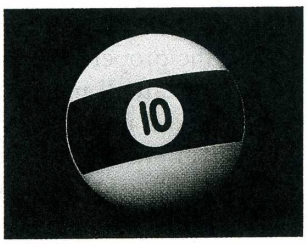
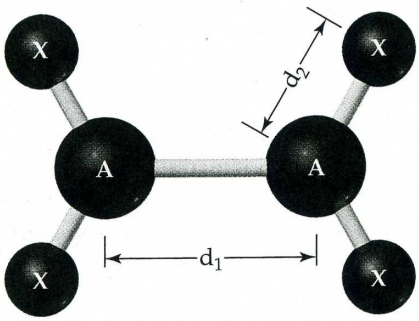
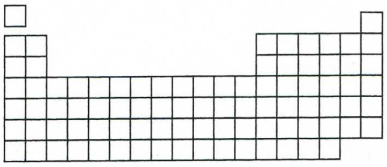
7.2 Neon has atomic number 10. What are the similarities
CQ and differences in describing the radius of a neon atom and the radius of the billiard ball illustrated here? Could you use billiard balls to illustrate the concept of bonding atomic radius? Explain. [Section 7.3]



7.3 Consider the A_2X_4 molecule depicted below, where A
CQ and X are elements. The A—A bond length in this molecule is d_1, and the four A—X bond lengths are each d_2. **(a)** In terms of d_1 and d_2, how could you define the bonding atomic radii of atoms A and X? **(b)** In terms of d_1 and d_2, what would you predict for the X—X bond length of an X_2 molecule? [Section 7.3]



7.4 Make a simple sketch of the shape of the main part of the
CQ periodic table, as shown. **(a)** Ignoring H and He, write a single straight arrow from the element with the smallest bonding atomic radius to the element with the largest. **(b)** Ignoring H and He, write a single straight arrow from the element with the smallest first ionization energy to the element with the largest. **(c)** What significant observation can you make from the arrows you drew in parts (a) and (b)? [Sections 7.3 and 7.4]

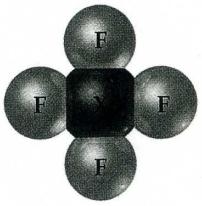


7.5 In the chemical process called *electron transfer*, an elec-
CQ tron is transferred from one atom or molecule to another (we will talk about electron transfer extensively in Chapter 20). A simple electron transfer reaction is

$$A(g) + A(g) \longrightarrow A^+(g) + A^-(g)$$

In terms of the ionization energy and electron affinity of atom A, what is the energy change for this reaction? [Sections 7.4 and 7.5]

7.6 An element X reacts with $F_2(g)$ to form the molecular
CQ product shown below. **(a)** Write a balanced equation for this reaction (do not worry about the phases for X and the product). **(b)** Do you think that X is a metal or nonmetal? Explain. [Section 7.6]



EXERCISES

Periodic Table; Effective Nuclear Charge

7.7 Why did Mendeleev leave blanks in his early version of
CQ the periodic table? How did he predict the properties of the elements that belonged in those blanks?

7.8 **(a)** In the period from about 1800 to about 1865, the
CQ atomic weights of many elements were accurately measured. Why was this important to Mendeleev's formulation of the periodic table? **(b)** What property of the atom did Moseley associate with the wavelength of X rays emitted from an element in his experiments? In what ways did this affect the meaning of the periodic table?

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

7.10 **(a)** How is the concept of effective nuclear charge used
CQ 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 1*s* electrons or the 2*s* electrons? Explain.

7.11 **(a)** If each core electron were totally effective in screening the valence electrons from the full charge of the nucleus and the valence electrons provided no screening for each other, what would be the values of the screening constant, *S*, and the effective nuclear charge, Z_{eff}, for the 4*s* electron in a potassium atom? **(b)** Detailed calculations show that the value of Z_{eff} for a K atom is 3.49+. Explain the difference between this number and the one you obtained in part (a).

7.12 **(a)** If the core electrons were totally effective at shielding the valence electrons from the full charge of the nucleus and the valence electrons provided no shielding for each other, what would be the values of *S* and Z_{eff} for a 3*p* electron in a sulfur atom? **(b)** Detailed calculations indicate that the value of *S* for a 3*p* electron in a sulfur atom is 10.52. Can you explain any difference between this number and the one you obtained in part (a)?

7.13 Which will experience the greater effective nuclear
CQ charge, the electrons in the $n = 3$ shell in Ar or the $n =$ shell in Kr? Which will be closer to the nucleus? Explain.

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

Atomic and Ionic Radii

7.15 Because an exact outer boundary cannot be measured or even calculated for an atom, how are atomic radii determined? What is the difference between a bonding radius and a nonbonding radius?

7.16 **(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?

7.17 The distance between W atoms in tungsten metal is 2.74 Å. What is the atomic radius of a tungsten atom in this environment? (This radius is called the *metallic radius*.)

7.18 Based on the radii presented in Figure 7.6, predict the distance between Ge atoms in solid germanium.

7.19 Estimate the As—I bond length from the data in Figure 7.6, and compare your value to the experimental As—I bond length in arsenic triiodide, AsI_3, 2.55 Å.

7.20 The experimental Bi—I bond length in bismuth triiodide, BiI_3, is 2.81 Å. Based on this value and data in Figure 7.6, predict the atomic radius of Bi.

7.21 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: F, P, S, As.

7.22 **(a)** Among the nonmetallic elements, the change in atomic radius in moving one place left or right in a row is smaller than the change in moving one row up or down. Explain these observations. **(b)** Arrange the following atoms in order of increasing atomic radius: Si, S, Ge, Se.

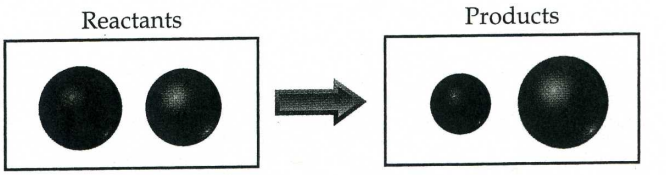
7.23 Using only the periodic table, arrange each set of atoms in order of increasing radius: **(a)** Ca, Mg, Be; **(b)** Ga, Br, Ge; **(c)** Al, Tl, Si.

7.24 Using only the periodic table, arrange each set of atoms in order of increasing radius: **(a)** Cs, K, Rb; **(b)** In, Te, Sn; **(c)** P, Cl, Sr.

7.25 **(a)** Why are monatomic cations smaller than their corresponding neutral atoms? **(b)** Why are monatomic anions larger than their corresponding neutral atoms? **(c)** Why does the size of ions increase as one proceeds down a column in the periodic table?

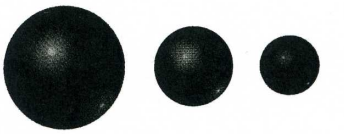
7.26 Explain the following variations in atomic or ionic radii: **(a)** $I^- > I > I^+$, **(b)** $Ca^{2+} > Mg^{2+} > Be^{2+}$, **(c)** $Fe > Fe^{2+} > Fe^{3+}$.

7.27 Consider a reaction represented by the following spheres:



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

7.28 Consider the following spheres:



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

7.29 **(a)** What is an isoelectronic series? **(b)** Which neutral atom is isoelectronic with each of the following ions: (*i*) N^{3-}, (*ii*) Ba^{2+}, (*iii*) Se^{2-}, (*iv*) Bi^{3+}?

7.30 Select the ions or atoms from the following sets that are isoelectronic with each other: **(a)** Na^+, Sr^{2+}, Br^-; **(b)** Y^{3+}, Br^-, Kr; **(c)** N^{3-}, P^{3-}, Ti^{4+}; **(d)** Fe^{3+}, Co^{3+}, Mn^{2+}.

7.31 **(a)** Why do the radii of isoelectronic ions decrease with increasing nuclear charge? **(b)** Which experiences the greatest effective nuclear charge, a 2*p* electron in F^-, a 2*p* electron in Ne, or a 2*p* electron in Na^+?

7.32 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.

7.33 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.

7.34 For each of the following statements, provide an explanation: **(a)** Cl^- is larger than Cl; **(b)** S^{2-} is larger than O^{2-}; **(c)** K^+ is larger than Ca^{2+}.

Ionization Energies; Electron Affinities

7.35 Write equations that show the processes that describe the first, second, and third ionization energies of a boron atom.

7.36 Write equations that show the process for **(a)** the first two ionization energies of tin and **(b)** the fourth ionization energy of titanium.

7.37 **(a)** Why are ionization energies always positive quantities? **(b)** Why does F have a larger first ionization energy than O? **(c)** Why is the second ionization energy of an atom always greater than its first ionization energy?

7.38 **(a)** Why does Li have a larger first ionization energy than Na? **(b)** The difference between the third and fourth ionization energies of scandium is much larger than the difference between the third and fourth ionization energies of titanium. Why? **(c)** Why does Li have a much larger second ionization energy than Be?

7.39 **(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?

7.40 **(a)** What is the trend in first ionization energies as one proceeds down the group 7A elements? Explain how this trend relates to the variation in atomic radii. **(b)** What is the trend in first ionization energies as one moves across the fourth period from K to Kr? How does this trend compare with the trend in atomic sizes?

7.41 **CQ** 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; **(b)** Be, Ca; **(c)** K, Co; **(d)** S, Ge; **(e)** Sn, Te.

7.42 **CQ** For each of the following pairs, indicate which element has the larger first ionization energy: **(a)** Rb, Mo; **(b)** N, P; **(c)** Ga, Cl; **(d)** Pb, Rn. (In each case use electron configuration and effective nuclear charge to explain your answer.)

7.43 Write the electron configurations for the following ions: **(a)** Si^{2+}, **(b)** Bi^{3+}, **(c)** Te^{2-}, **(d)** V^{3+}, **(e)** Hg^{2+}, **(f)** Ni^{2+}.

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

7.45 Write the electron configuration for **(a)** the Mn^{2+} ion and **(b)** the Si^{2-} ion. How many unpaired electrons does each contain?

7.46 Identify the element whose ions have the following electron configurations: **(a)** a 2+ ion with $[Ar]3d^9$, **(b)** a 1+ ion with $[Xe]4f^{14}5d^{10}6s^2$. How many unpaired electrons does each ion contain?

7.47 **CQ** Write equations, including electron configurations beneath the species involved, that explain the difference between the first ionization energy of Se(*g*) and the electron affinity of Se(*g*).

7.48 **CQ** The first ionization energy of Ar and the electron affinity of Ar are both positive values. What is the significance of the positive value in each case?

7.49 **CQ** The electron affinity of lithium is a negative value, whereas the electron affinity of beryllium is a positive value. Use electron configurations to account for this observation.

7.50 **CQ** While 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.

7.51 **CQ** 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?

7.52 **CQ** Write an equation for the process that corresponds to the electron affinity of the Mg^+ ion. Also write the electron configurations of the species involved. What process does this electron affinity equation correspond to? What is the magnitude of the energy change in the process? [*Hint:* The answer is in Table 7.2.]

Properties of Metals and Nonmetals

7.53 **CQ** How are metallic character and first ionization energy related?

7.54 **CQ** Arrange the following pure solid elements in order of increasing electrical conductivity: Ge, Ca, S, and Si. Explain the reasoning you used.

7.55 **CQ** For each of the following pairs, which element will have the greater metallic character: **(a)** Li or Be, **(b)** Li or Na, **(c)** Sn or P, **(d)** Al or B?

7.56 **CQ** **(a)** What data can you cite from this chapter to support a prediction that the metallic character of the group 5A elements will increase with increasing atomic number? **(b)** *Nonmetallic character* is the opposite of metallic character—nonmetallic character decreases as metallic character increases. Arrange the following elements in order of increasing nonmetallic character: Se, Ag, Sn, F, and C.

7.57 **CQ** Predict whether each of the following oxides is ionic or molecular: SO_2, MgO, Li_2O, P_2O_5, Y_2O_3, N_2O, and XeO_3. Explain the reasons for your choices.

7.58 **CQ** When metal oxides react with water, the oxygen generally ends up as the hydroxide ion, separate from the metal. In contrast, when nonmetallic oxides react with water, the oxygen ends up as part of the nonmetal species. (For example, upon reaction of CO_2 with water, the oxygen remains on the carbon in H_2CO_3.) **(a)** Give two examples each from metals and nonmetals to support these generalizations. **(b)** What connection is there between this contrasting behavior of metal and nonmetal oxides and ionization energies?

7.59 **CQ** **(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?

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

7.61 Chlorine reacts with oxygen to form Cl_2O_7. **(a)** What is the name of this product (see Table 2.6)? **(b)** Write a balanced equation for the formation $Cl_2O_7(l)$ from the elements. **(c)** Under usual conditions, Cl_2O_7 is a colorless liquid with a boiling point of 81°C. Is this boiling point expected or surprising? **(d)** Would you expect Cl_2O_7 to be more reactive toward $H^+(aq)$ or $OH^-(aq)$? Explain.

[7.62] An element X reacts with oxygen to form XO_2 and with chlorine to form XCl_4. XO_2 is a white solid that melts at high temperatures (above 1000°C). Under usual conditions, XCl_4 is a colorless liquid with a boiling point of 58°C. **(a)** XCl_4 reacts with water to form XO_2 and another product. What is the likely identity of the other product? **(b)** Do you think that element X is a metal, nonmetal, or metalloid? Explain. **(c)** By using a sourcebook such as the *CRC Handbook of Chemistry and Physics*, try to determine the identity of element X.

[7.63] Write balanced equations for the following reactions: **(a)** barium oxide with water, **(b)** iron(II) oxide with perchloric acid, **(c)** sulfur trioxide with water, **(d)** carbon dioxide with aqueous sodium hydroxide.

7.64 Write balanced equations for the following reactions: **(a)** potassium oxide with water, **(b)** diphosphorus trioxide with water, **(c)** chromium(III) oxide with dilute hydrochloric acid, **(d)** selenium dioxide with aqueous potassium hydroxide.

Group Trends in Metals and Nonmetals

7.65 **CQ** Compare the elements sodium and magnesium with respect to the following properties: **(a)** electron configuration, **(b)** most common ionic charge, **(c)** first ionization energy, **(d)** reactivity toward water, **(e)** atomic radius. Account for the differences between the two elements.

7.66 CQ **(a)** Compare the electron configurations and atomic radii (see Figure 7.6) of rubidium and silver. In what respects are their electronic configurations similar? Account for the difference in radii of the two elements. **(b)** As with rubidium, silver is most commonly found as the 1+ ion, Ag^+. However, silver is far less reactive. Explain these observations.

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

7.68 **(a)** Why is cesium more reactive toward water than is lithium? **(b)** One of the alkali metals reacts with oxygen to form a solid white substance. When this substance is dissolved in water, the solution gives a positive test for hydrogen peroxide, H_2O_2. When the solution is tested in a burner flame, a lilac-purple flame is produced. What is the likely identity of the metal? **(c)** Write a balanced chemical equation for reaction of the white substance with water.

7.69 Write a balanced equation for the reaction that occurs in each of the following cases: **(a)** Potassium metal burns in an atmosphere of chlorine gas. **(b)** Strontium oxide is added to water. **(c)** A fresh surface of lithium metal is exposed to oxygen gas. **(d)** Sodium metal is reacted with molten sulfur.

7.70 Write a balanced equation for the reaction that occurs in each of the following cases: **(a)** Potassium is added to water. **(b)** Barium is added to water. **(c)** Lithium is heated in nitrogen, forming lithium nitride. **(d)** Magnesium burns in oxygen.

7.71 CQ Use electron configurations to explain why hydrogen exhibits properties similar to those of both Li and F.

7.72 **(a)** As described in Section 7.7, the alkali metals react with hydrogen to form hydrides and react with halogens—for example, fluorine—to form halides. Compare the roles of hydrogen and the halogen in these reactions. In what sense are the forms of hydrogen and halogen in the products alike? **(b)** Write balanced equations for the reaction of fluorine with calcium and for the reaction of hydrogen with calcium. What are the similarities among the products of these reactions?

7.73 Compare the elements fluorine and chlorine with respect to the following properties: **(a)** electron configuration, **(b)** most common ionic charge, **(c)** first ionization energy, **(d)** reactivity toward water, **(e)** electron affinity, **(f)** atomic radius. Account for the differences between the two elements.

7.74 CQ Little is known about the properties of astatine, At, because of its rarity and high radioactivity. Nevertheless, it is possible for us to make many predictions about its properties. **(a)** Do you expect the element to be a gas, liquid, or solid at room temperature? Explain. **(b)** What is the chemical formula of the compound it forms with Na?

7.75 CQ Until the early 1960s the group 8A elements were called the inert gases. Why was this name given? Why is it inappropriate?

7.76 Why does xenon react with fluorine whereas neon does not?

7.77 Write a balanced equation for the reaction that occurs in each of the following cases: **(a)** Ozone decomposes to dioxygen. **(b)** Xenon reacts with fluorine. (Write three different equations.) **(c)** Sulfur reacts with hydrogen gas. **(d)** Fluorine reacts with water.

7.78 Write a balanced equation for the reaction that occurs in each of the following cases: **(a)** Chlorine reacts with water. **(b)** Barium metal is heated in an atmosphere of hydrogen gas. **(c)** Lithium reacts with sulfur. **(d)** Fluorine reacts with magnesium metal.

Additional Exercises

7.79 Consider the stable elements through lead ($Z = 82$). In how many instances are the atomic weights of the elements in the reverse order relative to the atomic numbers of the elements? What is the explanation for these cases?

7.80 In 1871, Mendeleev predicted the existence of an element that he called eka-aluminum, which would have the following properties: atomic weight of about 68 amu, density of about 5.9 g/cm^3, low melting point, high boiling point, and an oxide with stoichiometry M_2O_3. **(a)** In 1875 the element predicted by Mendeleev was discovered. By what name is the element known? **(b)** Use a reference such as the *CRC Handbook of Chemistry and Physics* or WebElements.com to check the accuracy of Mendeleev's predictions.

7.81 CQ **(a)** Which will have the lower energy, a 4*s* or a 4*p* electron in an As atom? **(b)** How can we use the concept of effective nuclear charge to explain your answer to part (a)?

7.82 **(a)** If the core electrons were totally effective at shielding the valence electrons and the valence electrons provided no shielding for each other, what would be the effective nuclear charge acting on the valence electron in P? **(b)** Detailed calculations indicate that the effective nuclear charge is 5.6+ for the 3*s* electrons and 4.9+ for the 3*p* electrons. Why are the values for the 3*s* and 3*p* electrons different? **(c)** If you remove a single electron from a P atom, which orbital will it come from? Explain.

7.83 CQ Nearly all the mass of an atom is in the nucleus, which has a very small radius. When atoms bond together (for example, two fluorine atoms in F_2), why is the distance separating the nuclei so much larger than the radii of the nuclei?

[7.84] Consider the change in effective nuclear charge experienced by a 2*p* electron as we proceed from C to N. **(a)** Based on a simple model in which core electrons screen the valence electrons completely and valence electrons do not screen other valence electrons, what do you predict for the change in Z_{eff} from C to N? **(b)** The actual calculated change in Z_{eff} from C to N is 0.70+. How can we explain the difference between this number and the one obtained in part (a)? **(c)** The calculated change in Z_{eff} from N to O is smaller than that from C to N. Can you provide an explanation for this observation?

7.85 CQ As we move across a period of the periodic table, why do the sizes of the transition elements change more gradually than those of the representative elements?

7.86 In the series of group 5A hydrides, of general formula MH_3, the measured bond distances are P—H, 1.419 Å; As—H, 1.519 Å; Sb—H, 1.707 Å. **(a)** Compare these values with those estimated by use of the atomic radii in Figure 7.6. **(b)** Explain the steady increase in M—H bond distance in this series in terms of the electronic configurations of the M atoms.

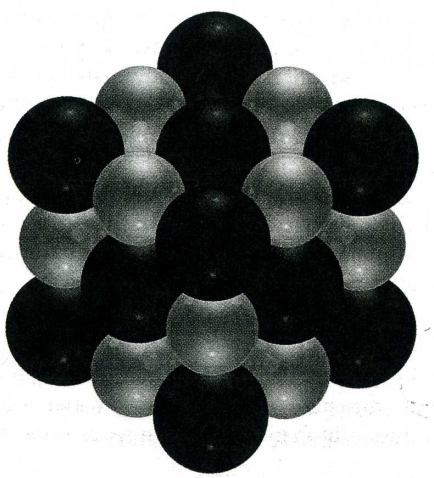
7.87 It is possible to produce compounds of the form $GeClH_3$, $GeCl_2H_2$, and $GeCl_3H$. What values do you predict for the Ge—H and Ge—Cl bond lengths in these compounds?

7.88 Note from the following table that the increase in atomic radius in moving from Zr to Hf is smaller than in moving from Y to La. Suggest an explanation for this effect.

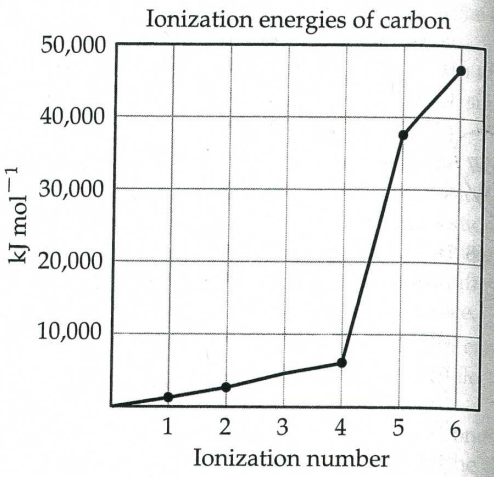
Atomic Radii (Å)			
Sc	1.44	Ti	1.36
Y	1.62	Zr	1.48
La	1.69	Hf	1.50

7.89 The "Chemistry and Life" box on ionic size in Section 7.3 compares the ionic radii of Zn^{2+} and Cd^{2+}. **(a)** The 2+ ion of which other element seems the most obvious one to compare to Zn^{2+} and Cd^{2+}? **(b)** With reference to Figure 2.24, is the element in part (a) essential for life? **(c)** Estimate the ionic radius of the 2+ ion of the element in part (a). Explain any assumptions you have made. **(d)** Would you expect the 2+ ion of the element in part (a) to be physiologically more similar to Zn^{2+} or to Cd^{2+}? **(e)** Use a sourcebook or a Web search to determine whether the element in part (a) is toxic to humans.

[7.90] The ionic substance strontium oxide, SrO, forms from the direct reaction of strontium metal with molecular oxygen. The arrangement of the ions in solid SrO is analogous to that in solid NaCl (see Figure 2.23) and is shown below. **(a)** Write a balanced equation for the formation of SrO(*s*) from the elements. **(b)** In the figure, do the large spheres represent Sr^{2+} ions or O^{2-} ions? Explain. **(c)** Based on the ionic radii in Figure 7.7, predict the length of the side of the cube in the figure. **(d)** The experimental density of SrO is 5.10 g/cm^3. Given your answer to part (c), what is the number of formula units of SrO that are contained in the cube in the figure? (We will examine structures like those in the figure more closely in Chapter 11.)



7.91 CQ Explain the variation in ionization energies of carbon, as displayed in the following graph:



[7.92] Listed here are the atomic and ionic (2+) radii for calcium and zinc:

Radii (Å)	
Ca 1.74	Ca^{2+} 0.99
Zn 1.31	Zn^{2+} 0.74

(a) Explain why the ionic radius in each case is smaller than the atomic radius. **(b)** Why is the atomic radius of calcium larger than that of zinc? **(c)** Suggest a reason why the difference in the ionic radii is much less than the difference in the atomic radii.

7.93 CQ Do you agree with the following statement? "A negative value for the electron affinity of an atom occurs when the outermost electrons only incompletely shield one another from the nucleus." If not, change it to make it more nearly correct in your view. Apply either the statement as given or your revised statement to explain why the electron affinity of bromine is −325 kJ/mol and that for its neighbor Kr is > 0.

7.94 CQ Use orbital diagrams to illustrate what happens when an oxygen atom gains two electrons. Why is it extremely difficult to add a third electron to the atom?

[7.95] Use electron configurations to explain the following observations: **(a)** The first ionization energy of phosphorus is greater than that of sulfur. **(b)** The electron affinity of nitrogen is lower (less negative) than those of both carbon and oxygen. **(c)** The second ionization energy of oxygen is greater than that of fluorine. **(d)** The third ionization energy of manganese is greater than those of both chromium and iron.

7.96 The following table gives the electron affinities, in kJ/mol, for the group 1B and group 2B metals:

Cu −119	Zn > 0
Ag −126	Cd > 0
Au −223	Hg > 0

(a) Why are the electron affinities of the group 2B elements greater than zero? **(b)** Why do the electron affinities of the group 1B elements become more negative as we move down the group? [*Hint:* Examine the trends in the electron affinity of other groups as we proceed down the periodic table.]

7.97 Hydrogen is an unusual element because it behaves in some ways like the alkali metal elements and in other ways like a nonmetal. Its properties can be explained in part by its electron configuration and by the values for its ionization energy and electron affinity. **(a)** Explain why the electron affinity of hydrogen is much closer to the values for the alkali elements than for the halogens. **(b)** Is the following statement true? "Hydrogen has the smallest bonding atomic radius of any element that forms chemical compounds." If not, correct it. If it is, explain in terms of electron configurations. **(c)** Explain why the ionization energy of hydrogen is closer to the values for the halogens than for the alkali metals.

[7.98] The first ionization energy of the oxygen molecule is the energy required for the following process:

$$O_2(g) \longrightarrow O_2^+(g) + e^-$$

The energy needed for this process is 1175 kJ/mol, very similar to the first ionization energy of Xe. Would you expect O_2 to react with F_2? If so, suggest a product or products of this reaction.

7.99 CQ Based on your reading of this chapter, arrange the following in order of increasing melting point: K, Br_2, Mg, and O_2. Explain the factors that determine the order.

[7.100] CQ There are certain similarities in properties that exist between the first member of any periodic family and the element located below it and to the right in the periodic table. For example, in some ways Li resembles Mg, Be resembles Al, and so forth. This observation is called the diagonal relationship. Using what we have learned in this chapter, offer a possible explanation for this relationship.

[7.101] CQ The elements at the bottom of groups 1A, 2A, 6A, 7A, and 8A—Fr, Ra, Po, At, and Rn—are all radioactive. As a result, much less is known about their physical and chemical properties than those of the elements above them. Based on what we have learned in this chapter, which of these five elements would you expect **(a)** to have the most metallic character, **(b)** to have the least metallic (that is, the most nonmetallic) character, **(c)** to have the largest first ionization energy, **(d)** to have the smallest first ionization energy, **(e)** to have the greatest (most negative) electron affinity, **(f)** to have the largest atomic radius, **(g)** to resemble least in appearance the element immediately above it, **(h)** to have the highest melting point, **(i)** to react most exothermically with water?

[7.102] A historian discovers a nineteenth-century notebook in which some observations, dated 1822, on a substance thought to be a new element, were recorded. Here are some of the data recorded in the notebook: Ductile, silver-white, metallic looking. Softer than lead. Unaffected by water. Stable in air. Melting point: 153°C. Density: 7.3 g/cm^3. Electrical conductivity: 20% that of copper. Hardness: About 1% as hard as iron. When 4.20 g of the unknown is heated in an excess of oxygen, 5.08 g of a white solid is formed. The solid could be sublimed by heating to over 800°C. **(a)** Using information in the text and a handbook of chemistry, and making allowances for possible variations in numbers from current values, identify the element reported. **(b)** Write a balanced chemical equation for the reaction with oxygen. **(c)** Judging from Figure 7.2, might this nineteenth-century investigator have been the first to discover a new element?

[7.103] CQ It has been discovered in recent years that many organic compounds that contain chlorine, including dioxins, which had been thought to be entirely of man-made origin, are formed in natural processes. More than 3000 natural organohalogen compounds, most involving chlorine and bromine, are known. These compounds, in which the halogen is attached to carbon, are nearly all nonionic materials. Why are these materials typically not ionic, as are the more abundant inorganic halogen compounds found in nature?

Integrative Exercises

[7.104] Moseley established the concept of atomic number by studying X rays emitted by the elements. The X rays emitted by some of the elements have the following wavelengths:

Element	Wavelength (Å)
Ne	14.610
Ca	3.358
Zn	1.435
Zr	0.786
Sn	0.491

(a) Calculate the frequency, ν, of the X rays emitted by each of the elements, in Hz. **(b)** Using graph paper (or suitable computer software), plot the square root of ν versus the atomic number of the element. What do you observe about the plot? **(c)** Explain how the plot in part (b) allowed Moseley to predict the existence of undiscovered elements. **(d)** Use the result from part (b) to predict the X-ray wavelength emitted by iron. **(e)** A particular element emits X rays with a wavelength of 0.980 Å. What element do you think it is?

[7.105] **(a)** Write the electron configuration for Li, and estimate the effective nuclear charge experienced by the valence electron. **(b)** The energy of an electron in a one-electron atom or ion equals $(-2.18 \times 10^{-18}\ \text{J})\left(\frac{Z^2}{n^2}\right)$, where Z is the nuclear charge and n is the principal quantum number of the electron. Estimate the first ionization energy of Li. **(c)** Compare the result of your calculation with the value reported in Table 7.4, and explain the difference. **(d)** What value of the effective nuclear charge gives the proper value for the ionization energy? Does this agree with your explanation in (c)?

[7.106] One way to measure ionization energies is photoelectron spectroscopy (PES), a technique based on the photoelectric effect. (Section 6.2) In PES, monochromatic light is directed onto a sample, causing electrons to be emitted. The kinetic energy of the emitted electrons is measured. The difference between the energy of the photons and the kinetic energy of the electrons corresponds to the energy needed to remove the electrons (that is, the ionization energy). Suppose that a PES experiment is performed in which mercury vapor is irradiated with ultraviolet light of wavelength 58.4 nm. **(a)** What is the energy of a photon of this light, in eV? **(b)** Write an equation that shows the process corresponding to the first ionization energy of Hg. **(c)** The kinetic energy of the emitted electrons is measured to be 10.75 eV. What is the first ionization energy of Hg, in kJ/mol? **(d)** With reference to Figure 7.11, determine which of the halogen elements has a first ionization energy closest to that of mercury.

7.107 Consider the gas-phase transfer of an electron from a sodium atom to a chlorine atom:

$$Na(g) + Cl(g) \longrightarrow Na^+(g) + Cl^-(g)$$

(a) Write this reaction as the sum of two reactions, one that relates to an ionization energy and one that relates to an electron affinity. **(b)** Use the result from part (a), data in this chapter, and Hess's law to calculate the enthalpy of the above reaction. Is the reaction exothermic or endothermic? **(c)** The reaction between sodium metal and chlorine gas is highly exothermic and produces NaCl(*s*), whose structure was discussed in Section 2.6. Comment on this observation relative to the calculated enthalpy for the aforementioned gas-phase reaction.

[7.108] When magnesium metal is burned in air (Figure 3.5), two products are produced. One is magnesium oxide, MgO. The other is the product of the reaction of Mg with molecular nitrogen, magnesium nitride. When water is added to magnesium nitride, it reacts to form magnesium oxide and ammonia gas. **(a)** Based on the charge of the nitride ion (Table 2.5), predict the formula of magnesium nitride. **(b)** Write a balanced equation for the reaction of magnesium nitride with water. What is the driving force for this reaction? **(c)** In an experiment a piece of magnesium ribbon is burned in air in a crucible. The mass of the mixture of MgO and magnesium nitride after burning is 0.470 g. Water is added to the crucible, further reaction occurs, and the crucible is heated to dryness until the final product is 0.486 g of MgO. What was the mass percentage of magnesium nitride in the mixture obtained after the initial burning? **(d)** Magnesium nitride can also be formed by reaction of the metal with ammonia at high temperature. Write a balanced equation for this reaction. If a 6.3-g Mg ribbon reacts with 2.57 g $NH_3(g)$ and the reaction goes to completion, which component is the limiting reactant? What mass of $H_2(g)$ is formed in the reaction? **(e)** The standard enthalpy of formation of solid magnesium nitride is -461.08 kJ mol^{-1}. Calculate the standard enthalpy change for the reaction between magnesium metal and ammonia gas.

7.109 **(a)** The experimental Bi—Br bond length in bismuth tribromide, $BiBr_3$, is 2.63 Å. Based on this value and the data in Figure 7.6, predict the atomic radius of Bi. **(b)** Bismuth tribromide is soluble in acidic solution. It is formed by treating solid bismuth(III) oxide with aqueous hydrobromic acid. Write a balanced chemical equation for this reaction. **(c)** While bismuth(III) oxide is soluble in acidic solutions, it is insoluble in basic solutions such as NaOH(*aq*). On the basis of these properties, is bismuth characterized as a metallic, metalloid, or nonmetallic element? **(d)** Treating bismuth with fluorine gas forms BiF_5. Use the electron configuration of Bi to explain the formation of a compound with this formulation. **(e)** While it is possible to form BiF_5 in the manner just described, pentahalides of bismuth are not known for the other halogens. Explain why the pentahalide might form with fluorine, but not with the other halogens. How does the behavior of bismuth relate to the fact that xenon reacts with fluorine to form compounds, but not with the other halogens?

7.110 Potassium superoxide, KO_2, is often used in oxygen masks (such as those used by firefighters) because KO_2 reacts with CO_2 to release molecular oxygen. Experiments indicate that 2 mol of $KO_2(s)$ react with each mole of $CO_2(g)$. **(a)** The products of the reaction are $K_2CO_3(s)$ and $O_2(g)$. Write a balanced equation for the reaction between $KO_2(s)$ and $CO_2(g)$. **(b)** Indicate the oxidation number for each atom involved in the reaction in part (a). What elements are being oxidized and reduced? **(c)** What mass of $KO_2(s)$ is needed to consume 18.0 g $CO_2(g)$? What mass of $O_2(g)$ is produced during this reaction?

eMEDIA EXERCISES

These exercises make use of the interactive objects available online in OneKey or the Companion Website, and on your Accelerator CD. Access to these resources comes in your MediaPak.

7.111 The **Periodic Trends: Atomic Radii** animation (*7.3*) describes the trends in the sizes of atoms on the periodic table—from left to right and from top to bottom. **(a)** What factors influence atomic radius? **(b)** Based on the factors that influence the atomic radius, explain why the radius of gallium is smaller than that of aluminum.

7.112 The **Gain and Loss of Electrons** animation (*7.3*) illustrates how addition or subtraction of an electron affects the size of an atom. The first ionization of aluminum produces the Al^+ ion, which is smaller than the neutral Al atom. The second ionization of aluminum produces the Al^{2+} ion, which is smaller still. The third ionization of aluminum produces the Al^{3+} ion, and the 3+ cation is even smaller than the 2+ cation. Of the first, second, and third ionizations, which would you expect to cause the biggest change in size? Explain your reasoning.

7.113 According to the information given in the **Periodic Trends: Ionization Energy** animation (*7.4*), you might expect fluorine and chlorine to have two of the highest