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

The diagram on the right represents an aqueous solution of one of the following compounds: MgCl₂, KCl, or K₂SO₄. 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 the ionic substances given in the problem to determine the relative numbers and charges of the ions that each contains. We then correlate these charged ionic species with the ones shown in the diagram.

Solve: The diagram shows twice as many cations as anions, consistent with the formulation K₂SO₄.

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

Practice Exercise

If you were to draw diagrams (such as that shown on the right) representing aqueous solutions of each of the following ionic compounds, how many anions would you show if the diagram contained six cations?

(a) NiSO₄, (b) Ca(NO₃)₂, (c) Na₃PO₄, (d) Al₂(SO₄)₃

Answers: (a) 6, (b) 12, (c) 2, (d) 9
Sample Exercise 4.2 Using Solubility Rules

Classify the following ionic compounds as soluble or insoluble in water: (a) sodium carbonate (Na₂CO₃), (b) lead sulfate (PbSO₄).

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.

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, Na₂CO₃ is soluble in water.
(b) Table 4.1 indicates that although most sulfates are water soluble, the sulfate of Pb²⁺ is an exception. Thus, PbSO₄ is insoluble in water.

Practice Exercise

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

Answers: (a) insoluble, (b) soluble, (c) soluble
Sample Exercise 4.3 Predicting a Metathesis Reaction

(a) Predict the identity of the precipitate that forms when solutions of BaCl₂ and K₂SO₄ 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 down the ions present in the reactants and to 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. Products also allows us to write the equation for the reaction.

Solve:

(a) The reactants contain Ba²⁺, Cl⁻, K⁺, and SO₄²⁻ ions. If we exchange the anions, we will have BaSO₄ and KCl. According to Table 4.1, most compounds of SO₄²⁻ are soluble but those of Ba²⁺ are not. Thus, BaSO₄ is insoluble and will precipitate from solution. KCl, on the other hand, is soluble.

(b) From part (a) we know the chemical formulas of the products, BaSO₄ and KCl. The balanced equation with phase labels shown is

\[ \text{BaCl}_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2 \text{KCl}(aq) \]

Practice Exercise

(a) What compound precipitates when solutions of Fe₂(SO₄)₃ and LiOH are mixed?
(b) Write a balanced equation for the reaction. (c) Will a precipitate form when solutions of Ba(NO₃)₂ and KOH are mixed?

Answers: (a) Fe(OH)₃; (b) Fe₂(SO₄)₃(aq) + 6 LiOH(aq) → 2 Fe(OH)₃(s) + 3 Li₂SO₄(aq); (c) no (both possible products are water soluble)
Sample Exercise 4.4 Writing a Net Ionic Equation

Write the net ionic equation for the precipitation reaction that occurs when 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 first need to 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, \( \text{Ca}^{2+} \), and chloride ions, \( \text{Cl}^- \); hence an aqueous solution of the substance is \( \text{CaCl}_2(aq) \). Sodium carbonate is composed of \( \text{Na}^+ \) ions and \( \text{CO}_3^{2-} \) ions; hence an aqueous solution of the compound is \( \text{Na}_2\text{CO}_3(aq) \). In the molecular equations for precipitation reactions, the anions and cations appear to exchange partners. Thus, we put \( \text{Ca}^{2+} \) and \( \text{CO}_3^{2-} \) together to give \( \text{CaCO}_3 \) and \( \text{Na}^+ \) and \( \text{Cl}^- \) together to give \( \text{NaCl} \). According to the solubility guidelines in Table 4.1, \( \text{CaCO}_3 \) is insoluble and \( \text{NaCl} \) is soluble. The balanced molecular equation is

\[
\text{CaCl}_2(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow \text{CaCO}_3(s) + 2 \text{NaCl}(aq)
\]
Sample Exercise 4.4 Writing a Net Ionic Equation

Solution (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\(_2\)CO\(_3\), 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

\[
\text{Ca}^{2+}(aq) + 2 \text{Cl}^-(aq) + 2 \text{Na}^+(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{CaCO}_3(s) + 2 \text{Na}^+(aq) + 2 \text{Cl}^-(aq)
\]

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

\[
\text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{CaCO}_3(s)
\]

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, then they all 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.

**Answers:** \(3 \text{Ag}^+(aq) + \text{PO}_4^{3-}(aq) \rightarrow \text{Ag}_3\text{PO}_4(s)\)
Sample Exercise 4.5 Comparing Acid Strengths

The following diagrams represent aqueous solutions of three acids (HX, HY, and HZ) with water molecules omitted for clarity. Rank them 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 examine the drawings to determine the relative numbers of uncharged molecular species present. The strongest acid is the one with the most H⁺ ions and fewest undissociated acid 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 H⁺ ions and fewer molecules than HX, it is a stronger acid.
Sample Exercise 4.5 Writing a Net Ionic Equation

Practice Exercise

Imagine a diagram showing 10 Na\(^+\) ions and 10 OH\(^-\) ions. If this solution were mixed with the one pictured on the previous slide for HY, what would the diagram look like that represents the solution after any possible reaction? (H\(^+\) ions will react with ions to form H\(_2\)O.)

**Answer:** The final diagram would show 10 Na\(^+\) ions, 2 OH\(^-\) ions, 8 Y\(^-\) ions, and 8 H\(_2\)O molecules.
Sample Exercise 4.6 Identifying Strong, Weak, and Nonelectrolytes

Classify each of the following dissolved substances as a strong electrolyte, weak electrolyte, or nonelectrolyte: CaCl₂, HNO₃, C₂H₅OH (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.

Solve: Two compounds fit the criteria for ionic compounds: CaCl₂ 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, HNO₃ and HCOOH, are acids. Nitric acid, HNO₃, 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, C₂H₅OH, is neither an acid nor a base, so it is a nonelectrolyte.

Comment: Although C₂H₅OH has an OH group, it is not a metal hydroxide; thus, it is 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). The COOH group is called the “carboxylic acid group” (Chapter 16). Molecules that have this group are all weak acids.
Sample Exercise 4.6 Predicting a Metathesis Reaction

Practice Exercise

Consider solutions in which 0.1 mol of each of the following compounds is dissolved in 1 L of water: Ca(NO₃)₂ (calcium nitrate), C₆H₁₂O₆ (glucose), CH₃COONa (sodium acetate), and CH₃COOH (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.

Answers: C₆H₁₂O₆ (nonelectrolyte) < CH₃COOH (weak electrolyte, existing mainly in the form of molecules with few ions) < CH₃COONa (strong electrolyte that provides two ions, Na⁺ and CH₃COO⁻) < Ca(NO₃)₂ (strong electrolyte that provides three ions, Ca²⁺ and 2 NO₃⁻)
Sample Exercise 4.7 Writing Chemical Equations for a Neutralization Reaction

(a) Write a balanced molecular equation for the reaction between aqueous solutions of acetic acid (CH₃COOH) and barium hydroxide, Ba(OH)₂. (b) Write the net ionic equation for this reaction.

Solution

Analyze: We are given the chemical formulas for an acid and a base and asked to write a balanced molecular equation and then 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, H₂O 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 will contain the cation of the base (Ba²⁺) and the anion of the acid (CH₃COO⁻). Thus, the formula of the salt is Ba(CH₃COO)₂. According to the solubility guidelines in Table 4.1, this compound is soluble. The unbalanced molecular equation for the neutralization reaction is

\[
\text{CH}_3\text{COOH}(aq) + \text{Ba(OH)}_2(aq) \rightarrow \text{H}_2\text{O}(l) + \text{Ba(CH}_3\text{COO)}_2(aq)
\]

To balance this molecular equation, we must provide two molecules of CH₃COOH to furnish the two CH₃COO⁻ ions and to supply the two H⁺ ions needed to combine with the two CH⁻ ions of the base. The balanced molecular equation is

\[
2\ \text{CH}_3\text{COOH}(aq) + \text{Ba(OH)}_2(aq) \rightarrow 2\ \text{H}_2\text{O}(l) + \text{Ba(CH}_3\text{COO)}_2(aq)
\]
Sample Exercise 4.7 Writing Chemical Equations for a Neutralization Reaction

Solution (continued)

(b) To write the net ionic equation, we must determine whether each compound in aqueous solution is a strong electrolyte. CH₃COOH is a weak electrolyte (weak acid), Ba(OH)₂ is a strong electrolyte, and Ba(CH₃COO)₂ is also a strong electrolyte (ionic compound). Thus, the complete ionic equation is

\[
2 \text{CH}_3\text{COOH}(aq) + \text{Ba}^{2+}(aq) + 2 \text{OH}^-(aq) \rightarrow \text{2H}_2\text{O}(l) + \text{Ba}^{2+}(aq) + 2 \text{CH}_3\text{COO}^-(aq)
\]

Eliminating the spectator ions gives

\[
2 \text{CH}_3\text{COOH}(aq) + 2 \text{OH}^-(aq) \rightarrow 2 \text{H}_2\text{O}(l) + 2 \text{CH}_3\text{COO}^-(aq)
\]

Simplifying the coefficients gives the net ionic equation:

\[
\text{CH}_3\text{COOH}(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CH}_3\text{COO}^-(aq)
\]

Check: We can determine whether the molecular equation is correctly balanced by counting the number of atoms of each kind on both sides of the arrow. (There are 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 CH₃COO 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.
Sample Exercise 4.7 Writing Chemical Equations for a Neutralization Reaction

Practice Exercise

(a) Write a balanced molecular equation for the reaction of carbonic acid (H₂CO₃) and potassium hydroxide (KOH). (b) Write the net ionic equation for this reaction.

Answers: (a) H₂CO₃(aq) + 2 KOH(aq) → 2 H₂O(l) + K₂CO₃(aq);
(b) H₂CO₃(aq) + 2 OH⁻(aq) → 2 H₂O(l) + CO₃²⁻(aq). (H₂CO₃ is a weak acid and therefore a weak electrolyte, whereas KOH, a strong base, and K₂CO₃, an ionic compound, are strong electrolytes.)
Sample Exercise 4.8 Determining Oxidation Numbers

Determine the oxidation number of sulfur in each of the following: (a) $\text{H}_2\text{S}$, (b) $\text{S}_8$, (c) $\text{SCl}_2$, (d) $\text{Na}_2\text{SO}_3$, (e) $\text{SO}_4^{2-}$.

Solution

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

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 above to assign oxidation numbers.

Solve:

(a) When bonded to a nonmetal, hydrogen has an oxidation number of +1 (rule 3b). Because the $\text{H}_2\text{S}$ 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 $\text{SO}_4^{2-}$ 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.

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.
Sample Exercise 4.8 Determining Oxidation Numbers

Practice Exercise

What is the oxidation state of the boldfaced element in each of the following: (a) $\text{P}_2\text{O}_5$, (b) $\text{NaH}$, (c) $\text{Cr}_2\text{O}_7^{2-}$, (d) $\text{SnBr}_4$, (e) $\text{BaO}_2$?

*Answers:* (a) +5, (b) −1, (c) +6, (d) +4, (e) −1
Sample Exercise 4.9 Writing Molecular and Net Ionic 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 H₂ gas. To write the balanced equations, we must write the chemical formulas for the two reactants and then determine the formula of the salt. The salt is composed of the cation formed by the metal and the anion of the acid.

**Solve:** The formulas of the given reactants are Al and HBr. The cation formed by Al is Al³⁺, and the anion from hydrobromic acid is Br⁻. Thus, the salt formed in the reaction is AlBr₃. Writing the reactants and products and then balancing the equation gives this molecular equation:

\[
2 \text{Al}(s) + 6 \text{HBr}(aq) \rightarrow 2 \text{AlBr}_3(aq) + 3 \text{H}_2(g)
\]

Both HBr and AlBr₃ are soluble strong electrolytes. Thus, the complete ionic equation is

\[
2 \text{Al}(s) + 6 \text{H}^+(aq) + 6 \text{Br}^-(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 6 \text{Br}^-(aq) + 3 \text{H}_2(g)
\]

Because Br⁻ is a spectator ion, the net ionic equation is

\[
2 \text{Al}(s) + 6 \text{H}^+(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{H}_2(g)
\]
Sample Exercise 4.9  Writing Molecular and Net Ionic Equations for Oxidation-Reduction Reactions

Solution (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?

Answers: (a) Mg\((s)\) + CoSO\(_{4}(aq)\) → MgSO\(_{4}(aq)\) + Co\((s)\); Mg\((s)\) + Co\(^{2+}(aq)\) → Mg\(^{2+}(aq)\) + Co\((s)\) (b) Mg is oxidized and Co\(^{2+}\) is reduced.
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, FeCl₂, and a metal, Mg—and asked if they react with each other.

Plan: A reaction will occur if Mg is above Fe in the activity series, Table 4.5. If the reaction occurs, the Fe²⁺ ion in FeCl₂ will be reduced to Fe, and the elemental Mg will be oxidized to Mg²⁺.

Solve: Because Mg is above Fe in the table, the reaction will occur. To write the formula for the salt that is produced in the reaction, we must remember the charges on common ions. Magnesium is always present in compounds as Mg²⁺: the chloride ion is Cl⁻. The magnesium salt formed in the reaction is MgCl₂, meaning the balanced molecular equation is

\[ \text{Mg}(s) + \text{FeCl}_2(aq) \rightarrow \text{MgCl}_2(aq) + \text{Fe}(s) \]

Both FeCl₂ and MgCl₂ are soluble strong electrolytes and can be written in ionic form. Cl⁻ then, is a spectator ion in the reaction. The net ionic equation is

\[ \text{Mg}(s) + \text{Fe}^{2+}(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{Fe}(s) \]

The net ionic equation shows that Mg is oxidized and Fe²⁺ is reduced in this reaction.

Check: Note that the net ionic equation is balanced with respect to both charge and mass.
Sample Exercise 4.10 Determining When an Oxidation-Reduction Reaction Can Occur

Practice Exercise
Which of the following metals will be oxidized by Pb(NO₃)₂: Zn, Cu, Fe?

*Answer:* Zn and Fe
Sample Exercise 4.11 Calculating Molarity

Calculate the molarity of a solution made by dissolving 23.4 g of sodium sulfate (Na₂SO₄) 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 (Na₂SO₄), and the volume of the solution (125 ml). We are asked to calculate the molarity of the solution.

Plan: We can calculate molarity using Equation 4.33. 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 Na₂SO₄ is obtained by using its molar mass:

\[
\text{Moles Na}_2\text{SO}_4 = \left(\frac{23.4 \text{ g Na}_2\text{SO}_4}{142 \text{ g Na}_2\text{SO}_4}\right) \left(\frac{1 \text{ mol Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4}\right) = 0.165 \text{ mol Na}_2\text{SO}_4
\]

Converting the volume of the solution to liters:

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

Thus, the molarity is

\[
\text{Molarity} = \frac{0.165 \text{ mol Na}_2\text{SO}_4}{0.125 \text{ L soln}} = 1.32 \frac{\text{ mol Na}_2\text{SO}_4}{L \text{ soln}} = 1.32 \text{ M}
\]

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 (mol/L) 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.
Sample Exercise 4.11 Calculating Molarity

Practice Exercise

Calculate the molarity of a solution made by dissolving 5.00 g of glucose (C\textsubscript{6}H\textsubscript{12}O\textsubscript{6}) in sufficient water to form exactly 100 mL of solution.

Answer: 0.278 M
Sample Exercise 4.12 Calculating Molar Concentrations of Ions

What are the molar concentrations of each of the ions 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 concentrations of the ions.

Solve: Calcium nitrate is composed of calcium (Ca\(^{2+}\)) ions 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 in the compound, 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 \(2 \times 0.025 \text{ M} = 0.050 \text{ M}\) in NO\(^{3-}\):

\[
\frac{\text{mol NO}_3^-}{\text{L}} = \left( \frac{0.025 \text{ mol Ca(NO}_3)_2}{\text{L}} \right) \left( \frac{2 \text{ mol NO}_3^-}{1 \text{ mol Ca(NO}_3)_2} \right) = 0.050 \text{ M}
\]

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 K\(^+\)
Sample Exercise 4.13 Using Molarity to Calculate Grams of Solute

How many grams of Na₂SO₄ are required to make 0.350 L of 0.500 M Na₂SO₄?

Solution

Analyze: We are given the volume of the solution (0.350 L), its concentration (0.500 M), and the identity of the solute Na₂SO₄ and asked to calculate the number of grams of the solute in the solution.

Plan: We can use the definition of molarity (Equation 4.33) to determine the number of moles of solute, and then convert moles to grams using the molar mass of the solute.

\[ M_{\text{Na}_2\text{SO}_4} = \frac{\text{moles Na}_2\text{SO}_4}{\text{liters soln}} \]

Solve: Calculating the moles of Na₂SO₄ using the molarity and volume of solution gives

\[ M_{\text{Na}_2\text{SO}_4} = \frac{\text{moles Na}_2\text{SO}_4}{\text{liters soln}} \]

\[ = \frac{\text{liters soln} \times M_{\text{Na}_2\text{SO}_4}}{} \]

\[ = (0.350 \text{ L soln}) \left( \frac{0.500 \text{ mol Na}_2\text{SO}_4}{1 \text{ L soln}} \right) \]

\[ = 0.175 \text{ mol Na}_2\text{SO}_4 \]

Because each mole of Na₂SO₄ weighs 142 g, the required number of grams of Na₂SO₄ is

\[ \text{grams Na}_2\text{SO}_4 = (0.175 \text{ mol Na}_2\text{SO}_4) \left( \frac{142 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} \right) = 24.9 \text{ g Na}_2\text{SO}_4 \]

Check: The magnitude of the answer, the units, and the number of significant figures are all appropriate.
Sample Exercise 4.13 Using Molarity to Calculate Grams of Solute

Practice Exercise

(a) How many grams of Na₂SO₄ are there in 15 mL of 0.50 M Na₂SO₄? (b) How many milliliters of 0.50 M Na₂SO₄ solution are needed to provide 0.038 mol of this salt?

Answers: (a) 1.1 g, (b) 76 mL
Sample Exercise 4.14 Preparing A solution by Dilution

How many milliliters of 3.0 \( M \) \( \text{H}_2\text{SO}_4 \) are needed to make 450 mL of 0.10 \( M \) \( \text{H}_2\text{SO}_4 \)?

**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, \( \text{H}_2\text{SO}_4 \), 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 4.35. Let’s compare the two methods.

**Solve:** Calculating the moles of \( \text{H}_2\text{SO}_4 \) in the dilute solution:

\[
\text{moles } \text{H}_2\text{SO}_4 \text{ in dilute solution} = (0.450 \text{ L soln}) \left( \frac{0.10 \text{ mol } \text{H}_2\text{SO}_4}{1 \text{ L soln}} \right) = 0.045 \text{ mol } \text{H}_2\text{SO}_4
\]

Calculating the volume of the concentrated solution that contains 0.045 mol \( \text{H}_2\text{SO}_4 \):

\[
\text{L conc soln} = (0.045 \text{ mol } \text{H}_2\text{SO}_4) \left( \frac{1 \text{ L soln}}{3.0 \text{ mol } \text{H}_2\text{SO}_4} \right) = 0.015 \text{ L soln}
\]

Converting liters to milliliters gives 15 mL.

If we apply Equation 4.35, we get the same result:

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

\[
V_{\text{conc}} = \frac{(0.10 \text{ M})(450 \text{ mL})}{3.0 \text{ M}} = 15 \text{ mL}
\]
Sample Exercise 4.14 Preparing A solution by Dilution

Solution (continued)

Either way, we see that if we start with 15 mL of 3.0 M H$_2$SO$_4$ 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.

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$_2$Cr$_2$O$_7$ 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?

Answers: (a) 0.0200 L = 20.0 mL, (b) 5.0 mL, (c) 0.40 M
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:** We can use the molarity and volume of the HNO$_3$ solution to calculate the number of moles of HNO$_3$. We then use the balanced equation to relate the moles of HNO$_3$ to moles of Ca(OH)$_2$. Finally, we can convert moles of Ca(OH)$_2$ to grams. These steps can be summarized as follows:

\[
L_{\text{HNO}_3} \times M_{\text{HNO}_3} \Rightarrow \text{mol HNO}_3 \Rightarrow \text{mol Ca(OH)}_2 \Rightarrow \text{g Ca(OH)}_2
\]

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

\[
\text{moles HNO}_3 = L_{\text{HNO}_3} \times M_{\text{HNO}_3} = (0.0250 \, \text{L})(0.100 \, \frac{\text{mol HNO}_3}{\text{L}}) \\
= 2.50 \times 10^{-3} \, \text{mol HNO}_3
\]

Because this is an acid–base neutralization reaction, HNO$_3$ and Ca(OH)$_2$ react to form H$_2$O and the salt containing Ca$^{2+}$ and NO$_3^-$:

\[
2 \, \text{HNO}_3(aq) + \text{Ca(OH)}_2(s) \rightarrow 2 \, \text{H}_2\text{O}(l) + \text{Ca(NO}_3)_2(aq)
\]

Thus, 2 mol HNO$_3$ \( \equiv \) 1 mol Ca(OH)$_2$. Therefore,

\[
\text{grams Ca(OH)}_2 = (2.50 \times 10^{-3} \, \text{mol HNO}_3)(\frac{1 \, \text{mol Ca(OH)}_2}{2 \, \text{mol HNO}_3})(\frac{74.1 \, \text{g Ca(OH)}_2}{1 \, \text{mol Ca(OH)}_2}) \\
= 0.0926 \, \text{g Ca(OH)}_2
\]
Sample Exercise 4.15 Using Mass Relations In a Neutralization Reaction

Solution

Check: The size of the answer is reasonable. A small volume of dilute acid will require only a small amount of base to neutralize it.

Practice Exercise

(a) How many grams of NaOH are needed to neutralize 20.0 mL of 0.150 M H₂SO₄ solution? (b) How many liters of 0.500 M HCl(aq) are needed to react completely with 0.100 mol of Pb(NO₃)₂(aq), forming a precipitate of PbCl₂(s)?

Answers: (a) 0.240 g, (b) 0.400 L
Sample Exercise 4.16 Determining the Quality of Solute by Titration

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

\[
\text{Ag}^{+}(aq) + \text{Cl}^{-}(aq) \rightarrow \text{AgCl(s)}
\]

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 first to calculate the number of grams of Cl\(^{-}\) in the sample and, second, to calculate the mass percent of Cl\(^{-}\) in the sample.

(a) Plan: We begin by using the volume and molarity of Ag\(^{+}\) to calculate the number of moles of Ag\(^{+}\) used in the titration. We can then use the balanced equation to determine the moles of Cl\(^{-}\) in the sample and from that the grams of Cl\(^{-}\).

Solve:

\[
\text{moles Ag}^{+} = (20.2 \text{ mL soln}) \left( \frac{1 \text{ L soln}}{1000 \text{ mL soln}} \right) \left( 0.100 \frac{\text{mol Ag}^{+}}{\text{L soln}} \right) = 2.02 \times 10^{-3} \text{ mol Ag}^{+}
\]

From the balanced equation we see that formula. Using this information and the molar mass of Cl\(^{-}\), we have

\[
\text{grams Cl}^{-} = (2.02 \times 10^{-3} \text{ mol Ag}^{+}) \left( \frac{1 \text{ mol Cl}^{-}}{1 \text{ mol Ag}^{+}} \right) \left( 35.5 \text{ g Cl}^{-} \right) = 7.17 \times 10^{-2} \text{ g Cl}^{-}
\]
Sample Exercise 4.16 Determining the Quality of Solute by Titration

Solution (continued)

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

\[
\text{Solve: Percent Cl}^{-} = \frac{7.17 \times 10^{-3} \text{ g}}{10.0 \text{ g}} \times 100\% = 0.717\% \text{ Cl}^{-}
\]

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\(^-\).

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 as follows:

\[
\text{MnO}_4^-(aq) + 5 \text{Fe}^{2+}(aq) + 8 \text{H}^+(aq) \rightarrow \text{Mn}^{2+}(aq) + 5 \text{Fe}^{3+}(aq) + 4 \text{H}_2\text{O}(l)
\]

(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?

Answers: (a) 1.057 \times 10^{-3} \text{ mol MnO}_4^- \ (b) 5.286 \times 10^{-3} \text{ mol Fe}^{2+}, \ (c) 0.2952 \text{ g}, \ (d) 33.21\%
Sample Exercise 4.17 Determining Solution Concentration Via an Acid-Base Titration

One commercial method used to peel potatoes is to soak them in a solution of NaOH for a short time, remove them from the NaOH, and spray off the peel. The concentration of NaOH is normally in the range of 3 to 6 M. The NaOH is analyzed periodically. In one such analysis, 45.7 mL of 0.500 M H_2SO_4 is required to neutralize a 20.0-mL sample 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 that reacts completely with a 20.0-mL sample of NaOH. We are asked to calculate the molarity of the NaOH solution.

Plan: We can use the volume and molarity of the H_2SO_4 to calculate the number of moles of this substance. Then, we can use this quantity and the balanced equation for the reaction to calculate the number of moles of NaOH. Finally, we can use the moles of NaOH and the volume of this solution to calculate molarity.

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

\[
\text{moles H}_2\text{SO}_4 = (45.7 \text{ mL soln}) \left( \frac{1 \text{ L soln}}{1000 \text{ mL soln}} \right) \left( 0.500 \frac{\text{mol H}_2\text{SO}_4}{\text{L soln}} \right)
\]

\[
= 2.28 \times 10^{-2} \text{ mol H}_2\text{SO}_4
\]

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

\[
\text{H}_2\text{SO}_4(aq) + 2 \text{NaOH}(aq) \rightarrow 2 \text{H}_2\text{O}(l) + \text{Na}_2\text{SO}_4(aq)
\]

According to the balanced equation, 1 mol H_2SO_4 \approx 2 mol NaOH. Therefore,

\[
\text{moles NaOH} = (2.28 \times 10^{-2} \text{ mol H}_2\text{SO}_4) \left( 2 \frac{\text{mol NaOH}}{1 \text{mol H}_2\text{SO}_4} \right)
\]

\[
= 4.56 \times 10^{-2} \text{ mol NaOH}
\]
Sample Exercise 4.17 Determining Solution Concentration Via an Acid-Base Titration

Solution (continued)

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

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

\[
= 2.28 \frac{\text{mol NaOH}}{\text{L soln}} = 2.28 \text{ M}
\]

Practice Exercise

What is the molarity of an NaOH solution if 48.0 mL is needed to neutralize 35.0 mL of 0.144 M H\textsubscript{2}SO\textsubscript{4}?

Answers: 0.210 M
Sample Integrative Exercise

Note: Integrative exercises require skills from earlier chapters as well as ones from the present chapter. A sample of 70.5 mg of potassium phosphate is added to 15.0 mL of 0.050 M silver nitrate, resulting in the formation of a precipitate. (a) Write the molecular equation for the reaction. (b) What is the limiting reactant in the reaction? (c) Calculate the theoretical yield, in grams, of the precipitate that forms.

Solution

(a) Potassium phosphate and silver nitrate are both ionic compounds. Potassium phosphate contains K\(^+\) and PO\(_4^{3-}\) ions, so its chemical formula is K\(_3\)PO\(_4\). Silver nitrate contains Ag\(^+\) and NO\(_3^-\) ions, so its chemical formula is AgNO\(_3\). Because both reactants are strong electrolytes, the solution contains K\(^+\), PO\(_4^{3-}\), Ag\(^+\), and NO\(_3^-\) ions before the reaction occurs. According to the solubility guidelines in Table 4.1, and PO\(_4^{3-}\) form an insoluble compound, so Ag\(_3\)PO\(_4\) will precipitate from the solution. In contrast, K\(^+\) and NO\(_3^-\) will remain in solution because KNO\(_3\) is water soluble. Thus, the balanced molecular equation for the reaction is

\[
K_3PO_4(aq) + 3 \text{AgNO}_3(aq) \rightarrow \text{Ag}_3\text{PO}_4(s) + 3 \text{KNO}_3(aq)
\]

(b) To determine the limiting reactant, we must examine the number of moles of each reactant. (Section 3.7) The number of moles of K\(_3\)PO\(_4\) is calculated from the mass of the sample using the molar mass as a conversion factor. (Section 3.4) The molar mass of K\(_3\)PO\(_4\) is 3(39.1) + 31.0 + 4(16.0) = 212.3 g/mol. Converting milligrams to grams and then to moles, we have

\[
(70.5 \text{ mg K}_3\text{PO}_4) \left(\frac{10^{-3} \text{ g K}_3\text{PO}_4}{1 \text{ mg K}_3\text{PO}_4}\right) \left(\frac{1 \text{ mol K}_3\text{PO}_4}{212.3 \text{ g K}_3\text{PO}_4}\right) = 3.32 \times 10^{-4} \text{ mol K}_3\text{PO}_4
\]
Sample Integrative Exercise

Solution (continued)

We determine the number of moles of AgNO₃ from the volume and molarity of the solution. (Section 4.5)
Converting milliliters to liters and then to moles, we have

\[
15.0 \text{ mL} \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left( \frac{0.050 \text{ mol AgNO₃}}{1 \text{ L}} \right) = 7.5 \times 10^{-4} \text{ mol AgNO₃}
\]

Comparing the amounts of the two reactants, we find that there are \((7.5 \times 10^{-4})/(3.32 \times 10^{-4}) = 2.3\) times as many moles of AgNO₃ as there are moles of K₃PO₄. According to the balanced equation, however, 1 mol K₃PO₄ requires 3 mol of AgNO₃. Thus, there is insufficient AgNO₃ to consume the K₃PO₄, and AgNO₃ is the limiting reactant.

(c) The precipitate is Ag₃PO₄, whose molar mass is 3(107.9) + 31.0 + 4(16.0) = 418.7 g/mol. To calculate the number of grams of Ag₃PO₄ that could be produced in this reaction (the theoretical yield), we use the number of moles of the limiting reactant, converting mol AgNO₃ \(\Rightarrow\) mol Ag₃PO₄ \(\Rightarrow\) g Ag₃PO₄.

We use the coefficients in the balanced equation to convert moles of AgNO₃ to moles Ag₃PO₄, and we use the molar mass of Ag₃PO₄ to convert the number of moles of this substance to grams.

\[
(7.5 \times 10^{-4} \text{ mol AgNO₃}) \left( \frac{1 \text{ mol Ag₃PO₄}}{3 \text{ mol AgNO₃}} \right) \left( \frac{418.7 \text{ g Ag₃PO₄}}{1 \text{ mol Ag₃PO₄}} \right) = 0.10 \text{ g Ag₃PO₄}
\]

The answer has only two significant figures because the quantity of AgNO₃ is given to only two significant figures.