Sample Exercise 7.2 Predicting Relative Sizes of Atomic Radii

Referring to the periodic table, arrange (as much as possible) the atoms B, C, Al, and Si in order of increasing size.

Solution

Analyze and Plan We are given the chemical symbols for four elements and told to use their relative positions in the periodic table to predict the relative size of their atomic radii. We can use the two periodic trends described in the text to help with this problem.

Solve

B and C are in the same period, with C to the right of B. Therefore, we expect the radius of C to be smaller than that of B because radii usually decrease as we move across a period.

radius C < radius B

Al and Si are in the same period, with Si to the right of Al.

radius Si < radius Al

The radius increases as we move down a group with Al and B belonging to the same group, as do C and Si.

radius B < radius Al
radius C < radius Si

Combining these comparisons, we can conclude that C has the smallest radius and Al the largest. Unfortunately, the two periodic trends available to us do not supply enough information to determine the relative sizes of B and Si.

radius C < radius B ~ radius Si < radius Al

Sample Exercise 7.2 Predicting Relative Sizes of Atomic Radii

Continued

Check Referring back to Figure 7.7, we can obtain numerical values for each atomic radius that allow us to say that the radius of Si is greater than that of B.

$$C(0.76\ \text{Å}) < B\ (0.84\ \text{Å}) < Si\ (1.11\ \text{Å}) < Al\ (1.21\ \text{Å})$$

If you examine Figure 7.7 carefully, you will discover that for the sand *p*-block elements the increase in radius moving one element down a column tends to be greater than the increase moving one element left across a row. There are exceptions, however.

Comment Note that the trends we have just discussed are for the *s*- and *p*-block elements. As seen in Figure 7.7, the transition elements do not show a regular decrease moving across a period.

Practice Exercise 1

By referring to the periodic table, but not to Figure 7.7, place the following atoms in order of increasing bonding atomic radius: N, O, P, Ge.

(a) N < O < P < Ge
(b) P < N < O < Ge
(c) O < N < Ge < P
(d) O < N < P < Ge
(e) N < P < Ge < O

Sample Exercise 7.2 Predicting Relative Sizes of Atomic Radii

Continued

Practice Exercise 2

Arrange Be, C, K, and Ca in order of increasing atomic radius.

Sample Exercise 7.3 Predicting Relative Sizes of Atomic and Ionic Radii

Arrange Mg^{2+}, Ca^{2+}, and Ca in order of decreasing radius.

Solution

Cations are smaller than their parent atoms, and so $Ca^{2+} < Ca$. Because Ca is below Mg in group 2A, Ca^{2+} is larger than Mg^{2+}. Consequently, $Ca > Ca^{2+} > Mg^{2+}$.

Practice Exercise 1

Arrange the following atoms and ions in order of increasing ionic radius: F, S^{2-}, Cl, and Se^{2-}.

(a) $F < S^{2-} < Cl < Se^{2-}$
(b) $F < Cl < S^{2-} < Se^{2-}$
(c) $F < S^{2-} < Se^{2-} < Cl$
(d) $Cl < F < Se^{2-} < S^{2-}$
(e) $S^{2-} < F < Se^{2-} < Cl$

Practice Exercise 2

Which of the following atoms and ions is largest: S^{2-}, S, O^{2-}?

Sample Exercise 7.4 Ionic Radii in an Isoelectronic Series

Arrange the ions K^+, Cl^-, Ca^{2+}, and S^{2-} in order of decreasing size.

Solution

This is an isoelectronic series, with all ions having 18 electrons. In such a series, size decreases as nuclear charge (atomic number) increases. The atomic numbers of the ions are S 16, Cl 17, K 19, Ca 20. Thus, the ions decrease in size in the order $S^{2-} > Cl^- > K^+ > Ca^{2+}$.

Practice Exercise 1

Arrange the following ions in order of increasing ionic radius: Br^-, Rb^+, Se^{2-}, Sr^{2+}, Te^{2-}.

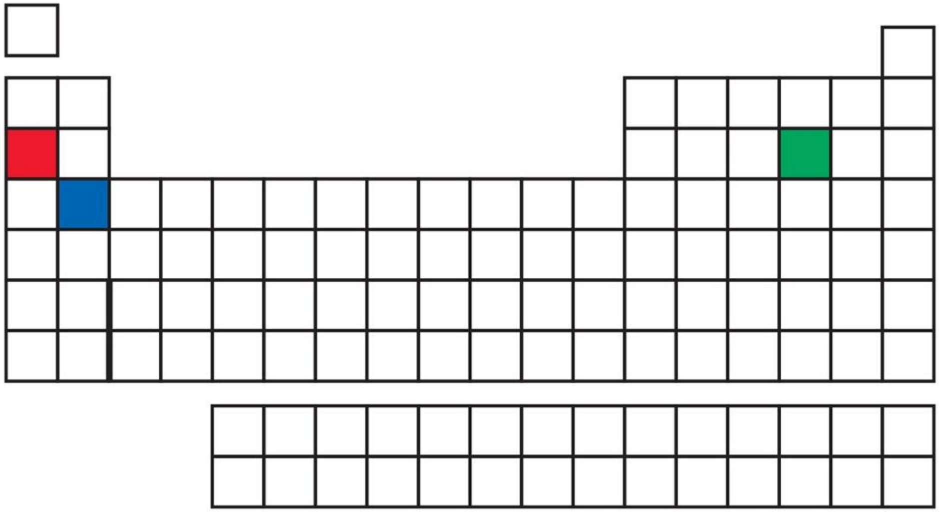
(a) $Sr^{2+} < Rb^+ < Br^- < Se^{2-} < Te^{2-}$
(b) $Br^- < Sr^{2+} < Se^{2-} < Te^{2-} < Rb^+$
(c) $Rb^+ < Sr^{2+} < Se^{2-} < Te^{2-} < Br^-$
(d) $Rb^+ < Br^- < Sr^{2+} < Se^{2-} < Te^{2-}$
(e) $Sr^{2+} < Rb^+ < Br^- < Te^{2-} < Se^{2-}$

Practice Exercise 2

In the isoelectronic series Ca^{2+}, Cs^+, Y^{3+}, which ion is largest?

Sample Exercise 7.5 Trends in Ionization Energy

Three elements are indicated in the periodic table in the margin. Which one has the largest second ionization energy?



Solution

Analyze and Plan The locations of the elements in the periodic table allow us to predict the electron configurations. The greatest ionization energies involve removal of core electrons. Thus, we should look first for an element with only one electron in the outermost occupied shell.

Solve The red box represents Na, which has one valence electron. The second ionization energy of this element is associated, therefore, with the removal of a core electron. The other elements indicated, S (green) and Ca (blue), have two or more valence electrons. Thus, Na should have the largest second ionization energy.

Check A chemistry handbook gives these I_2 values: Ca, 1145 kJ/mol; S, 2252 kJ/mol; Na, 4562 kJ/mol.

Sample Exercise 7.5 Trends in Ionization Energy

Continued

Practice Exercise 1

The third ionization energy of bromine is the energy required for which of the following processes?

(a) $Br(g) \longrightarrow Br^+(g) + e^-$
(b) $Br^+(g) \longrightarrow Br^{2+}(g) + e^-$
(c) $Br(g) \longrightarrow Br^{2+}(g) + 2\ e^-$
(d) $Br(g) \longrightarrow Br^{3+}(g) + 3\ e^-$
(e) $Br^{2+}(g) \longrightarrow Br^{3+}(g) + e^-$

Practice Exercise 2

Which has the greater third ionization energy, Ca or S?

Sample Exercise 7.6 Periodic Trends in Ionization Energy

Referring to the periodic table, arrange the atoms Ne, Na, P, Ar, K in order of increasing first ionization energy.

Solution

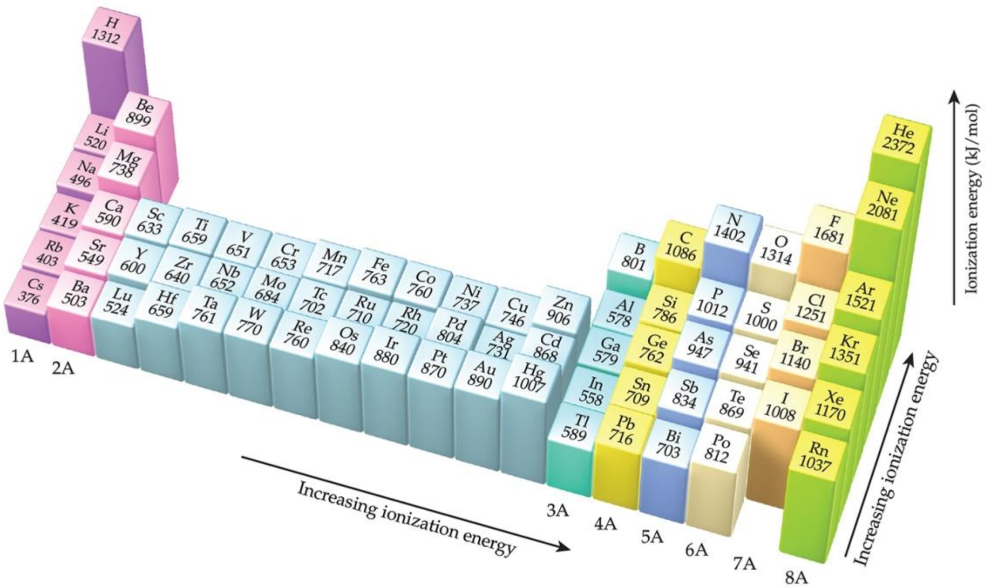
Analyze and Plan We are given the chemical symbols for five elements. To rank them according to increasing first ionization energy, we need to locate each element in the periodic table. We can then use their relative positions and the trends in first ionization energies to predict their order.

Solve Ionization energy increases as we move left to right across a period and decreases as we move down a group. Because Na, P, and Ar are in the same period, we expect I_1 to vary in the order Na < P < Ar. Because Ne is above Ar in group 8A, we expect Ar < Ne. Similarly, K is directly below Na in group 1A, and so we expect K < Na.

From these observations, we conclude that the ionization energies follow the order

$$\text{K} < \text{Na} < \text{P} < \text{Ar} < \text{Ne}$$

Check The values shown in Figure 7.10 confirm this prediction.



Sample Exercise 7.6 Periodic Trends in Ionization Energy

Continued

Practice Exercise 1

Consider the following statements about first ionization energies:

(i) Because the effective nuclear charge for Mg is greater than that for Be, the first ionization energy of Mg is greater than that of Be.

(ii) The first ionization energy of O is less than that of N because in O we must pair electrons in one of the 2*p* orbitals.

(iii) The first ionization energy of Ar is less than that of Ne because a 3*p* electron in Ar is farther from the nucleus than a 2*p* electron in Ne.

Which of the statements (i), (ii), and (iii) is or are true?

(a) Only one of the statements is true.

(b) Statements (i) and (ii) are true.

(c) Statements (i) and (iii) are true.

(d) Statements (ii) and (iii) are true.

(e) All three statements are true.

Practice Exercise 2

Which has the lowest first ionization energy, B, Al, C, or Si? Which has the highest?

Sample Exercise 7.7 Electron Configurations of Ions

Write the electron configurations for (a) Ca^{2+}, (b) Co^{3+}, and (c) S^{2-}.

Solution

Analyze and Plan We are asked to write electron configurations for three ions. To do so, we first write the electron configuration of each parent atom and then remove or add electrons to form the ions. Electrons are first removed from the orbitals having the highest value of n. They are added to the empty or partially filled orbitals having the lowest value of *n*.

Solve

(a) Calcium (atomic number 20) has the electron configuration $[Ar]4s^2$. To form a 2+ ion, the two outer 4*s* electrons must be removed, giving an ion that is isoelectronic with Ar:

$$Ca^{2+}: [Ar]$$

(b) Cobalt (atomic number 27) has the electron configuration $[Ar]4s^2 3d^7$. To form a 3+ ion, three electrons must be removed. As discussed in the text, the 4*s* electrons are removed before the 3*d* electrons. Consequently, we remove the two 4*s* electrons and one of the 3*d* electrons, and the electron configuration for Co^{3+} is

$$Co^{3+} : [Ar]3d^6$$

(c) Sulfur (atomic number 16) has the electron configuration $[Ne]3s^2 3p^4$. To form a 2– ion, two electrons must be added. There is room for two additional electrons in the 3*p* orbitals. Thus, the S^{2-} electron configuration is

$$S^{2-} : [Ne]3s^2 3p^6 = [Ar]$$

Sample Exercise 7.7 Electron Configurations of Ions

Continued

Comment Remember that many of the common ions of the *s*- and *p*-block elements, such as Ca^{2+} and S^{2-}, have the same number of electrons as the closest noble gas. (Section 2.7)

Practice Exercise 1

The ground-state electron configuration of a Tc atom is $[Kr]5s^24d^5$. What is the electron configuration of a Tc^{3+} ion?
(a) $[Kr]4d^4$ **(b)** $[Kr]5s^24d^2$ **(c)** $[Kr]5s^14d^3$ **(d)** $[Kr]5s^24d^8$ **(e)** $[Kr]4d^{10}$

Practice Exercise 2

Write the electron configurations for **(a)** Ga^{3+}, **(b)** Cr^{3+}, and **(c)** Br^-.

Sample Exercise 7.8 Properties of Metal Oxides

(a) Would you expect scandium oxide to be a solid, liquid, or gas at room temperature?
(b) Write the balanced chemical equation for the reaction of scandium oxide with nitric acid.

Solution

Analyze and Plan We are asked about one physical property of scandium oxide—its state at room temperature—and one chemical property—how it reacts with nitric acid.

Solve

(a) Because scandium oxide is the oxide of a metal, we expect it to be an ionic solid. Indeed it is, with the very high melting point of 2485 °C.

(b) In compounds, scandium has a 3+ charge, Sc^{3+}, and the oxide ion is O^{2-}. Consequently, the formula of scandium oxide is Sc_2O_3. Metal oxides tend to be basic and, therefore, to react with acids to form a salt plus water. In this case, the salt is scandium nitrate, $Sc(NO_3)_3$:

$$Sc_2O_3(s) + 6\ HNO_3(aq) \longrightarrow 2\ Sc(NO_3)_3(aq) + 3\ H_2O(l)$$

Sample Exercise 7.8 Properties of Metal Oxides

Continued

Practice Exercise 1

Suppose that a metal oxide of formula M_2O_3 were soluble in water. What would be the major product or products of dissolving the substance in water?

(a) $MH_3(aq) + O_2(g)$
(b) $M(s) + H_2(g) + O_2(g)$
(c) $M^{3+}(aq) + H_2O_2(aq)$
(d) $M(OH)_2(aq)$
(e) $M(OH)_3(aq)$

Practice Exercise 2

Write the balanced chemical equation for the reaction between copper(II) oxide and sulfuric acid.

Sample Exercise 7.9 Reactions of Nonmetal Oxides

Write a balanced chemical equation for the reaction of solid selenium dioxide, $SeO_2(s)$, with **(a)** water, **(b)** aqueous sodium hydroxide.

Solution

Analyze and Plan We note that selenium is a nonmetal. We therefore need to write chemical equations for the reaction of a nonmetal oxide with water and with a base, NaOH. Nonmetal oxides are acidic, reacting with water to form an acid and with bases to form a salt and water.

Solve

(a) The reaction between selenium dioxide and water is like that between carbon dioxide and water (Equation 7.13):

$$SeO_2(s) + H_2O(l) \longrightarrow H_2SeO_3(aq)$$

(It does not matter that SeO_2 is a solid and CO_2 is a gas under ambient conditions; the point is that both are water-soluble nonmetal oxides.)

(b) The reaction with sodium hydroxide is like that in Equation 7.15:

$$SeO_2(s) + 2\ NaOH(aq) \longrightarrow Na_2SeO_3(aq) + H_2O(l)$$

Sample Exercise 7.9 Reactions of Nonmetal Oxides

Continued

Practice Exercise 1

Consider the following oxides: SO_2, Y_2O_3, MgO, Cl_2O, N_2O_5.
How many are expected to form acidic solutions in water?
(a) 1 **(b)** 2 **(c)** 3 **(d)** 4 **(e)** 5

Practice Exercise 2

Write a balanced chemical equation for the reaction of solid tetraphosphorus hexoxide with water.

Sample Exercise 7.10 Reactions of an Alkali Metal

Write a balanced equation for the reaction of cesium metal with (a) $Cl_2(g)$, (b) $H_2O(l)$, (c) $H_2(g)$.

Solution

Analyze and Plan Because cesium is an alkali metal, we expect its chemistry to be dominated by oxidation of the metal to Cs^+ ions. Further, we recognize that Cs is far down the periodic table, which means it is among the most active of all metals and probably reacts with all three substances.

Solve The reaction between Cs and Cl_2 is a simple combination reaction between a metal and a nonmetal, forming the ionic compound CsCl:

$$2\,Cs(s) + Cl_2(g) \longrightarrow 2\,CsCl(s)$$

From Equations 7.18 and 7.16, we predict the reactions of cesium with water and hydrogen to proceed as follows:

$$2\,Cs(s) + 2\,H_2O(l) \longrightarrow 2\,CsOH(aq) + H_2(g)$$
$$2\,Cs(s) + H_2(g) \longrightarrow 2\,CsH(s)$$

All three reactions are redox reactions where cesium forms a Cs^+ ion. The Cl^-, OH^-, and H^- are all 1– ions, which means the products have 1:1 stoichiometry with Cs^+.

Sample Exercise 7.10 Reactions of an Alkali Metal

Continued

Practice Exercise 1

Consider the following three statements about the reactivity of an alkali metal M with oxygen gas:
(i) Based on their positions in the periodic table, the expected product is the ionic oxide M_2O.
(ii) Some of the alkali metals produce metal peroxides or metal superoxides when they react with oxygen.
(iii) When dissolved in water, an alkali metal oxide produces a basic solution.

Which of the statements (i), (ii), and (iii) is or are true?
(a) Only one of the statements is true.
(b) Statements (i) and (ii) are true.
(c) Statements (i) and (iii) are true.
(d) Statements (ii) and (iii) are true.
(e) All three statements are true.

Practice Exercise 2

Write a balanced equation for the expected reaction between potassium metal and elemental sulfur, S(*s*).

Sample Integrative Exercise Putting Concepts Together

The element bismuth (Bi, atomic number 83) is the heaviest member of group 5A. A salt of the element, bismuth subsalicylate, is the active ingredient in Pepto-Bismol®, an over-the-counter medication for gastric distress.

(a) Based on values presented in Figure 7.7 and Tables 7.5 and 7.6, what might you expect for the bonding atomic radius of bismuth?

(b) What accounts for the general increase in atomic radius going down the group 5A elements?

(c) Another major use of bismuth has been as an ingredient in low-melting metal alloys, such as those used in fire sprinkler systems and in typesetting. The element itself is a brittle white crystalline solid. How do these characteristics fit with the fact that bismuth is in the same periodic group with such nonmetallic elements as nitrogen and phosphorus?

(d) Bi_2O_3 is a basic oxide. Write a balanced chemical equation for its reaction with dilute nitric acid. If 6.77 g of Bi_2O_3 is dissolved in dilute acidic solution to make 0.500 L of solution, what is the molarity of the solution of Bi^{3+} ion?

(e) ^{209}Bi is the heaviest stable isotope of any element. How many protons and neutrons are present in this nucleus?

(f) The density of Bi at 25 °C is 9.808 g/cm^3. How many Bi atoms are present in a cube of the element that is 5.00 cm on each edge? How many moles of the element are present?

Sample Integrative Exercise **Putting Concepts Together**

Continued

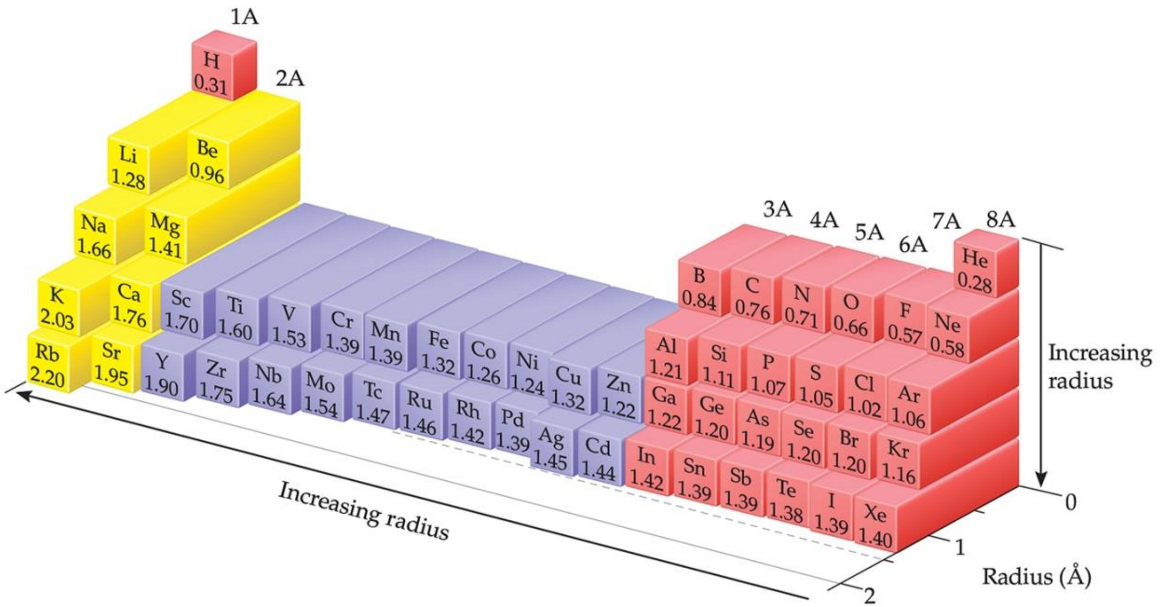


Table 7.5 Some Properties of the Alkaline Earth Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius (Å)	I_1 (kJ/mol)
Beryllium	$[He]2s^2$	1287	1.85	0.96	899
Magnesium	$[Ne]3s^2$	650	1.74	1.41	738
Calcium	$[Ar]4s^2$	842	1.55	1.76	590
Strontium	$[Kr]5s^2$	777	2.63	1.95	549
Barium	$[Xe]6s^2$	727	3.51	2.15	503

Table 7.6 Some Properties of the Group 6A Elements

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Oxygen	$[He]2s^22p^4$	−218	1.43 g/L	0.66	1314
Sulfur	$[Ne]3s^23p^4$	115	1.96 g/cm^3	1.05	1000
Selenium	$[Ar]3d^{10}4s^24p^4$	221	4.82 g/cm^3	1.20	941
Tellurium	$[Kr]4d^{10}5s^25p^4$	450	6.24 g/cm^3	1.38	869
Polonium	$[Xe]4f^{14}5d^{10}6s^26p^4$	254	9.20 g/cm^3	1.40	812

Sample Integrative Exercise Putting Concepts Together

Continued

Solution

(a) Bismuth is directly below antimony, Sb, in group 5A. Based on the observation that atomic radii increase as we go down a column, we would expect the radius of Bi to be greater than that of Sb, which is 1.39 Å. We also know that atomic radii generally decrease as we proceed from left to right in a period. Tables 7.5 and 7.6 each give an element in the same period, namely Ba and Po. We would therefore expect that the radius of Bi is smaller than that of Ba (2.15 Å) and larger than that of Po (1.40 Å). We also see that in other periods, the difference in radius between the neighboring group 5A and group 6A elements is relatively small. We might therefore expect that the radius of Bi is slightly larger than that of Po—much closer to the radius of Po than to the radius of Ba. The tabulated value for the atomic radius on Bi is 1.48 Å, in accord with our expectations.

(b) The general increase in radius with increasing atomic number in the group 5A elements occurs because additional shells of electrons are being added, with corresponding increases in nuclear charge. The core electrons in each case largely screen the outermost electrons from the nucleus, so the effective nuclear charge does not vary greatly as we go to higher atomic numbers. However, the principal quantum number, n, of the outermost electrons steadily increases, with a corresponding increase in orbital radius.

(c) The contrast between the properties of bismuth and those of nitrogen and phosphorus illustrates the general rule that there is a trend toward increased metallic character as we move down in a given group. Bismuth, in fact, is a metal. The increased metallic character occurs because the outermost electrons are more readily lost in bonding, a trend that is consistent with its lower ionization energy.

Sample Integrative Exercise Putting Concepts Together

Continued

(d) Following the procedures described in Section 4.2 for writing molecular and net ionic equations, we have the following:

Molecular equation:

$$Bi_2O_3(s) + 6\ HNO_3(aq) \longrightarrow 2\ Bi(NO_3)_3(aq) + 3\ H_2O(l)$$

Net ionic equation:

$$Bi_2O_3(s) + 6\ H^+(aq) \longrightarrow 2\ Bi^3 + (aq) + 3\ H_2O(l)$$

In the net ionic equation, nitric acid is a strong acid and $Bi(NO_3)_3$ is a soluble salt, so we need to show only the reaction of the solid with the hydrogen ion forming the $Bi^{3+}(aq)$ ion and water. To calculate the concentration of the solution, we proceed as follows (Section 4.5):

$$\frac{6.77\ \cancel{\text{g Bi}_2\text{O}_3}}{0.500\ \text{L soln}} \times \frac{1\ \cancel{\text{mol Bi}_2\text{O}_3}}{466.0\ \cancel{\text{g Bi}_2\text{O}_3}} \times \frac{2\ \text{mol Bi}^{3+}}{1\ \cancel{\text{mol Bi}_2\text{O}_3}} = \frac{0.0581\ \text{mol Bi}^{3+}}{\text{L soln}} = 0.0581\ M$$

(e) Recall that the atomic number of any element is the number of protons and electrons in a neutral atom of the element. (Section 2.3) Bismuth is element 83; there are therefore 83 protons in the nucleus. Because the atomic mass number is 209, there are 209 – 83 = 126 neutrons in the nucleus.

Sample Integrative Exercise Putting Concepts Together

Continued

(f) We can use the density and the atomic weight to determine the number of moles of Bi, and then use Avogadro's number to convert the result to the number of atoms. (Sections 1.4 and 3.4) The volume of the cube is $(5.00)^3\ cm^3 = 125\ cm^3$. Then we have

$(5.00)^3\ cm^3 = 125\ cm^3$. Then we have

$$125\ \cancel{cm^3\ Bi} \times \frac{9.808\ \cancel{g\ Bi}}{1\ \cancel{cm^3}} \times \frac{1\ mol\ Bi}{209.0\ \cancel{g\ Bi}} = 5.87\ mol\ Bi$$

$$5.87\ \cancel{mol\ Bi} \times \frac{6.022 \times 10^{23}\ atom\ Bi}{1\ \cancel{mol\ Bi}} = 3.53 \times 10^{24}\ atoms\ Bi$$