e) **Hg**

\[ [\text{Xe}] 4f^{14}5d^{10} \text{ (lost 6s}^2 \text{) } \]

(f) **Rh**

\[ [\text{Kr}] 4d^6 \text{ (lost the 4d}^7 \text{ e- and 5s}^2 \text{) } \]

7.49 Write the electron configuration for

(a) the **Ni**

\[ [\text{Ar}] 3d^8 \]

How many unpaired electrons does it contain?

\[ \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \quad \text{2 UNpaired e-} \]

3d 3d 3d 3d 3d

(b) the **Sn**

\[ [\text{Kr}] 5s^24d^{10} \]

How many unpaired electrons does it contain?

\[ \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \quad \text{0 UNpaired e-} \]

5s 4d 4d 4d 4d 4d

7.55 Consider the first ionization energy of neon and the electron affinity of fluorine.

(a) Write equations, including electron configurations, for each process. (b) These two quantities will have opposite signs. Which will be positive, and which will be negative?

\[ \text{Ne} \quad [\text{He}]2s^22p^6 \rightarrow \text{Ne}^+ \quad [\text{He}]2s^22p^5 + 1 \text{ e}^- \quad \text{IE always POSITIVE (requires E)} \]

\[ \text{F} \quad [\text{He}] 2s^22p^5 + 1 \text{ e}^- \rightarrow \text{F}^- \quad [\text{He}] 2s^22p^6 \quad \text{EA always NEGATIVE (E release)} \]