Periodic Properties of the Elements Problem Set

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

7.12 **(a)** How is the concept of effective nuclear charge used to simplify the numerous electron–electron repulsions in a many-electron atom? **(b)** Which experiences a greater effective nuclear charge in a Be atom, the $1s$ electrons or the $2s$ electrons? Explain.

7.13 Detailed calculations show that the value of Z_{eff} for the outermost electrons in Na and K atoms is 2.51+ and 3.49+, respectively. **(a)** What value do you estimate for Z_{eff} experienced by the outermost electron in both Na and K by assuming core electrons contribute 1.00 and valence electrons contribute 0.00 to the screening constant? **(b)** What values do you estimate for Z_{eff} using Slater's rules? **(c)** Which approach gives a more accurate estimate of Z_{eff} ? **(d)** Does either method of approximation account for the gradual increase in Z_{eff} that occurs upon moving down a group? **(e)** Predict Z_{eff} for the outermost electrons in the Rb atom based on the calculations for Na and K.

7.16 Arrange the following atoms in order of increasing effective nuclear charge experienced by the electrons in the $n = 3$ electron shell: K, Mg, P, Rh, and Ti. Explain the basis for your order.

7.17 **(a)** Because an exact outer boundary cannot be measured or even calculated for an atom, how are atomic radii determined? **(b)** What is the difference between a bonding radius and a nonbonding radius? **(c)** For a given element, which one is larger? **(d)** If a free atom reacts to become part of a molecule, would you say that the atom gets smaller or larger?

7.18 **(a)** Why does the quantum mechanical description of many-electron atoms make it difficult to define a precise atomic radius? **(b)** When nonbonded atoms come up against one another, what determines how closely the nuclear centers can approach?

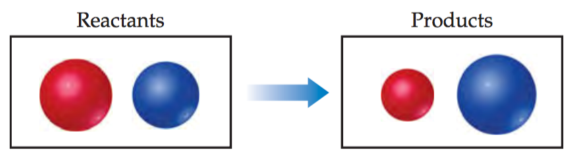
7.19 Tungsten has the highest melting point of any metal in the periodic table: 3422 °C. The distance between W atoms in tungsten metal is 2.74 Å. **(a)** What is the atomic radius of a tungsten atom in this environment? (This radius is called the *metallic radius*.) **(b)** If you put tungsten metal under high pressure, predict what would happen to the distance between W atoms.

7.23 How do the sizes of atoms change as we move **(a)** from left to right across a row in the periodic table, **(b)** from top to bottom in a group in the periodic table? **(c)** Arrange the following atoms in order of increasing atomic radius: O, Si, I, Ge.

7.25 Using only the periodic table, arrange each set of atoms in order from largest to smallest: **(a)** K, Li, Cs; **(b)** Pb, Sn, Si; **(c)** F, O, N.

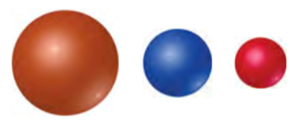
7.28 Explain the following variations in atomic or ionic radii: **(a)** $I^- > I > I^+$, **(b)** $Ca^{2+} > Mg^{2+} > Be^{2+}$, **(c)** $Fe > Fe^{2+} > Fe^{3+}$.

7.29 In the reaction



which sphere represents a metal and which represents a nonmetal? Explain your answer.

7.30 Which of these spheres represents F, which represents Br, and which represents Br^-?



7.31 **(a)** What is an isoelectronic series? **(b)** Which neutral atom is isoelectronic with each of the following ions: Ga^{3+}, Zr^{4+}, Mn^{7+}, I^-, Pb^{2+}?

7.32 Identify at least two ions that have the following ground-state electron configurations: **(a)** [Ar]; **(b)** $[Ar]3d^5$; **(c)** $[Kr]5s^24d^{10}$.

7.33 Some ions do not have a corresponding neutral atom that has the same electron configuration. For each of the following ions, identify the neutral atom that has the same number of electrons and determine if this atom has the same electron configuration. If such an atom does not exist, explain why. **(a)** Cl^-, **(b)** Sc^{3+}, **(c)** Fe^{2+}, **(d)** Zn^{2+}, **(e)** Sn^{4+}.

7.36 Consider S, Cl, and K and their most common ions. **(a)** List the atoms in order of increasing size. **(b)** List the ions in order of increasing size. **(c)** Explain any differences in the orders of the atomic and ionic sizes.

7.37 For each of the following sets of atoms and ions, arrange the members in order of increasing size: **(a)** Se^{2-}, Te^{2-}, Se; **(b)** Co^{3+}, Fe^{2+}, Fe^{3+}; **(c)** Ca, Ti^{4+}, Sc^{3+}; **(d)** Be^{2+}, Na^+, Ne.

7.39 Write equations that show the processes that describe the first, second, and third ionization energies of an aluminum atom. Which process would require the least amount of energy?

7.41 Identify each statement as true or false. If it is false, rewrite it so that it is true: **(a)** Ionization energies are always negative quantitites. **(b)** Oxygen has a larger first ionization energy than fluorine. **(c)** The second ionization energy of an atom is always greater than its first ionization energy.

7.43 **(a)** What is the general relationship between the size of an atom and its first ionization energy? **(b)** Which element in the periodic table has the largest ionization energy? Which has the smallest?

7.45 Based on their positions in the periodic table, predict which atom of the following pairs will have the smaller first ionization energy: **(a)** Cl, Ar; **(b)** Be, Ca; **(c)** K, Co; **(d)** S, Ge; **(e)** Sn, Te.

7.48 Write electron configurations for the following ions, and determine which have noble-gas configurations: **(a)** Cr^{3+}, **(b)** N^{3-}, **(c)** Sc^{3+}, **(d)** Cu^{2+}, **(e)** Tl^{+}, **(f)** Au^{+}.

7.51 The first ionization energy and electron affinity of Ar are both positive values. **(a)** What is the significance of the positive value in each case? **(b)** What are the units of electron affinity?

7.52 If the electron affinity for an element is a negative number, does it mean that the anion of the element is more stable than the neutral atom? Explain.

7.53 Although the electron affinity of bromine is a negative quantity, it is positive for Kr. Use the electron configurations of the two elements to explain the difference.

7.54 What is the relationship between the ionization energy of an anion with a 1− charge such as F^{-} and the electron affinity of the neutral atom, F?

7.57 How are metallic character and first ionization energy related?

7.59 Discussing this chapter, a classmate says, "An element that commonly forms a cation is a metal." Do you agree or disagree? Explain your answer.

7.61 Predict whether each of the following oxides is ionic or molecular: SnO_2, Al_2O_3, CO_2, Li_2O, Fe_2O_3, H_2O. Explain the reasons for your choices.

7.63 **(a)** What is meant by the terms *acidic oxide* and *basic oxide*? **(b)** How can we predict whether an oxide will be acidic or basic based on its composition?

7.64 Arrange the following oxides in order of increasing acidity: CO_2, CaO, Al_2O_3, SO_3, SiO_2, and P_2O_5.

7.69 Does the reactivity of a metal correlate with its first ionization energy? Explain.

7.71 **(a)** Why is calcium generally more reactive than magnesium? **(b)** Why is calcium generally less reactive than potassium?