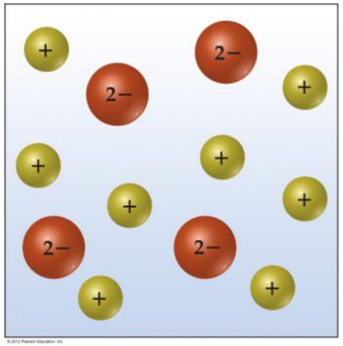
Sample Exercise 4.1 Relating Relative Numbers of Anions and Cations to Chemical Formulas

The accompanying diagram represents an aqueous solution of either $MgCl_2$, KCl, or K_2SO_4. Which solution does the drawing best represent?



Solution

Analyze We are asked to associate the charged spheres in the diagram with ions present in a solution of an ionic substance.

Plan We examine each ionic substance given to determine the relative numbers and charges of its ions .We then correlate these ionic species with the ones shown in the diagram.

Solve The diagram shows twice as many cations as anions, consistent with the formulation K_2SO_4.

Check Notice that the net charge in the diagram is zero, as it must be if it is to represent an ionic substance.

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Sample Exercise 4.1 Relating Relative Numbers of Anions and Cations to Chemical Formulas

Continued

Practice Exercise

If you were to draw diagrams representing aqueous solutions of **(a)** $NiSO_4$, **(b)** $Ca(NO_3)_2$, **(c)** Na_3PO_4, **(d)** $Al_2(SO_4)_3$, how many anions would you show if each diagram contained six cations?

Answer: **(a)** 6, **(b)** 12, **(c)** 2, **(d)** 9

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Sample Exercise 4.2 Using Solubility Rules

Classify these ionic compounds as soluble or insoluble in water: **(a)** sodium carbonate, Na_2CO_3, **(b)** lead sulfate, $PbSO_4$.

Solution

Analyze We are given the names and formulas of two ionic compounds and asked to predict whether they are soluble or insoluble in water.

Plan We can use Table 4.1 to answer the question. Thus, we need to focus on the anion in each compound because the table is organized by anions.

TABLE 4.1 • Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		**Important Exceptions**
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}

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Solve

(a) According to Table 4.1, most carbonates are insoluble. But carbonates of the alkali metal cations (such as sodium ion) are an exception to this rule and are soluble. Thus, Na_2CO_3 is soluble in water.

(b) Table 4.1 indicates that although most sulfates are water soluble, the sulfate of is an exception. Thus, $PbSO_4$ is insoluble in water.

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Sample Exercise 4.2 Using Solubility Rules

Continued

Practice Exercise

Classify the following compounds as soluble or insoluble in water: **(a)** cobalt(II) hydroxide, **(b)** barium nitrate, **(c)** ammonium phosphate.

Answer: **(a)** insoluble, **(b)** soluble, **(c)** soluble

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Sample Exercise 4.3 Predicting a Metathesis Reaction

(a) Predict the identity of the precipitate that forms when aqueous solutions of $BaCl_2$ and K_2SO_4 are mixed.
(b) Write the balanced chemical equation for the reaction.

Solution

Analyze We are given two ionic reactants and asked to predict the insoluble product that they form.

Plan We need to write the ions present in the reactants and exchange the anions between the two cations. Once we have written the chemical formulas for these products, we can use Table 4.1 to determine which is insoluble in water. Knowing the products also allows us to write the equation for the reaction.

TABLE 4.1 • Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}

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Solve

(a) The reactants contain Ba^{2+}, Cl^-, K^+, and SO_4^{2-} ions. Exchanging the anions gives us $BaSO_4$ and KCl. According to Table 4.1, most compounds of SO_4^{2-} are soluble but those of Ba^{2+} are not. Thus, $BaSO_4$ is insoluble and will precipitate from solution. KCl is soluble.

(b) From part **(a)** we know the chemical formulas of the products, $BaSO_4$ and KCl. The balanced equation is

$$BaCl_2(aq) + K_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,KCl(aq)$$

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Sample Exercise 4.3 Predicting a Metathesis Reaction

Continued

Practice Exercise

(a) What compound precipitates when aqueous solutions of $Fe_2(SO_4)_3$ and LiOH are mixed? **(b)** Write a balanced equation for the reaction. **(c)** Will a precipitate form when solutions of $Ba(NO_3)_2$ and KOH are mixed?

Answer: **(a)** $Fe(OH)_3$, **(b)** $Fe_2(SO_4)_3(aq) + 6\ LiOH(aq) \longrightarrow 2\ Fe(OH)_3(s) + 3\ Li_2SO_4(aq)$, **(c)** no (both possible products, $Ba(OH)_2$ and KNO_3, are water soluble)

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Sample Exercise 4.4 Writing a Net Ionic Equation

Write the net ionic equation for the precipitation reaction that occurs when aqueous solutions of calcium chloride and sodium carbonate are mixed.

Solution

Analyze Our task is to write a net ionic equation for a precipitation reaction, given the names of the reactants present in solution.

Plan We write the chemical formulas of the reactants and products and then determine which product is insoluble. We then write and balance the molecular equation. Next, we write each soluble strong electrolyte as separated ions to obtain the complete ionic equation. Finally, we eliminate the spectator ions to obtain the net ionic equation.

Solve Calcium chloride is composed of calcium ions, Ca^{2+}, and chloride ions, Cl^{-}; hence, an aqueous solution of the substance is $CaCl_2(aq)$. Sodium carbonate is composed of Na^{+} ions and $CO3^{2-}$ ions; hence, an aqueous solution of the compound is $Na_2CO_3(aq)$. In the molecular equations for precipitation reactions, the anions and cations appear to exchange partners. Thus, we put Ca^{2+} and CO_3^{2-} together to give $CaCO_3$ and Na^{+} and Cl^{-} together to give NaCl. According to the solubility guidelines in Table 4.1, $CaCO_3$ is insoluble and NaCl is soluble. The balancedmolecular equation is

$$CaCl_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2\ NaCl(aq)$$

TABLE 4.1 • Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^{-}	None
	CH_3COO^{-}	None
	Cl^{-}	Compounds of Ag^{+}, Hg_2^{2+}, and Pb^{2+}
	Br^{-}	Compounds of Ag^{+}, Hg_2^{2+}, and Pb^{2+}
	I^{-}	Compounds of Ag^{+}, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^{+}, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^{+} and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^{+} and the alkali metal cations
	OH^{-}	Compounds of NH_4^{+}, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}

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Sample Exercise 4.4 Writing a Net Ionic Equation

Continued

In a complete ionic equation, *only* dissolved strong electrolytes (such as soluble ionic compounds) are written as separate ions. As the (*aq*) designations remind us, $CaCl_2$, Na_2CO_3, and NaCl are all dissolved in the solution. Furthermore, they are all strong electrolytes. $CaCO_3$ is an ionic compound, but it is not soluble. We do not write the formula of any insoluble compound as its component ions. Thus, the complete ionic equation is

$$Ca^{2+}(aq) + 2\ Cl^-(aq) + 2\ Na^+(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s) + 2\ Na^+(aq) + 2\ Cl^-(aq)$$

Cl^- and Na^+ are spectator ions. Canceling them gives the following net ionic equation:

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

Check We can check our result by confirming that both the elements and the electric charge are balanced. Each side has one Ca, one C, and three O, and the net charge on each side equals 0.

Comment If none of the ions in an ionic equation is removed from solution or changed in some way, all ions are spectator ions and a reaction does not occur.

Practice Exercise

Write the net ionic equation for the precipitation reaction that occurs when aqueous solutions of silver nitrate and potassium phosphate are mixed.

Answer: $3\ Ag^+(aq) + PO_4^{3-}(aq) \longrightarrow Ag_3PO_4(s)$

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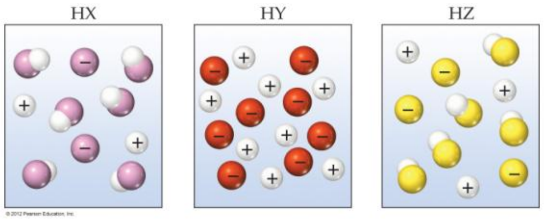
Sample Exercise 4.5 Comparing Acid Strengths

The following diagrams represent aqueous solutions of acids HX, HY, and HZ, with water molecules omitted for clarity. Rank the acids from strongest to weakest.



Solution

Analyze We are asked to rank three acids from strongest to weakest, based on schematic drawings of their solutions.

Plan We can determine the relative numbers of uncharged molecular species in the diagrams. The strongest acid is the one with the most H^+ ions and fewest undissociated molecules in solution. The weakest acid is the one with the largest number of undissociated molecules.

Solve The order is HY > HZ > HX. HY is a strong acid because it is totally ionized (no HY molecules in solution), whereas both HX and HZ are weak acids, whose solutions consist of a mixture of molecules and ions. Because HZ contains more ions and fewer molecules than HX, it is a stronger acid.

Practice Exercise

Imagine a diagram showing 10 Na^+ ions and 10 OH^- ions. If this solution were mixed with the one pictured above for HY, what species would be present in a diagram that represents the combined solutions after any possible reaction?

Answer: The diagram would show 10 Na^+ ions, 2 OH^- ions, 8 Y^- ions, and 8 H_2O molecules.

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Sample Exercise 4.6 Identifying Strong, Weak, and Nonelectrolytes

Classify these dissolved substances as strong, weak, or nonelectrolyte: $CaCl_2$, HNO_3, C_2H_5OH (ethanol), HCOOH (formic acid), KOH.

Solution

Analyze We are given several chemical formulas and asked to classify each substance as a strong electrolyte, weak electrolyte, or nonelectrolyte.

Plan The approach we take is outlined in Table 4.3. We can predict whether a substance is ionic or molecular based on its composition. As we saw in Section 2.7, most ionic compounds we encounter in this text are composed of a metal and a nonmetal, whereas most molecular compounds are composed only of nonmetals.

TABLE 4.3 • Summary of the Electrolytic Behavior of Common Soluble Ionic and Molecular Compounds

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids, weak bases	All other compounds

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Solve Two compounds fit the criteria for ionic compounds: $CaCl_2$ and KOH. Because Table 4.3 tells us that all ionic compounds are strong electrolytes, that is how we classify these two substances. The three remaining compounds are molecular. Two, HNO_3 and HCOOH, are acids. Nitric acid, HNO_3, is a common strong acid, as shown in Table 4.2, and therefore is a strong electrolyte. Because most acids are weak acids, our best guess would be that HCOOH is a weak acid (weak electrolyte). This is correct. The remaining molecular compound, C_2H_5OH, is neither an acid nor a base, so it is a nonelectrolyte.

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Sample Exercise 4.6 Identifying Strong, Weak, and Nonelectrolytes

Continued

Solve Two compounds fit the criteria for ionic compounds: $CaCl_2$ and KOH. Because Table 4.3 tells us that all ionic compounds are strong electrolytes, that is how we classify these two substances. The three remaining compounds are molecular. Two, HNO_3 and HCOOH, are acids. Nitric acid, HNO_3, is a common strong acid, as shown in Table 4.2, and therefore is a strong electrolyte. Because most acids are weak acids, our best guess would be that HCOOH is a weak acid (weak electrolyte). This is correct. The remaining molecular compound, C_2H_5OH, is neither an acid nor a base, so it is a nonelectrolyte.

TABLE 4.2 • Common Strong Acids and Bases

Strong Acids	Strong Bases
Hydrochloric, HCl	Group 1A metal hydroxides [LiOH, NaOH, KOH, RbOH, CsOH]
Hydrobromic, HBr	Heavy group 2A metal hydroxides [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$]
Hydroiodic, HI	
Chloric, $HClO_3$	
Perchloric, $HClO_4$	
Nitric, HNO_3	
Sulfuric, H_2SO_4	

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Comment Although C_2H_5OH has an OH group, it is not a metal hydroxide and so not a base. Rather, it is a member of a class of organic compounds that have C—OH bonds, which are known as alcohols. (Section 2.9) Organic compounds containing the COOH group are called carboxylic acids (Chapter 16).Molecules that have this group are weak acids.

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Sample Exercise 4.6 Identifying Strong, Weak, and Nonelectrolytes

Continued

Practice Exercise

Consider solutions in which 0.1 mol of each of the following compounds is dissolved in 1 L of water: $Ca(NO_3)_2$ (calcium nitrate), $C_6H_{12}O_6$ (glucose), $NaCH_3COO$ (sodium acetate), and CH_3COOH (acetic acid). Rank the solutions in order of increasing electrical conductivity, based on the fact that the greater the number of ions in solution, the greater the conductivity.

Answer: $C_6H_{12}O_6$ (nonelectrolyte) < CH_3COOH (weak electrolyte, existing mainly in the form of molecules with few ions) < $NaCH_3COO$ (strong electrolyte that provides two ions, Na^+ and CH_3COO^-) < $Ca(NO_3)_2$ (strong electrolyte that provides three ions, Ca^{2+} and 2 NO_3^-)

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Sample Exercise 4.7 Writing Chemical Equations for a Neutralization Reaction

For the reaction between aqueous solutions of acetic acid (CH_3COOH) and barium hydroxide, $Ba(OH)_2$, write **(a)** the balanced molecular equation, **(b)** the complete ionic equation, **(c)** the net ionic equation.

Solution

Analyze We are given the chemical formulas for an acid and a base and asked to write a balanced molecular equation, a complete ionic equation, and a net ionic equation for their neutralization reaction.

Plan As Equation 4.12 and the italicized statement that follows it indicate, neutralization reactions form two products, H_2O and a salt. We examine the cation of the base and the anion of the acid to determine the composition of the salt.

Solve

(a) The salt contains the cation of the base (Ba^{2+}) and the anion of the acid (CH_3COO^-). Thus, the salt formula is $Ba(CH_3COO)_2$. According to Table 4.1, this compound is soluble in water. The unbalanced molecular equation for the neutralization reaction is

$$CH_3COOH(aq) + Ba(OH)_2(aq) \longrightarrow H_2O(l) + Ba(CH_3COO)_2(aq)$$

TABLE 4.1 • Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}

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Sample Exercise 4.7 Writing Chemical Equations for a Neutralization Reaction

Continued

To balance this equation, we must provide two molecules of CH_3COOH to furnish the two CH_3COO^- ions and to supply the two H^+ ions needed to combine with the two OH^- ions of the base. The balanced molecular equation is

$$2\,CH_3COOH(aq) + Ba(OH)_2(aq) \longrightarrow 2\,H_2O(l) + Ba(CH_3COO)_2(aq)$$

(b) To write the complete ionic equation, we identify the strong electrolytes and break them into ions. In this case $Ba(OH)_2$ and $Ba(CH_3COO)_2$ are both water-soluble ionic compounds and hence strong electrolytes. Thus, the complete ionic equation is

$$2\,CH_3COOH(aq) + Ba^{2+}(aq) + 2\,OH^-(aq) \longrightarrow 2\,H_2O(l) + Ba^{2+}(aq) + 2\,CH_3COO^-(aq)$$

(c) Eliminating the spectator ion, Ba^{2+}, and simplifying coefficients gives the net ionic equation:

$$2\,CH_3COOH(aq) + 2\,OH^-(aq) \longrightarrow 2\,H_2O(l) + 2\,CH_3COO^-(aq)$$

$$CH_3COOH\,(aq) + OH^-(aq) \longrightarrow H_2O(l) + CH_3COO^-(aq)$$

Check We can determine whether the molecular equation is balanced by counting the number of atoms of each kind on both sides of the arrow (10 H, 6 O, 4 C, and 1 Ba on each side). However, it is often easier to check equations by counting groups: There are 2 CH_3COO groups, as well as 1 Ba, and 4 additional H atoms and 2 additional O atoms on each side of the equation. The net ionic equation checks out because the numbers of each kind of element and the net charge are the same on both sides of the equation.

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Sample Exercise 4.7 Writing Chemical Equations for a Neutralization Reaction

Continued

Practice Exercise

For the reaction of phosphorous acid (H_3PO_3) and potassium hydroxide (KOH), write **(a)** the balanced molecular equation and **(b)** the net ionic equation.

Answer: **(a)** $H_3PO_3(aq) + 3\ KOH(aq) \longrightarrow 3\ H_2O(l) + K_3PO_3(aq)$,
(b) $H_3PO_3(aq) + 3\ OH^-(aq) \longrightarrow 3\ H_2O(l) + PO_3^{3-}(aq)$. ($H_3PO_3$ is a weak acid and therefore a weak electrolyte, whereas KOH, a strong base, and K_3PO_3, an ionic compound, are strong electrolytes.)

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Sample Exercise 4.8 Determining Oxidation Numbers

Determine the oxidation number of sulfur in **(a)** H_2S, **(b)** S_8, **(c)** SCl_2, **(d)** Na_2SO_3, **(e)** SO_4^{2-}.

Solution

Analyze We are asked to determine the oxidation number of sulfur in two molecular species, in the elemental form, and in two substances containing ions.

Plan In each species the sum of oxidation numbers of all the atoms must equal the charge on the species. We will use the rules outlined previously to assign oxidation numbers.

Solve

(a) When bonded to a nonmetal, hydrogen has an oxidation number of +1 (rule 3b). Because the H_2S molecule is neutral, the sum of the oxidation numbers must equal zero (rule 4). Letting x equal the oxidation number of S, we have $2(+1) + x = 0$. Thus, S has an oxidation number of –2.

(b) Because this is an elemental form of sulfur, the oxidation number of S is 0 (rule 1).

(c) Because this is a binary compound, we expect chlorine to have an oxidation number of –1 (rule 3c). The sum of the oxidation numbers must equal zero (rule 4). Letting x equal the oxidation number of S, we have $x + 2(-1) = 0$. Consequently, the oxidation number of S must be +2.

(d) Sodium, an alkali metal, always has an oxidation number of +1 in its compounds (rule 2). Oxygen has a common oxidation state of –2 (rule 3a). Letting x equal the oxidation number of S, we have $2(+1) + x + 3(-2) = 0$. Therefore, the oxidation number of S in this compound is +4.

(e) The oxidation state of O is –2 (rule 3a). The sum of the oxidation numbers equals –2, the net charge of the SO_4^{2-} ion (rule 4). Thus, we have $x + 4(-2) = -2$. From this relation we conclude that the oxidation number of S in this ion is +6.

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Sample Exercise 4.8 Determining Oxidation Numbers

Continued

Comment These examples illustrate that the oxidation number of a given element depends on the compound in which it occurs. The oxidation numbers of sulfur, as seen in these examples, range from –2 to +6.

Practice Exercise

What is the oxidation state of the boldfaced element in **(a)** $\mathbf{P}_2O_5$, **(b)** $Na\mathbf{H}$, **(c)** $\mathbf{Cr}_2O_7^{2-}$, **(d)** $\mathbf{Sn}Br_4$, **(e)** $Ba\mathbf{O}_2$?

Answer: **(a)** +5, **(b)** –1, **(c)** +6, **(d)** +4, **(e)** –1

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Sample Exercise 4.9 Writing Equations for Oxidation-Reduction Reactions

Write the balanced molecular and net ionic equations for the reaction of aluminum with hydrobromic acid.

Solution

Analyze We must write two equations—molecular and net ionic—for the redox reaction between a metal and an acid.

Plan Metals react with acids to form salts and H_2 gas. To write the balanced equations, we must write the chemical formulas for the two reactants and then determine the formula of the salt, which is composed of the cation formed by the metal and the anion of the acid.

Solve The reactants are Al and HBr. The cation formed by Al is Al^{3+}, and the anion from hydrobromic acid is Br^-. Thus, the salt formed in the reaction is $AlBr_3$.Writing the reactants and products and then balancing the equation gives the molecular equation:

$$2\,Al(s) + 6\,HBr(aq) \longrightarrow 2\,AlBr_3(aq) + 3\,H_2(g)$$

Both HBr and $AlBr_3$ are soluble strong electrolytes. Thus, the complete ionic equation is

$$2\,Al(s) + 6\,H^+(aq) + 6\,Br^-(aq) \longrightarrow 2\,Al^{3+}(aq) + 6\,Br^-(aq) + 3\,H_2(g)$$

Because Br^- is a spectator ion, the net ionic equation is

$$2\,Al(s) + 6\,H^+(aq) \longrightarrow 2\,Al^{3+}(aq) + 3\,H_2(g)$$

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Sample Exercise 4.9 Writing Equations for Oxidation-Reduction Reactions

Continued

Comment The substance oxidized is the aluminum metal because its oxidation state changes from 0 in the metal to +3 in the cation, thereby increasing in oxidation number. The H^+ is reduced because its oxidation state changes from +1 in the acid to 0 in H_2.

Practice Exercise

(a) Write the balanced molecular and net ionic equations for the reaction between magnesium and cobalt(II) sulfate. **(b)** What is oxidized and what is reduced in the reaction?

Answer: **(a)** $Mg(s) + CoSO_4(aq) \longrightarrow MgSO_4(aq) + Co(s)$; $Mg(s) + Co^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Co(s)$,
(b) Mg is oxidized and Co^{2+} is reduced.

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Sample Exercise 4.10 Determining When an Oxidation-Reduction Reaction Can Occur

Will an aqueous solution of iron(II) chloride oxidize magnesium metal? If so, write the balanced molecular and net ionic equations for the reaction.

Solution

Analyze We are given two substances—an aqueous salt, $FeCl_2$, and a metal, Mg—and asked if they react with each other.

Plan A reaction occurs if the reactant that is a metal in its elemental form (Mg) is located above the reactant that is a metal in its oxidized form (Fe^{2+}) in Table 4.5. If the reaction occurs, the Fe^{2+} ion in $FeCl_2$ is reduced to Fe, and the Mg is oxidized to Mg^{2+}.

TABLE 4.5 • Activity Series of Metals in Aqueous Solution

Metal	Oxidation Reaction
Lithium	$Li(s) \longrightarrow Li^+(aq) + e^-$
Potassium	$K(s) \longrightarrow K^+(aq) + e^-$
Barium	$Ba(s) \longrightarrow Ba^{2+}(aq) + 2e^-$
Calcium	$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^-$
Sodium	$Na(s) \longrightarrow Na^+(aq) + e^-$
Magnesium	$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$
Aluminum	$Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$
Manganese	$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^-$
Zinc	$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$
Chromium	$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^-$
Iron	$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$
Cobalt	$Co(s) \longrightarrow Co^{2+}(aq) + 2e^-$
Nickel	$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^-$
Tin	$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^-$
Lead	$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^-$
Hydrogen	$H_2(g) \longrightarrow 2H^+(aq) + 2e^-$
Copper	$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$
Silver	$Ag(s) \longrightarrow Ag^+(aq) + e^-$
Mercury	$Hg(l) \longrightarrow Hg^{2+}(aq) + 2e^-$
Platinum	$Pt(s) \longrightarrow Pt^{2+}(aq) + 2e^-$
Gold	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^-$

Ease of oxidation increases

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Sample Exercise 4.10 Determining When an Oxidation-Reduction Reaction Can Occur

Continued

Solve Because Mg is above Fe in the table, the reaction occurs. To write the formula for the salt produced in the reaction, we must remember the charges on common ions. Magnesium is always present in compounds as Mg^{2+}; the chloride ion is Cl^-. The magnesium salt formed in the reaction is $MgCl_2$, meaning the balanced molecular equation is

$$Mg(s) + FeCl_2(aq) \longrightarrow MgCl_2(aq) + Fe(s)$$

Both $FeCl_2$ and $MgCl_2$ are soluble strong electrolytes and can be written in ionic form, which shows us that Cl^- is a spectator ion in the reaction. The net ionic equation is

$$Mg(s) + Fe^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Fe(s)$$

The net ionic equation shows that Mg is oxidized and Fe^{2+} is reduced in this reaction.

Check Note that the net ionic equation is balanced with respect to both charge and mass.

Practice Exercise

Which of the following metals will be oxidized by $Pb(NO_3)_2$: Zn, Cu, Fe?

Answer: Zn and Fe

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Sample Exercise 4.11 Calculating Molarity

Calculate the molarity of a solution made by dissolving 23.4 g of sodium sulfate (Na_2SO_4) in enough water to form 125 mL of solution.

Solution

Analyze We are given the number of grams of solute (23.4 g), its chemical formula (Na_2SO_4), and the volume of the solution (125 mL) and asked to calculate the molarity of the solution.

Plan We can calculate molarity using Equation 4.32. To do so, we must convert the number of grams of solute to moles and the volume of the solution from milliliters to liters.

Solve The number of moles of Na_2SO_4 is obtained by using its molar mass:

$$\text{Moles } Na_2SO_4 = (23.4\ \cancel{\text{g } Na_2SO_4})\left(\frac{1\text{ mol } Na_2SO_4}{142\ \cancel{\text{g } Na_2SO_4}}\right) = 0.165\text{ mol } Na_2SO_4$$

Converting the volume of the solution to liters:

$$\text{Liters soln} = (125\ \cancel{\text{mL}})\left(\frac{1\text{ L}}{1000\ \cancel{\text{mL}}}\right) = 0.125\text{ L}$$

Thus, the molarity is

$$\text{Molarity} = \frac{0.165\text{ mol } Na_2SO_4}{0.125\text{ L soln}} = 1.32\frac{\text{mol } Na_2SO_4}{\text{L soln}} = 1.32\ M$$

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Sample Exercise 4.11 Calculating Molarity

Continued

Check Because the numerator is only slightly larger than the denominator, it is reasonable for the answer to be a little over 1 M. The units are appropriate for molarity, and three significant figures are appropriate for the answer because each of the initial pieces of data had three significant figures.

Practice Exercise

Calculate the molarity of a solution made by dissolving 5.00 g of glucose ($C_6H_{12}O_6$) in sufficient water to form exactly 100 mL of solution.

Answer: 0.278 M

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Sample Exercise 4.12 Calculating Molar Concentrations of Ions

What is the molar concentration of each ion present in a 0.025 *M* aqueous solution of calcium nitrate?

Solution

Analyze We are given the concentration of the ionic compound used to make the solution and asked to determine the concentrations of the ions in the solution.

Plan We can use the subscripts in the chemical formula of the compound to determine the relative ion concentrations.

Solve Calcium nitrate is composed of calcium ions (Ca^{2+}) and nitrate ions (NO_3^-), so its chemical formula is $Ca(NO_3)_2$. Because there are two NO_3^- ions for each Ca^{2+} ion, each mole of $Ca(NO_3)_2$ that dissolves dissociates into 1 mol of Ca^{2+} and 2 mol of NO_3^-. Thus, a solution that is 0.025 *M* in $Ca(NO_3)_2$ is 0.025 *M* in Ca^{2+} and 2 × 0.025 *M* = 0.050 *M* in NO_3^-:

$$\frac{\text{mol } NO_3^-}{\text{L}} = \left(\frac{0.025 \cancel{\text{mol } Ca(NO_3)_2}}{\text{L}}\right)\left(\frac{2 \text{ mol } NO_3^-}{1 \cancel{\text{mol } Ca(NO_3)_2}}\right) = 0.050\ M$$

Check The concentration of NO_3^- ions is twice that of Ca^{2+} ions, as the subscript 2 after the NO_3^- in the chemical formula $Ca(NO_3)_2$ suggests it should be.

Practice Exercise

What is the molar concentration of K^+ ions in a 0.015 *M* solution of potassium carbonate?

Answer: 0.030 *M*

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Sample Exercise 4.13 Using Molarity to Calculate Grams of Solute

How many grams of Na_2SO_4 are required to make 0.350 L of 0.500 *M* Na_2SO_4?

Solution

Analyze We are given the volume of the solution (0.350 L), its concentration (0.500 *M*), and the identity of the solute Na_2SO_4 and asked to calculate the number of grams of the solute in the solution.

Plan We can use the definition of molarity (Equation 4.32) to determine the number of moles of solute, and then convert moles to grams using the molar mass of the solute.

$$M_{Na_2SO_4} = \frac{\text{moles } Na_2SO_4}{\text{liters soln}}$$

Solve Calculating the moles of Na_2SO_4 using the molarity and volume of solution gives

$$M_{Na_2SO_4} = \frac{\text{moles } Na_2SO_4}{\text{liters soln}}$$

$$\text{Moles } Na_2SO_4 = \text{liters soln} \times M_{Na_2SO_4}$$

$$= (0.350\ \cancel{\text{L soln}})\left(\frac{0.500\text{ mol } Na_2SO_4}{1\ \cancel{\text{L soln}}}\right)$$

$$= 0.175\text{ mol } Na_2SO_4$$

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Sample Exercise 4.13 Using Molarity to Calculate Grams of Solute

Continued

Because each mole of Na_2SO_4 has a mass of 142 g, the required number of grams of Na_2SO_4 is

$$\text{Grams } Na_2SO_4 = (0.175\ \cancel{\text{mol } Na_2SO_4})\left(\frac{142 \text{ g } Na_2SO_4}{1\ \cancel{\text{mol } Na_2SO_4}}\right) = 24.9 \text{ g } Na_2SO_4$$

Check The magnitude of the answer, the units, and the number of significant figures are all appropriate.

Practice Exercise

(a) How many grams of Na_2SO_4 are there in 15 mL of 0.50 *M* Na_2SO_4? **(b)** How many milliliters of 0.50 *M* Na_2SO_4 solution are needed to provide 0.038 mol of this salt?

Answer: **(a)** 1.1 g, **(b)** 76 mL

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Sample Exercise 4.14 Preparing a Solution by Dilution

How many milliliters of 3.0 *M* H_2SO_4 are needed to make 450 mL of 0.10 *M* H_2SO_4?

Solution

Analyze We need to dilute a concentrated solution. We are given the molarity of a more concentrated solution (3.0 *M*) and the volume and molarity of a more dilute one containing the same solute (450 mL of 0.10 *M* solution). We must calculate the volume of the concentrated solution needed to prepare the dilute solution.

Plan We can calculate the number of moles of solute, H_2SO_4, in the dilute solution and then calculate the volume of the concentrated solution needed to supply this amount of solute. Alternatively, we can directly apply Equation 4.34. Let's compare the two methods.

Solve Calculating the moles of H_2SO_4 in the dilute solution:

$$\text{Moles } H_2SO_4 \text{ in dilute solution} = (0.450\ \cancel{\text{L soln}})\left(\frac{0.10 \text{ mol } H_2SO_4}{1\ \cancel{\text{L soln}}}\right)$$
$$= 0.045 \text{ mol } H_2SO_4$$

Calculating the volume of the concentrated solution that contains 0.045 mol H_2SO_4:

$$\text{L conc soln} = (0.045\ \cancel{\text{mol } H_2SO_4})\left(\frac{1 \text{ L soln}}{\cancel{3.0 \text{ mol } H_2SO_4}}\right) = 0.015 \text{ L soln}$$

Converting liters to milliliters gives 15 mL.

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Sample Exercise 4.14 Preparing a Solution by Dilution

Continued

If we apply Equation 4.34, we get the same result:

$$(3.0\ M)(V_{\text{conc}}) = (0.10\ M)(450\ \text{mL})$$

$$(V_{\text{conc}}) = \frac{(0.10\ \cancel{M})(450\ \text{mL})}{3.0\ \cancel{M}} = 15\ \text{mL}$$

Either way, we see that if we start with 15 mL of 3.0 *M* H_2SO_4 and dilute it to a total volume of 450 mL, the desired 0.10 *M* solution will be obtained.

Check The calculated volume seems reasonable because a small volume of concentrated solution is used to prepare a large volume of dilute solution.

Comment The first approach can also be used to find the final concentration when two solutions of different concentrations are mixed, whereas the second approach, using Equation 4.34, can be used only for diluting a concentrated solution with pure solvent.

Practice Exercise

(a) What volume of 2.50 *M* lead(II) nitrate solution contains 0.0500 mol of Pb^{2+}? **(b)** How many milliliters of 5.0 *M* $K_2Cr_2O_7$ solution must be diluted to prepare 250 mL of 0.10 *M* solution? **(c)** If 10.0 mL of a 10.0 *M* stock solution of NaOH is diluted to 250 mL, what is the concentration of the resulting stock solution?

Answer: **(a)** 0.0200 L = 20.0 mL, **(b)** 5.0 mL, **(c)** 0.40 *M*

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Sample Exercise 4.15 Using Mass Relations in a Neutralization Reaction

How many grams of $Ca(OH)_2$ are needed to neutralize 25.0 mL of 0.100 *M* HNO_3?

Solution

Analyze The reactants are an acid, HNO_3, and a base, $Ca(OH)_2$. The volume and molarity of HNO_3 are given, and we are asked how many grams of $Ca(OH)_2$ are needed to neutralize this quantity of HNO_3.

Plan Following the steps outlined by the green arrows in Figure 4.18, we use the molarity and volume of the HNO_3 solution (substance B in Figure 4.18) to calculate the number of moles of HNO_3. We then use the balanced equation to relate moles of HNO_3 to moles of $Ca(OH)_2$ (substance A). Finally, we use the molar mass to convert moles to grams of Ca(OH)2:

$$V_{HNO_3} \times M_{HNO_3} \Rightarrow \text{mol } HNO_3 \Rightarrow \text{mol } Ca(OH)_2 \Rightarrow \text{g } Ca(OH)_2$$

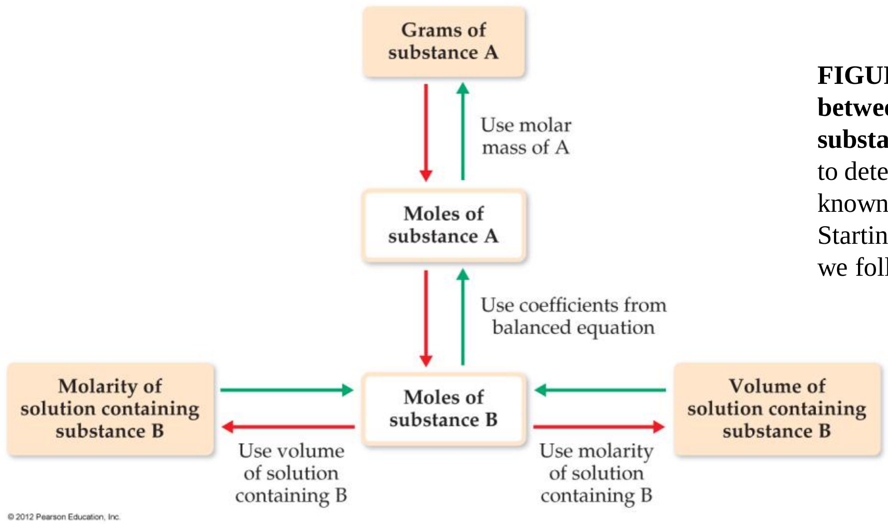

FIGURE 4.18 Procedure for solving stoichiometry problems involving reactions between a pure substance A and a solution containing a known concentration of substance B. Starting from a known mass of substance A, we follow the red arrows to determine either the volume of the solution containing B (if the molarity of B is known) or the molarity of the solution containing B (if the volume of B is known). Starting from either a known volume or known molarity of the solution containing B, we follow the green arrows to determine the mass of substance A.

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Sample Exercise 4.15 Using Mass Relations in a Neutralization Reaction

Continued

Solve The product of the molar concentration of a solution and its volume in liters gives the number of moles of solute:

$$\text{Moles HNO}_3 = V_{\text{HNO}_3} \times M_{\text{HNO}_3} = (0.0250\ \cancel{\text{L}})\left(\frac{0.100\ \text{mol HNO}_3}{\cancel{\text{L}}}\right)$$
$$= 2.50 \times 10^{-3}\ \text{mol HNO}_3$$

Because this is a neutralization reaction, HNO_3 and $Ca(OH)_2$ react to form H_2O and the salt containing Ca^{2+} and NO_3^-:

$$2\ \text{HNO}_3(aq) + \text{Ca(OH)}_2(s) \longrightarrow 2\ \text{H}_2\text{O}(l) + \text{Ca(NO}_3)_2(aq)$$

Thus, 2 mol $HNO_3 \simeq$ mol $Ca(OH)_2$. Therefore,

$$\text{Grams Ca(OH)}_2 = (2.50 \times 10^{-3}\ \cancel{\text{mol HNO}_3})\left(\frac{\cancel{1\ \text{mol Ca(OH)}_2}}{\cancel{2\ \text{mol HNO}_3}}\right)\left(\frac{74.1\ \text{g Ca(OH)}_2}{\cancel{1\ \text{mol Ca(OH)}_2}}\right)$$
$$= 0.0926\ \text{g Ca(OH)}_2$$

Check The answer is reasonable because a small volume of dilute acid requires only a small amount of base to neutralize it.

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Sample Exercise 4.15 Using Mass Relations in a Neutralization Reaction

Continued

Practice Exercise

(a) How many grams of NaOH are needed to neutralize 20.0 mL of 0.150 *M* H_2SO_4 solution? **(b)** How many liters of 0.500 *M* $HCl(aq)$ are needed to react completely with 0.100 mol of $Pb(NO_3)_2(aq)$, forming a precipitate of $PbCl_2(s)$?

Answer: **(a)** 0.240 g, **(b)** 0.400 L

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Sample Exercise 4.16 Determining Solution Concentration by an Acid–Base Titration

One commercial method used to peel potatoes is to soak them in a NaOH solution for a short time, then remove them and spray off the peel. The NaOH concentration is normally 3 to 6 *M*, and the solution must be analyzed periodically. In one such analysis, 45.7 mL of 0.500 *M* H_2SO_4 is required to neutralize 20.0 mL of NaOH solution. What is the concentration of the NaOH solution?

Solution

Analyze We are given the volume (45.7 mL) and molarity (0.500 *M*) of an H_2SO_4 solution (the standard solution) that reacts completely with 20.0 mL of NaOH solution. We are asked to calculate the molarity of the NaOH solution.

Plan Following the steps of Figure 4.20, we use the H_2SO_4 volume and molarity to calculate the number of moles of H_2SO_4. Then we can use this quantity and the balanced equation for the reaction to calculate moles of NaOH. Finally, we can use moles of NaOH and the NaOH volume to calculate NaOH molarity.

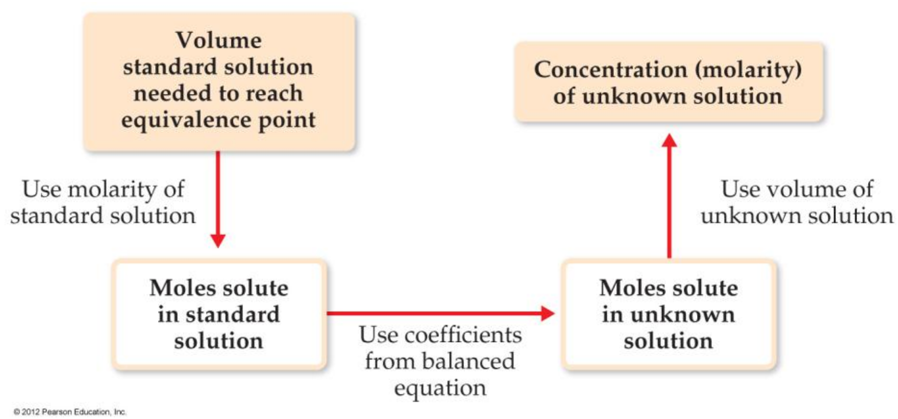

FIGURE 4.20 Procedure for determining the concentration of a solution from titration with a standard solution.

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Sample Exercise 4.16 Determining Solution Concentration by an Acid–Base Titration

Continued

Solve The number of moles of H_2SO_4 is the product of the volume and molarity of this solution:

$$\text{Moles } H_2SO_4 = (45.7 \cancel{\text{mL soln}})\left(\frac{1 \cancel{\text{L soln}}}{1000 \cancel{\text{mL soln}}}\right)\left(\frac{0.500 \text{ mol } H_2SO_4}{\cancel{\text{L soln}}}\right)$$

$$= 2.28 \times 10^{-2} \text{ mol } H_2SO_4$$

Acids react with metal hydroxides to form water and a salt. Thus, the balanced equation for the neutralization reaction is

$$H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow 2\,H_2O(l) + Na_2SO_4(aq)$$

According to the balanced equation, $1 \text{ mol } H_2SO_4 \simeq 2 \text{ mol NaOH}$. Therefore,

$$\text{Moles NaOH} = (2.28 \times 10^{-2} \cancel{\text{mol } H_2SO_4})\left(\frac{2 \text{ mol NaOH}}{\cancel{1 \text{ mol } H_2SO_4}}\right)$$

$$= 4.56 \times 10^{-2} \text{ mol NaOH}$$

Knowing the number of moles of NaOH in 20.0 mL of solution allows us to calculate the molarity of this solution:

$$\text{Molarity NaOH} = \frac{\text{mol NaOH}}{\text{L soln}} = \left(\frac{4.56 \times 10^{-2} \text{ mol NaOH}}{20.0 \cancel{\text{mL soln}}}\right)\left(\frac{1000 \cancel{\text{mL soln}}}{1 \text{ L soln}}\right)$$

$$= 2.28 \frac{\text{mol NaOH}}{\text{L soln}} = 2.28\,M$$

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Sample Exercise 4.16 Determining Solution Concentration by an Acid–Base Titration

Continued

Practice Exercise

What is the molarity of an NaOH solution if 48.0 mL neutralizes 35.0 mL of 0.144 M H_2SO_4?

Answer: 0.210 M

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Sample Exercise 4.17 Determining the Quantity of Solute by Titration

The quantity of Cl^- in a municipal water supply is determined by titrating the sample with Ag^+. The precipitation reaction taking place during the titration is

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

The end point in this type of titration is marked by a change in color of a special type of indicator. **(a)** How many grams of chloride ion are in a sample of the water if 20.2 mL of 0.100 *M* Ag^+ is needed to react with all the chloride in the sample? **(b)** If the sample has a mass of 10.0 g, what percent Cl^- does it contain?

Solution

Analyze We are given the volume (20.2 mL) and molarity (0.100 *M*) of a solution of Ag^+ and the chemical equation for reaction of this ion with Cl–.We are asked to calculate the number of grams of Cl^- in the sample and the mass percent of Cl^- in the sample.

(a) Plan We can use the procedure outlined by the green arrows in Figure 4.18.We begin by using the volume and molarity of Ag^+ to calculate the number of moles of Ag^+ used in the titration. We then use the balanced equation to determine the moles of Cl^- in the sample and from that the grams of Cl^-.

Solve

$$\text{Moles } Ag^+ = (20.2\ \cancel{\text{mL soln}})\left(\frac{1\ \cancel{\text{L soln}}}{1000\ \cancel{\text{mL soln}}}\right)\left(\frac{0.100\ \text{mol } Ag^+}{\cancel{\text{L soln}}}\right)$$

$$= 2.02 \times 10^{-3}\ \text{mol } Ag^+$$

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Sample Exercise 4.17 Determining the Quantity of Solute by Titration

Continued

From the balanced equation we see that 1 mol $Ag^+ \simeq$ 1 mol Cl^-. Using this information and the molar mass of Cl, we have

$$\text{Grams } Cl^- = (2.02 \times 10^{-3} \cancel{\text{mol Ag}^+})\left(\frac{1 \cancel{\text{mol Cl}^-}}{1 \cancel{\text{mol Ag}^+}}\right)\left(\frac{35.5 \text{ g Cl}^-}{\cancel{\text{mol Cl}^-}}\right)$$

$$= 7.17 \times 10^{-2} \text{ g Cl}^-$$

(b) Plan To calculate the percentage of Cl^- in the sample, we compare the number of grams of Cl^- in the sample, 7.17×10^{-2} g, with the original mass of the sample, 10.0 g.

Solve

$$\text{Percent } Cl^- = \frac{7.17 \times 10^{-2} \text{ g}}{10.0 \text{ g}} \times 100\% = 0.717\% \text{ Cl}^-$$

Comment Chloride ion is one of the most common ions in water and sewage. Ocean water contains 1.92% Cl^-. Whether water containing Cl^- tastes salty depends on the other ions present. If the only accompanying ions are Na^+, a salty taste may be detected with as little as 0.03% Cl–.

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Sample Exercise 4.17 Determining the Quantity of Solute by Titration

Continued

Practice Exercise

A sample of an iron ore is dissolved in acid, and the iron is converted to Fe^{2+}. The sample is then titrated with 47.20 mL of 0.02240 *M* MnO_4^- solution. The oxidation-reduction reaction that occurs during titration is

$$MnO_4^-(aq) + 5\,Fe^{2+}(aq) + 8\,H^+(aq) \qquad\qquad Mn^{2+}(aq) + 5\,Fe^{3+}(aq) + 4\,H_2O(l)$$

(a) How many moles of MnO_4^- were added to the solution? **(b)** How many moles of Fe^{2+} were in the sample? **(c)** How many grams of iron were in the sample? **(d)** If the sample had a mass of 0.8890 g, what is the percentage of iron in the sample?

Answer: **(a)** 1.057×10^{-3} mol MnO_4^-, **(b)** 5.286×10^{-3} mol Fe^{2+}, **(c)** 0.2952 g, **(d)** 33.21%

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