4. CHEMICAL REACTIONS

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

4.1 a. According to Table 4.1, all compounds that contain sodium, Na^+, are soluble. Thus, NaBr is soluble in water.

b. According to Table 4.1, most compounds that contain hydroxides, OH^-, are insoluble in water. However, $Ba(OH)_2$ is listed as one of the exceptions to this rule, so it is soluble in water.

c. Calcium carbonate is $CaCO_3$. According to Table 4.1, most compounds that contain carbonate, CO_3^{2-}, are insoluble. $CaCO_3$ is not one of the exceptions, so it is insoluble in water.

4.2 a. The problem states that HNO_3 is a strong electrolyte, but $Mg(OH)_2$ is a solid, so retain its formula. On the product side, $Mg(NO_3)_2$ is a soluble ionic compound, but water is a nonelectrolyte, so retain its formula. The resulting complete ionic equation is

$$2H^+(aq) + 2NO_3^-(aq) + Mg(OH)_2(s) \rightarrow 2H_2O(l) + Mg^{2+}(aq) + 2NO_3^-(aq)$$

The corresponding net ionic equation is

$$2H^+(aq) + Mg(OH)_2(s) \rightarrow 2H_2O(l) + Mg^{2+}(aq)$$

(continued)

b. Both reactants are soluble ionic compounds, and on the product side, $NaNO_3$ is also a soluble ionic compound. $PbSO_4$ is a solid, so retain its formula. The resulting complete ionic equation is

$$Pb^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + SO_4^{2-}(aq) \rightarrow$$
$$PbSO_4(s) + 2Na^+(aq) + 2NO_3^-(aq)$$

The corresponding net ionic equation is

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$$

4.3 The formulas of the compounds are NaI and $Pb(C_2H_3O_2)_2$. Exchanging anions, you get sodium acetate, $NaC_2H_3O_2$, and lead(II) iodide, PbI_2. The equation for the exchange reaction is

$$NaI + Pb(C_2H_3O_2)_2 \rightarrow NaC_2H_3O_2 + PbI_2$$

From Table 4.1, you see that NaI is soluble, $Pb(C_2H_3O_2)_2$ is soluble, $NaC_2H_3O_2$ is soluble, and PbI_2 is insoluble. Thus, lead(II) iodide precipitates. The balanced molecular equation with phase labels is

$$Pb(C_2H_3O_2)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaC_2H_3O_2(aq)$$

To get the net ionic equation, you write the soluble ionic compounds as ions, and cancel the spectator ions, ($C_2H_3O_2^-$ and Na^+). The final result is

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

4.4 a. H_3PO_4 is not listed as a strong acid in Table 4.3, so it is a weak acid.

b. Hypochlorous acid, HClO, is not one of the strong acids listed in Table 4.3, therefore we assume that HClO is a weak acid.

c. As noted in Table 4.3, $HClO_4$ is a strong acid.

d. As noted in Table 4.3, $Sr(OH)_2$ is a strong base.

4.5 The salt consists of the cation from the base (Li^+) and the anion from the acid (CN^-); its formula is LiCN. You will need to add H_2O as a product to complete and balance the molecular equation:

$$HCN(aq) + LiOH(aq) \rightarrow LiCN(aq) + H_2O(l)$$

Note that LiOH (a strong base) and LiCN (a soluble ionic substance) are strong electrolytes; HCN is a weak electrolyte (it is not one of the strong acids in Table 4.3). After eliminating the spectator ions (Li^+ and CN^-), the net ionic equation is

$$HCN(aq) + OH^-(aq) \rightarrow H_2O(l) + CN^-(aq)$$

4.6 The first step in the neutralization is described by the following molecular equation:

$$H_2SO_4(aq) + KOH(aq) \rightarrow KHSO_4(aq) + H_2O(l)$$

The corresponding net ionic equation is

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

The reaction of the acid salt $KHSO_4$ is given by the following molecular equation:

$$KHSO_4(aq) + KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$$

The corresponding net ionic equation is

$$HSO_4^-(aq) + OH^-(aq) \rightarrow H_2O(l) + SO_4^{2-}(aq)$$

4.7 First, write the molecular equation for the exchange reaction noting that the products of the reaction would be soluble $Ca(NO_3)_2$ and H_2CO_3. The carbonic acid decomposes to water and carbon dioxide gas. The molecular equation for the process is

$$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

The corresponding net ionic equation is

$$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$$

4.8 a. For potassium dichromate, $K_2Cr_2O_7$,

$$2 \text{ x (oxidation number of K)} + 2 \text{ x (oxidation number of Cr)}$$

$$+ 7 \text{ x (oxidation number of O)} = 0$$

For oxygen,

$$2 \text{ x } (+1) + 2 \text{ x (oxidation number of Cr)} + 7 \text{ x } (-2) = 0$$

Therefore,

$$2 \text{ x oxidation number of Cr} = -2 \text{ x } (+1) - 7 \text{ x } (-2) = +12$$

or, oxidation number of Cr = +6.

b. For the permanganate ion, MnO_4^-,

$$\text{(Oxidation number of Mn)} + 4 \text{ x (oxidation number of O)} = -1$$

For oxygen,

$$\text{(oxidation number of Mn)} + 4 \text{ x } (-2) = -1$$

Therefore,

$$\text{Oxidation number of Mn} = -1 - [4 \text{ x } (-2)] = +7$$

4.9 Identify the oxidation states of the elements.

$$\overset{0}{Ca} + \overset{0}{Cl_2} \rightarrow \overset{+2}{Ca}\overset{-1}{Cl_2}$$

Break the reaction into two half reactions making sure that both mass and charge are balanced.

$$Ca \rightarrow Ca^{2+} + 2e^-$$

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

Since each half-reaction has two electrons, it is not necessary to multiply the reactions by any factors to cancel them out. Adding the two half-reactions together and canceling out the electrons, you get

$$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$$

4.10 Convert mass of NaCl (molar mass, 58.44 g) to moles of NaCl. Then divide moles of solute by liters of solution. Note that 25.0 mL = 0.0250 L.

$$0.0678 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 1.1\underline{6}0 \times 10^{-3} \text{ mol NaCl}$$

$$\text{Molarity} = \frac{1.1\underline{6}0 \times 10^{-3} \text{ mol NaCl}}{0.0250 \text{ L soln}} = 0.046\underline{4}1 = 0.0464\text{M}$$

4.11 Convert grams of NaCl (molar mass, 58.44 g) to moles NaCl and then to volume of NaCl solution.

$$0.0958 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \frac{1 \text{ L soln}}{0.163 \text{ mol NaCl}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= 10.\underline{0}6 = 10.1 \text{ mL NaCl}$$

4.12 One (1) liter of solution is equivalent to 0.15 mol NaCl. The amount of NaCl in 50.0 mL of solution is

$$50.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.15 \text{ mol NaCl}}{1 \text{ L soln}} = 0.007\underline{5}0 \text{ mol NaCl}$$

Convert to grams using the molar mass of NaCl (58.44 g/mol).

$$0.00750 \text{ mol NaCl} \times \frac{58.4 \text{ g NaCl}}{1 \text{ mol NaCl}} = 0.4\underline{3}8 = 0.44 \text{ g NaCl}$$

4.13 Use the rearranged version of the dilution formula from the text to calculate the initial volume of 1.5 M sulfuric acid required:

$$V_i = \frac{M_f V_f}{M_i} = \frac{0.18 \text{ M} \times 100.0 \text{ mL}}{1.5 \text{ M}} = 1\underline{2}.0 = 12 \text{ mL}$$

4.14 There are two different reactions taking place in forming the CaC_2O_4 (molar mass 128.10 g/mol) precipitate. These are

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

$$CaCl_2(aq) + Na_2C_2O_4(aq) \rightarrow CaC_2O_4(s) + 2NaCl(aq)$$

The overall stoichiometry of the reactions is one mol $CaCO_3$/one mol CaC_2O_4. Also note that each $CaCO_3$ contains one Ca atom, so this gives an overall conversion factor of one mol Ca/one mol CaC_2O_4. The mass of Ca can now be calculated.

$$0.1402 \text{ g } CaC_2O_4 \times \frac{1 \text{ mol } CaC_2O_4}{128.10 \text{ g } CaC_2O_4} \times \frac{1 \text{ mol Ca}}{1 \text{ mol } CaC_2O_4} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}}$$

$$= 0.0438\underline{6}6 \text{ g Ca}$$

Now, calculate the percentage of calcium in the 128.3 mg (0.1283 g) limestone:

$$\frac{0.0438\underline{6}6 \text{ g Ca}}{0.1283 \text{ g limestone}} \times 100\% = 34.1\underline{9}0 = 34.19\%$$

4.15 Convert the volume of Na_3PO_4 to moles using the molarity of Na_3PO_4. Note that 45.7 ml = 0.0457 L.

$$0.0457 \text{ L } Na_3PO_4 \times \frac{0.265 \text{ mol } Na_3PO_4}{1 \text{ L}} = 0.012\underline{1}1 \text{ mol } Na_3PO_4$$

Finally, calculate the amount of $NiSO_4$ required to react with this amount of Na_3PO_4:

$$0.12\underline{1}1 \text{ mol } Na_3PO_4 \times \frac{3 \text{ mol } NiSO_4}{2 \text{ mol } Na_3PO_4} \times \frac{1 \text{ L } NiSO_4}{0.375 \text{ M } NiSO_4} = 0.048\underline{4}4 \text{ L (48.4 mL)}$$

4.16 Convert the volume of NaOH solution (0.0391 L) to moles NaOH (from the molarity of NaOH). Then, convert moles NaOH to moles $HC_2H_3O_2$ (from the chemical equation). Finally, convert moles of $HC_2H_3O_2$ (molar mass 60.05 g/mol) to grams $HC_2H_3O_2$.

$$0.0391 \text{ L NaOH} \times \frac{0.108 \text{ mol NaOH}}{1 \text{ L}} \times \frac{1 \text{ mol } HC_2H_3O_2}{1 \text{ mol NaOH}} \times \frac{60.05 \text{ g } HC_2H_3O_2}{1 \text{ mol } HC_2H_3O_2}$$

$$= 0.25\underline{3}59 \text{ g}$$

The mass percentage of acetic acid in the vinegar can now be calculated.

$$\text{Percentage Mass} = \frac{0.25\underline{3}59 \text{ g } HC_2H_3O_2}{5.00 \text{ g vinegar}} \times 100\% = 5.0\underline{7}1 = 5.07\%$$

■ Answers to Concept Checks

4.1 The left beaker contains two types of individual atoms (ions) and no solid; therefore, it must represent the soluble, LiI. Because LiI is a soluble ionic compound, it is an electrolyte. The beaker on the right represents a molecular compound that is soluble but not dissociated in solution. Therefore, it must be the CH_3OH. Because the CH_3OH is not dissociated in solution, and no ions are present, it is a nonelectrolyte.

4.2 a. In order to solve this part of the problem, keep in mind that this is an exchange (metathesis) reaction. Since you are given the products in the picture, you need to work backward to determine the reactants. Starting with the solid $SrSO_4(s)$, you know that the SO_4^{2-} anion started the reaction with a different cation (not Sr^{2+}). Since Na^+ is the only option, you can conclude that one of the reactants must be Na_2SO_4. Based on solubility rules, you know that Na_2SO_4 is soluble, so you represent it as $Na_2SO_4(aq)$. The remaining cation and anion indicate that the other reactant is the soluble $Sr(C_2H_3O_2)_2$. Observing the soluble and insoluble species in the picture, you can conclude that the molecular equation is

$$Na_2SO_4(aq) + Sr(C_2H_3O_2)_2(aq) \rightarrow SrSO_4(s) + 2NaC_2H_3O_2(aq)$$

b. Writing the strong electrolytes in the form of ions and the solid with its molecular formula, the complete ionic equation for the reaction is

$$2Na^+(aq) + SO_4^{2-}(aq) + Sr^{2+}(aq) + 2C_2H_3O_2^-(aq)$$

$$\rightarrow SrSO_4(s) + 2Na^+(aq) + 2C_2H_3O_2^-(aq)$$

c. After canceling the spectator ions, the net ionic equation for the reaction is

$$Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$$

4.3 a. MOH must be a base since OH^- is being produced in solution. It must be a strong base because the reaction indicates that MOH is completely soluble (strong electrolyte). In order to maintain charge balance in the formula, the element M must be a 1+ cation, probably a metal from Group IA of the periodic table. Examples of bases that fall into this category include NaOH and KOH.

b. This must be an acid since H^+ is being produced in solution. It is a weak acid because the double arrow is used, indicating only a partial ionization in solution. From the chemical reaction, A^- represents an anion with a 1- charge. Acetic acid, $HC_2H_3O_2$, is a weak acid of this type.

c. This must be an acid since H^+ is being produced in solution. $H_2A(aq)$ is a weak acid because the equation indicates only partial ionization in solution. A^{2-} represents an anion with a 2- charge. Carbonic acid, H_2CO_3, is a weak acid of this type.

(continued)

d. Examples of M include Na^+, K^+, and Li^+. Examples of A for reaction b include F^-, $C_2H_3O_2^-$, and CN^-. Examples of A for reaction c. include S^{2-}, CO_3^{2-}, and $C_4H_4O_6^{2-}$.

4.4 a. In order to answer this question, you need to compare the number of atoms of X per unit of volume. In order to compare volumes, use the lines on the sides of the beakers. Beaker A has concentration of five atoms per two volume units, 5/2 or 2.5/1. Beaker B has a concentration of ten atoms per one volume unit, 10/1. Beaker C has a concentration of ten atoms per two volume units, 10/2 or 5/1. Beaker D has a concentration of five atoms per volume unit, 5/1. Comparing the concentrations, the ranking from lowest to highest concentration is: Beaker A > Beaker C = Beaker D > Beaker B.

b. To make the concentrations of X equal in each beaker, they all have to be made to match the beaker with the lowest concentration. This is Beaker A, which has five atoms of X in one-half a beaker of solution. To make the concentrations equal, do the following: double the volume of Beakers C and D, and quadruple the volume of Beaker B. Overall, Beakers A and B will contain a full beaker of solution, and Beakers C and D will contain a half-beaker of solution.

4.5 a. Since flask C required three times the amount of titrant (NaOH) as acid A, you have learned that acid C has three times as many acidic protons as acid A. Since flask B required two times the amount of titrant as acid A, you have also learned that acid B has two times as many acidic protons as acid A.

b. If you assume that acid A contains a monoprotic acid, then you know the number of moles of A in the flask. After performing the titration, you know that the moles of NaOH must equal the moles of acid in flask A. You take the number of moles of NaOH and divide it by the volume of NaOH added during the titration to determine the concentration of the NaOH solution.

■ Answers to Review Questions

4.1 Some electrolyte solutions are strongly conducting because they are almost completely ionized and others are weakly conducting because they are weakly ionized. The former solutions will have many more ions to conduct electricity than will the latter solutions if both are present at the same concentrations.

4.2 A strong electrolyte is an electrolyte that exists in solution almost entirely as ions. An example is NaCl. When NaCl dissolves in water, it dissolves almost completely to give Na^+ and Cl^- ions. A weak electrolyte is an electrolyte that dissolves in water to give a relatively small percentage of ions. An example is NH_3. When NH_3 dissolves in water, it reacts very little with the water, so the level of NH_3 is relatively high, and the level of the NH_4^+ and OH^- ions is relatively low.

4.3 Soluble means the ability of a substance to dissolve in water. A compound is insoluble if it does not dissolve appreciably in water. An example of a soluble ionic compound is sodium chloride, NaCl, and for an insoluble ionic compound, an example is calcium carbonate, $CaCO_3$.

4.4 The advantage of using a molecular equation to represent an ionic equation is that it states explicitly what chemical species have been added and what chemical species are obtained as products. It also makes stoichiometric calculations easy to perform. The disadvantages are (1) the molecular equation does not represent the fact that the reaction actually involves ions, and (2) the molecular equation does not indicate which species exist as ions and which exist as molecular solids or molecular gases.

4.5 A spectator ion is an ion that does not take part in the reaction. In the following ionic reaction, the Na^+ and Cl^- are spectator ions:

$$Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

4.6 A net ionic equation is an ionic equation from which spectator ions have been canceled. The value of such an equation is that it shows the reaction that actually occurs at the ionic level. An example is the ionic equation representing the reaction of calcium chloride ($CaCl_2$) with potassium carbonate(K_2CO_3).

$$CaCl_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ KCl(aq):$$

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s) \qquad \text{(net ionic equation)}$$

4.7 The three major types of chemical reactions are precipitation reactions, acid-base reactions, and oxidation-reduction reactions. Oxidation-reduction reactions can be further classified as combination reactions, decomposition reactions, displacement reactions, and combustion reactions. Brief descriptions and examples of each are given below.

A precipitation reaction is a reaction that appears to involve the exchange of parts of the reactants. An example is: $2KCl\ (aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$.

An acid-base reaction, or neutralization reaction, results in an ionic compound and possibly water. An example is: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$.

A combination reaction is a reaction in which two substances combine to form a third substance. An example is: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$.

A decomposition reaction is a reaction in which a single compound reacts to give two or more substances. An example is: $2HgO(s) \xrightarrow{\Delta} 2Hg(l) + O_2(g)$.

(continued)

A displacement reaction, or single replacement reaction, is a reaction in which an element reacts with a compound displacing an element from it. An example is: $Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$.

A combustion reaction is a reaction of a substance with oxygen, usually with rapid release of heat to produce a flame. The products include one or more oxides. An example is: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$.

4.8 To prepare crystalline AgCl and $NaNO_3$, first make solutions of $AgNO_3$ and NaCl by weighing equivalent molar amounts of both solid compounds. Then mix the two solutions together, forming a precipitate of silver chloride and a solution of soluble sodium nitrate. Filter off the silver chloride, and wash it with water to remove the sodium nitrate solution. Then allow it to dry to obtain pure crystalline silver chloride. Finally, take the filtrate containing the sodium nitrate, and evaporate it, leaving pure crystalline sodium nitrate.

4.9 An example of a neutralization reaction is

$$\underset{\text{acid}}{HBr} + \underset{\text{base}}{KOH} \rightarrow \underset{\text{salt}}{KBr} + H_2O(l)$$

4.10 An example of a polyprotic acid is carbonic acid, H_2CO_3. The successive neutralization is given by the following molecular equations:

$$H_2CO_3(aq) + NaOH(aq) \rightarrow NaHCO_3(aq) + H_2O(l)$$

$$NaHCO_3(aq) + NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$$

4.11 Since an oxidation-reduction reaction is an electron transfer reaction, one substance must lose the electrons and be oxidized while another substance must gain electrons and be reduced.

4.12 A displacement reaction is an oxidation-reduction reaction in which a free element reacts with a compound, displacing an element from it.

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

Ag^+ is the oxidizing agent, and Cu is the reducing agent.

4.13 The number of moles present does not change when the solution is diluted.

4.14 The reaction is

$$HCl + NaOH \rightarrow NaCl + H_2O$$

After titration, the volume of hydrochloric acid is converted to moles of HCl using the molarity. Since the stoichiometry of the reaction is one mole HCl to one mole NaOH, these quantities are equal.

moles HCl = moles NaOH = molarity x volume

You could then multiply by the molar mass of NaOH to obtain the amount in the mixture.

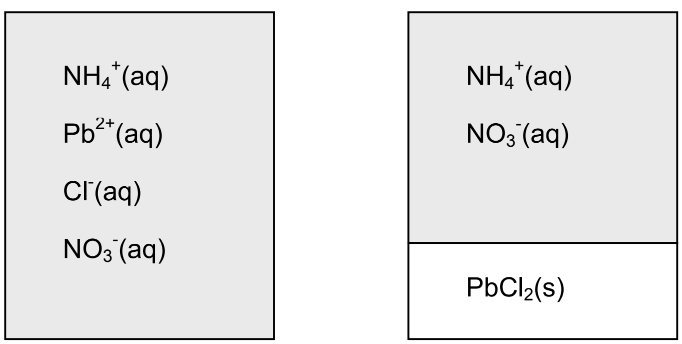
■ Answers to Conceptual Problems

4.15 a. Any soluble salt that will form a precipitate when reacted with Ag^+ ions in solution will work, for example: $CaCl_2$, Na_2S, $(NH_4)_2CO_3$.

b. No, no precipitate will form.

c. You would underestimate the amount of silver present in the solution.

4.16 a.



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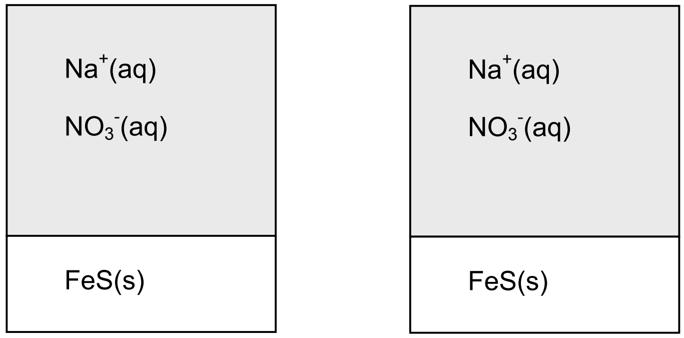
b.



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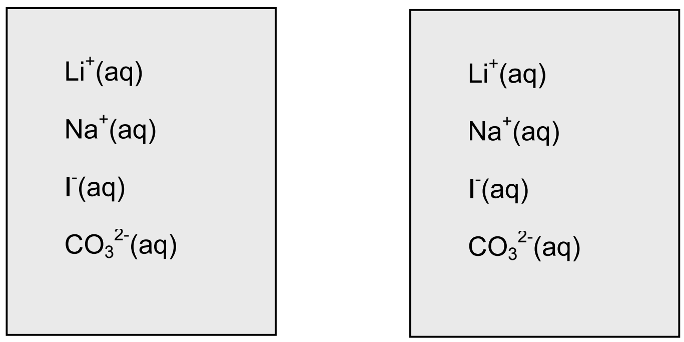
c.



t = 0 t > 0

4.17 a. $3Ca(C_2H_3O_2)_2(aq) + 2(NH_4)_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NH_4C_2H_3O_2(aq)$

b. $3Ca^{2+}(aq) + 6C_2H_3O_2^-(aq) + 6NH_4^+(aq) + 2PO_4^{3-}(aq)$
$\rightarrow Ca_3(PO_4)_2(s) + 6NH_4^+(aq) + 6C_2H_3O_2^-(aq)$

c. $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

4.18 a. Sample 1, Sample 2, Sample 3

b. They all will require the same volume of 0.1 M NaOH.

4.19 Probably not, since the ionic compound that is a nonelectrolyte is not soluble.

4.20 A good starting point is to identify the solution that contains the base. Since bases produce OH^- in aqueous solution, we would expect to see OH^- present in the BZ solution. The center beaker depicts OH^- in the solution, so it must be the base. By default, the remaining two beakers must contain acid. This is confirmed by the presence of H_3O^+ in both the left and right beakers. Keeping in mind that weak acids only partially dissociate, for the weak acid HA, we would expect to observe HA, H_3O^+, and A^- in the solution. In the case of the strong acid, HX, that completely dissociates, we would expect to observe only H_3O^+ and X^- in the solution. The beaker on the right only has H_3O^+ and one other species in solution, so it must be the strong acid HX. Examining the beaker on the left, there are three species present, which indicates that it must be the weak acid.

4.21 a. Since both solutions are made with compounds that contain chloride ions, the total chloride ion concentration is highest.

b. First, determine the concentrations of the compounds after mixing together. Use the dilution relationship, $M_1V_1 = M_2V_2$. Since equal volumes of equal molar solutions are mixed, the resulting concentrations are 0.50 M KBr and 0.50 M K_3PO_4.There is one Br^- ion per KBr, so the concentration of Br^- is 0.50 M. There is also one PO_4^{3-} ion per K_3PO_4, so its concentration is also 0.50 M. Potassium ion can be determined as follows.

$$0.50 \text{ M KBr} \times \frac{1 \text{ mol K}^+}{1 \text{ mol KBr}} + 0.50 \text{ M K}_3\text{PO}_4 \times \frac{3 \text{ mol K}^+}{1 \text{ mol K}_3\text{PO}_4} = 2.0 \text{ M K}^+$$

4.22 Since equal moles of the compounds are used, the highest Cl^- concentration is for the compound(s) with the largest subscript on Cl in the formula. Thus the order, from highest to lowest, is $AlCl_3 > PbCl_2 > KCl = NaCl = HCl > NH_3 = KOH = HCN$.

■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

4.23 a. Insoluble b. Soluble c. Soluble d. Soluble

4.24 a. Insoluble b. Soluble c. Soluble d. Soluble

4.25 a. Insoluble b. Soluble; The ions present would be NH_4^+ and SO_4^{2-}.

c. Insoluble d. Soluble; The ions present would be Na^+ and CO_3^{2-}.

4.26 a. Soluble; The ions present would be NH_4^+ and SO_4^{2-}. b. Insoluble

c. Insoluble d. Soluble; The ions present would be Ca^{2+} and NO_3^-.

4.27 a. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

b. $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$

c. $S^{2-}(aq) + 2H^+(aq) \rightarrow H_2S(g)$

d. $OH^-(aq) + NH_4^+(aq) \rightarrow NH_3(g) + H_2O(l)$

4.28 a. $H^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq)$

b. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

c. $Pb^{2+}(aq) + 2Br^-(aq) \rightarrow PbBr_2(s)$

d. $MgCO_3(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(l) + CO_2(g)$

4.29 Molecular equation: $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$

Net ionic equation: $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

4.30 Molecular equation: $K_2CO_3(aq) + 2HBr(aq) \rightarrow CO_2(g) + H_2O(l) + 2KBr(aq)$

Net ionic equation: $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

4.31 a. $FeSO_4(aq) + NaCl(aq) \rightarrow$ NR

b. $Na_2CO_3(aq) + MgBr_2(aq) \rightarrow MgCO_3(s) + 2NaBr(aq)$

$CO_3^{2-}(aq) + Mg^{2+}(aq) \rightarrow MgCO_3(s)$

c. $MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$

$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

d. $NiCl_2(aq) + NaBr(aq) \rightarrow$ NR

4.32 a. $AgNO_3(aq) + NaI(aq) \rightarrow AgI(s) + NaNO_3(aq)$

$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

b. $Ba(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KNO_3(aq)$

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

c. $Mg(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow NR$

d. $CaCl_2(aq) + Al(NO_3)_3(aq) \rightarrow NR$

4.33 a. $Ba(NO_3)_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + 2LiNO_3(aq)$

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

b. $Ca(NO_3)_2(aq) + NaBr(aq) \rightarrow NR$

c. $Al_2(SO_4)_3(aq) + 6NaOH(aq) \rightarrow 2Al(OH)_3(s) + 3Na_2SO_4(aq)$

$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

d. $3CaBr_2(aq) + 2Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NaBr(aq)$

$3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

4.34 a. $ZnCl_2(aq) + Na_2S(aq) \rightarrow ZnS(s) + 2NaCl(aq)$

$Zn^{2+}(aq) + S^{2-}(aq) \rightarrow ZnS(s)$

b. $CaCl_2(aq) + Na_2S(aq) \rightarrow NR$

c. $MgSO_4(aq) + KI(aq) \rightarrow NR$

d. $MgSO_4(aq) + K_2CO_3(aq) \rightarrow MgCO_3(s) + K_2SO_4(aq)$

$Mg^{2+}(aq) + CO_3^{2-}(aq) \rightarrow MgCO_3(s)$

4.35 a. Weak acid b. Strong base c. Strong acid d. Weak acid

4.36 a. Weak base b. Weak acid c. Strong base d. Strong acid

4.37 a. $NaOH(aq) + HNO_3(aq) \rightarrow H_2O(l) + NaNO_3(aq)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

b. $2HCl(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + BaCl_2(aq)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

c. $2HC_2H_3O_2(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(C_2H_3O_2)_2(aq)$

$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow H_2O(l) + C_2H_3O_2^-(aq)$

d. $NH_3(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$

$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$

4.38 a. $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

$Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$

b. $2HBr(aq) + Sr(OH)_2(s) \rightarrow 2H_2O(l) + SrBr_2(aq)$

$2H^+(aq) + Sr(OH)_2(s) \rightarrow 2H_2O(l) + Sr^{2+}(aq)$

c. $Ba(OH)_2(aq) + 2HC_2H_3O_2(aq) \rightarrow Ba(C_2H_3O_2)_2(aq) + 2H_2O(l)$

$OH^-(aq) + HC_2H_3O_2(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

d. $HNO_3(aq) + KOH(aq) \rightarrow H_2O(l) + KNO_3(aq)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

4.39 a. $2HBr(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + CaBr_2(aq)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

b. $3HNO_3(aq) + Al(OH)_3(s) \rightarrow 3H_2O(l) + Al(NO_3)_3(aq)$

$3H^+(aq) + Al(OH)_3(s) \rightarrow 3H_2O(l) + Al^{3+}(aq)$

c. $2HCN(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(CN)_2(aq)$

$HCN(aq) + OH^-(aq) \rightarrow H_2O(l) + CN^-(aq)$

d. $HCN(aq) + LiOH(aq) \rightarrow H_2O(l) + LiCN(aq)$

$HCN(aq) + OH^-(aq) \rightarrow H_2O(l) + CN^-(aq)$

4.40 a. $HClO_4(aq) + LiOH(aq) \rightarrow H_2O(l) + LiClO_4(aq)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

b. $2HNO_2(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + Ba(NO_2)_2(aq)$

$HNO_2(aq) + OH^-(aq) \rightarrow H_2O(l) + NO_2^-(aq)$

c. $HNO_2(aq) + NaOH(aq) \rightarrow H_2O(l) + NaNO_2(aq)$

$HNO_2(aq) + OH^-(aq) \rightarrow H_2O(l) + NO_2^-(aq)$

d. $2HCN(aq) + Sr(OH)_2(aq) \rightarrow 2H_2O(l) + Sr(CN)_2(aq)$

$HCN(aq) + OH^-(aq) \rightarrow H_2O(l) + CN^-(aq)$

4.41 a. $2KOH(aq) + H_3PO_4(aq) \rightarrow K_2HPO_4(aq) + 2H_2O(l)$

$2OH^-(aq) + H_3PO_4(aq) \rightarrow HPO_4^{2-}(aq) + 2H_2O(l)$

b. $3H_2SO_4(aq) + 2Al(OH)_3(s) \rightarrow 6H_2O(l) + Al_2(SO_4)_3(aq)$

$3H^+(aq) + Al(OH)_3(s) \rightarrow 3H_2O(l) + Al^{3+}(aq)$

c. $2HC_2H_3O_2(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(C_2H_3O_2)_2(aq)$

$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow H_2O(l) + C_2H_3O_2^-(aq)$

d. $H_2SO_3(aq) + NaOH(aq) \rightarrow H_2O(l) + NaHSO_3(aq)$

$H_2SO_3(aq) + OH^-(aq) \rightarrow HSO_3^-(aq) + H_2O(l)$

4.42 a. $Ca(OH)_2(aq) + 2H_2SO_4(aq) \rightarrow 2H_2O(l) + Ca(HSO_4)_2(aq)$

$OH^-(aq) + 2H^+ + SO_4^-(aq) \rightarrow H_2O(l) + HSO_4^-(aq)$ (ionizes to $H^+ + SO_4^{2-}$)

b. $2H_3PO_4(aq) + Ca(OH)_2(aq) \rightarrow H_2O(l) + Ca(H_2PO_4)_2(aq)$

$H_3PO_4(aq) + OH^-(aq) \rightarrow H_2O(l) + H_2PO_4^-(aq)$

c. $NaOH(aq) + H_2SO_4(aq) \rightarrow NaHSO_4(aq) + H_2O(l)$

$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

d. $Sr(OH)_2(aq) + 2H_2CO_3(aq) \rightarrow Sr(HCO_3)_2(aq) + 2H_2O(l)$

$OH^-(aq) + H_2CO_3(aq) \rightarrow HCO_3^-(aq) + H_2O(l)$

4.43 Molecular equations: $2H_2SO_3(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(HSO_3)_2(aq)$

$Ca(HSO_3)_2(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + 2CaSO_3(s)$

Ionic equations: $H_2SO_3(aq) + OH^-(aq) \rightarrow H_2O(l) + HSO_3^-(aq)$

$Ca^{2+}(aq) + HSO_3^-(aq) + OH^-(aq) \rightarrow CaSO_3(s) + H_2O(l)$

4.44 Molecular equations: $2H_3PO_4(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(H_2PO_4)_2(aq)$

$Ca(H_2PO_4)_2(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + 2CaHPO_4(aq)$

$2CaHPO_4(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca_3(PO_4)_2(s)$

Ionic equations: $H_3PO_4(aq) + OH^-(aq) \rightarrow H_2O(l) + H_2PO_4^-(aq)$

$H_2PO_4^-(aq) + OH^-(aq) \rightarrow H_2O(l) + HPO_4^{2-}(aq)$

$2HPO_4^{2-}(aq) + 2OH^-(aq) + 3Ca^{2+}(aq) \rightarrow 2H_2O(l) + Ca_3(PO_4)_2(s)$

4.45 a. Molecular equation: $CaS(aq) + 2HBr(aq) \rightarrow CaBr_2(aq) + H_2S(g)$

Ionic equation: $S^{2-}(aq) + 2H^+(aq) \rightarrow H_2S(g)$

b. Molecular equation: $MgCO_3(s) + 2HNO_3(aq) \rightarrow$
$Mg(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

Ionic equation: $MgCO_3(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$

c. Molecular equation: $K_2SO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + SO_2(g) + H_2O(l)$

Ionic equation: $SO_3^{2-}(aq) + 2H^+(aq) \rightarrow SO_2(g) + H_2O(l)$

4.46 a. Molecular equation: $BaCO_3(s) + 2HNO_3(aq) \rightarrow CO_2(g) + H_2O(l) + Ba(NO_3)_2(aq)$

Ionic equation: $BaCO_3(s) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l) + Ba^{2+}(aq)$

b. Molecular equation: $K_2S(aq) + 2HCl(aq) \rightarrow H_2S(g) + 2KCl(aq)$

Ionic equation: $S^{2-}(aq) + 2H^+(aq) \rightarrow H_2S(g)$

c. Molecular equation: $CaSO_3(s) + 2HI(aq) \rightarrow SO_2(g) + H_2O(l) + CaI_2(aq)$

Ionic equation: $CaSO_3(s) + 2H^+(aq) \rightarrow SO_2(g) + H_2O(l) + Ca^{2+}(aq)$

4.47 Molecular equation: $FeS(s) + 2HCl(aq) \rightarrow H_2S(g) + FeCl_2(aq)$

Ionic equation: $FeS(s) + 2H^+(aq) \rightarrow H_2S(g) + Fe^{2+}(aq)$

4.48 Molecular equation: $BaCO_3(s) + 2HBr(aq) \rightarrow BaBr_2(aq) + CO_2(g) + H_2O(l)$

Ionic equation: $BaCO_3(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + CO_2(g) + H_2O(l)$

4.49 a. Because all three O's = a total of -6, both Ga's = +6; thus, the oxidation number of Ga = +3.

b. Because both O's = a total of -4, the oxidation number of Nb = +4.

c. Because the four O's = a total of -8 and K = +1, the oxidation number of Br = +7.

d. Because the four O's = a total of -8 and the 2 K's = +2, the oxidation number of Mn = +6.

4.50 a. Because the three O's = a total of -6, the oxidation number of Cr = +6.

b. Because the two Cl's = a total of -2, both Hg's = +2; thus, the oxidation number of Hg = +1.

c. Because the three O's = a total of -6 and the 3 H's = a total of +3, the oxidation number of Ga = +3.

d. Because the four O's = a total of -8 and the 3 Na's = a total of +3, the oxidation number of P = +5.

4.51 a. Because the charge of -1 = $[x_N + 2$ (from 2 H's)], x_N must equal -3.

b. Because the charge of -1 = $[x_I - 6$ (from 3 O's)], x_I must equal +5.

c. Because the charge of -1 = $[x_{Al} - 8$ (4 O's) + 4 (4 H's)], x_{Al} must equal +3.

d. Because the charge of 0 = $[x_{Cl} - 8$ (4 O's) + 1 (1 H's)], x_{Cl} must equal +7.

4.52 a. Because the charge of -1 = $[x_N - 4$ (from 2 O's)], x_N must equal +3.

b. Because the charge of -2 = $[x_{Cr} - 8$ (from 4 O's)], x_{Cr} must equal +6.

c. Because the charge of -2 = $[x_{Zn} - 8$ (4 O's) + 4 (4 H's)], x_{Zn} must equal +2.

d. Because the charge of -1 = $[x_{As} - 6$ (3 O's) + 2 (2 H's)], x_{As} must equal +3.

4.53 a. From the list of common polyatomic anions in Table 2.6, the formula of the ClO_3 anion must be ClO_3^-. Thus, the oxidation state of Mn is Mn^{2+} (see also Table 4.5). Since the oxidation state of O is -2 and the net ionic charge is -1, the oxidation state of chlorine is determined by $x_{Cl} - 6 = -1$, so x_{Cl} must equal +5.

b. From the list of common polyatomic anions in Table 2.6, the formula of the CrO_4 anion must be CrO_4^{2-}. Thus, the oxidation state of Fe is Fe^{3+}. Since the oxidation state of O is -2 and the net ionic charge is -2, the oxidation state of Cr is determined by $x_{Cr} - 8 = -2$, so x_{Cr} must equal +6.

c. From the list of common polyatomic anions in Table 2.6, the formula of the Cr_2O_7 anion must be $Cr_2O_7^{2-}$. Thus, the oxidation state of Hg is Hg^{2+}. Since the oxidation state of O is -2 and the net ionic charge is -2, the oxidation state of Cr determined by $2x_{Cr} - 14 = -2$, so x_{Cr} must equal +6.

d. From the list of common polyatomic anions in Table 2.6, the formula of the PO_4 anion must be PO_4^{3-}. Thus, the oxidation state of Co is Co^{2+}. Since the oxidation state of O is -2 and the net ionic charge is -3, the oxidation state of P is determined by $x_P - 8 = -3$, so x_P must equal +5.

4.54 a. From the formula of ClO_3^- in the list of common polyatomic anions in Table 2.6, the formula of the BrO_3 anion must be BrO_3^-. Thus, the oxidation state of Hg is +1, and its formula must be Hg_2^{2+} (from Table 4.5). Since the oxidation state of O is -2 and the net ionic charge is -1, the oxidation state of Br is determined by $x_{Br} - 6 = -1$, so x_{Br} must equal +5.

b. From the list of common polyatomic anions in Table 2.6, the formula of the SO_4 anion must be SO_4^{2-}. Thus, the oxidation state of Cr is +3. Since the oxidation state of O is -2 and the net ionic charge is -2, the oxidation state of S is determined by $x_S - 8 = -2$, so x_S must equal +6.

c. From the formula of SO_4^{2-} in the list of common polyatomic anions in Table 2.6, the formula of the SeO_4 anion must be SeO_4^{2-}. Thus, the oxidation state of Co is +2. Since the oxidation state of O is -2 and the net ionic charge is -2, the oxidation state of Se is determined by $x_{Se} - 8 = -2$, so x_{Se} must be +6.

d. The formula of the hydroxide anion is OH^-. Thus, the oxidation state of Pb is +2. As usual, the oxidation state of H is +1 and the oxidation state of O is -2.

4.55 a. Phosphorus changes from an oxidation number of zero in P_4 to +5 in P_4O_{10}, losing electrons and acting as a reducing agent. Oxygen changes from an oxidation number of zero in O_2 to -2 in P_4O_{10}, gaining electrons and acting as an oxidizing agent.

b. Cobalt changes from an oxidation number of zero in Co(s) to +2 in $CoCl_2$, losing electrons and acting as a reducing agent. Chlorine changes from an oxidation number of zero in Cl_2 to -1 in $CoCl_2$, gaining electrons and acting as an oxidizing agent.

(continued)

4.56 a. Carbon changes from an oxidation number of zero in C to +2 in CO, losing electrons and acting as a reducing agent. Zinc changes from an oxidation number of +2 in ZnO to zero in Zn, gaining electrons and acting as an oxidizing agent.

b. Iron changes from an oxidation number of zero in Fe(s) to +2 in FeS, losing electrons and acting as a reducing agent. Sulfur changes from an oxidation number of 0 in S_8 to -2 in FeS, gaining electrons and acting as an oxidizing agent.

4.57 a. Al changes from oxidation number zero to +3; Al is the reducing agent. F changes from oxidation number zero to -1; F_2 is the oxidizing agent.

b. Hg changes from oxidation state +2 to 0; Hg^{2+} is the oxidizing agent. N changes from oxidation state +3 to +5; NO_2^- is the reducing agent.

4.58 a. C changes from oxidation number +2 to +4; the CO is the reducing agent. Fe changes from oxidation number +3 to 0; the Fe_2O_3 is the oxidizing agent.

b. S changes from oxidation number -2 to +6; the PbS is the reducing agent. O changes from oxidation number -1 to -2; the H_2O_2 is the oxidizing agent.

4.59 a. First, identify the species being oxidized and reduced, and assign the appropriate oxidation states. Since $CuCl_2$ and $AlCl_3$ are both soluble ionic compounds, Cl^- is a spectator ion and can be removed from the equation. The resulting net ionic equation is

$$\overset{+2}{Cu^{2+}}(aq) + \overset{0}{Al}(s) \rightarrow \overset{+3}{Al^{3+}}(aq) + \overset{0}{Cu}(s)$$

Next, write the half-reactions in an unbalanced form.

$$Al \rightarrow Al^{3+} \qquad \text{(oxidation)}$$
$$Cu^{2+} \rightarrow Cu \qquad \text{(reduction)}$$

Next, balance the charge in each equation by adding electrons to the more positive side to create balanced half-reactions.

$$Al \rightarrow Al^{3+} + 3e^- \qquad \text{(oxidation half-reaction)}$$
$$Cu^{2+} + 2e^- \rightarrow Cu \qquad \text{(reduction half-reaction)}$$

Multiply each half-reaction by a factor that will cancel out the electrons.

$$2 \times (Al \rightarrow Al^{3+} + 3e^-)$$
$$\underline{3 \times (Cu^{2+} + 2e^- \rightarrow Cu)}$$
$$3Cu^{2+} + 2Al + \cancel{6e^-} \rightarrow 2Al^{3+} + 3Cu + \cancel{6e^-}$$

(continued)

Therefore, the balanced oxidation-reduction reaction is

$$3Cu^{2+} + 2Al \rightarrow 2Al^{3+} + 3Cu$$

Finally, add six Cl^- ions to each side, and add phase labels. The resulting balanced equation is

$$3CuCl_2(aq) + 2Al(s) \rightarrow 2AlCl_3(aq) + 3Cu(s)$$

b. First, identify the species being oxidized and reduced, and assign the appropriate oxidation states.

$$\overset{+3}{Cr^{3+}}(aq) + \overset{0}{Zn}(s) \rightarrow \overset{0}{Cr}(s) + \overset{+2}{Zn^{2+}}(aq)$$

Next, write the half-reactions in an unbalanced form.

$$Zn \rightarrow Zn^{2+} \qquad \text{(oxidation)}$$

$$Cr^{3+} \rightarrow Cr \qquad \text{(reduction)}$$

Next, balance the charge in each equation by adding electrons to the more positive side to create balanced half-reactions.

$$Zn \rightarrow Zn^{2+} + 2e^- \qquad \text{(oxidation half-reaction)}$$

$$Cr^{3+} + 3e^- \rightarrow Cr \qquad \text{(reduction half-reaction)}$$

Multiply each half-reaction by a factor that will cancel out the electrons.

$$\begin{array}{l} 3 \times (Zn \rightarrow Zn^{2+} + 2e^-) \\ \underline{\qquad\qquad 2 \times (Cr^{3+} + 3e^- \rightarrow Cr)} \\ 2Cr^{3+} + 3Zn + \cancel{6e^-} \rightarrow 2Cr + 3Zn^{2+} + \cancel{6e^-} \end{array}$$

Therefore, the balanced oxidation-reduction reaction, including phase labels, is

$$2Cr^{3+}(aq) + 3Zn(s) \rightarrow 2Cr(s) + 3Zn^{2+}(aq)$$

4.60 a. First, identify the species being oxidized and reduced, and assign the appropriate oxidation states. Since FeI_3 and MgI_2 are both soluble ionic compounds, I^- is a spectator ion and can be removed from the equation. The resulting net ionic equation is

$$\overset{+3}{Fe^{3+}}(aq) + \overset{0}{Mg}(s) \rightarrow \overset{0}{Fe}(s) + \overset{+2}{Mg^{2+}}(aq)$$

Next, write the half-reactions in an unbalanced form.

$$Mg \rightarrow Mg^{2+} \qquad \text{(oxidation)}$$

$$Fe^{3+} \rightarrow Fe \qquad \text{(reduction)}$$

(continued)

Next, balance the charge in each equation by adding electrons to the more positive side to create balanced half-reactions.

$Mg \rightarrow Mg^{2+} + 2e^-$ (oxidation half-reaction)

$Fe^{3+} + 3e^- \rightarrow Fe$ (reduction half-reaction)

Multiply each half-reaction by a factor that will cancel out the electrons.

$$
\begin{array}{l}
3 \times (Mg \rightarrow Mg^{2+} + 2e^-) \\
\underline{\qquad\qquad 2 \times (Fe^{3+} + 3e^- \rightarrow Fe)} \\
2Fe^{3+} + 3Mg + \cancel{6e^-} \rightarrow 2Fe + 3Mg^{2+} + \cancel{6e^-}
\end{array}
$$

Therefore, the balanced oxidation-reduction reaction is

$$2Fe^{3+} + 3Mg \rightarrow 2Fe + 3Mg^{2+}$$

Finally, add six I^- ions to each side, and add phase labels. The resulting balanced equation is

$$2FeI_3(aq) + 3Mg(s) \rightarrow 2Fe(s) + 3MgI_2(aq)$$

b. First, identify the species being oxidized and reduced, and assign the appropriate oxidation states.

$$\overset{0}{H_2}(g) + \overset{+1}{Ag^+}(aq) \rightarrow \overset{0}{Ag}(s) + \overset{+1}{H^+}(aq)$$

Next, write the half-reactions in an unbalanced form, making sure that mass is balanced.

$H_2 \rightarrow 2H^+$ (oxidation)

$Ag^+ \rightarrow Ag$ (reduction)

Next, balance the charge in each equation by adding electrons to the more positive side to create balanced half-reactions.

$H_2 \rightarrow 2H^+ + 2e^-$ (oxidation half-reaction)

$Ag^+ + e^- \rightarrow Ag$ (reduction half-reaction)

Multiply each half-reaction by a factor that will cancel out the electrons.

$$
\begin{array}{l}
1 \times (H_2 \rightarrow 2H^+ + 2e^-) \\
\underline{\qquad\qquad 2 \times (Ag^+ + e^- \rightarrow Ag)} \\
H_2 + 2Ag^+ + \cancel{2e^-} \rightarrow 2Ag + 2H^+ + \cancel{2e^-}
\end{array}
$$

Therefore, the balanced oxidation-reduction reaction, including phase labels, is

$$H_2(g) + 2Ag^+(aq) \rightarrow 2Ag(s) + 2H^+(aq)$$

4.61 $\text{Molarity} = \dfrac{\text{moles solute}}{\text{liters solution}} = \dfrac{0.0512\text{ mol}}{0.0250\text{ L}} = 2.0\underline{4}8 = 2.05\text{ M}$

4.62 $\text{Molarity} = \dfrac{\text{moles solute}}{\text{liters solution}} = \dfrac{0.0285\text{ mol}}{0.0500\text{ L}} = 0.57\underline{0}0 = 0.570\text{ M}$

4.63 Find the number of moles of solute ($KMnO_4$) using the molar mass of 158.03 g $KMnO_4$ per one mol $KMnO_4$:

$$0.798\text{ g }KMnO_4 \times \frac{1\text{ mol }KMnO_4}{158.03\text{ g }KMnO_4} = 5.0\underline{4}97 \times 10^{-3}\text{ mol }KMnO_4$$

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters of solution}} = \frac{5.0497 \times 10^{-3}\text{ mol}}{0.0500\text{ L}} = 0.10\underline{0}99 = 0.101\text{ M}$$

4.64 Find the number of moles of solute ($H_2C_2O_4$) using the molar mass of 90.04 g $H_2C_2O_4$ per one mol $H_2C_2O_4$:

$$1.192\text{ g }H_2C_2O_4 \times \frac{1\text{ mol }H_2C_2O_4}{90.04\text{ g }H_2C_2O_4} = 0.0132\underline{3}8\text{ mol }H_2C_2O_4$$

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters solution}} = \frac{0.013238\text{ mol}}{0.1000\text{ L}} = 0.132\underline{3}8 = 0.1324\text{ M}$$

4.65 $0.150\text{ mol }CuSO_4 \times \dfrac{1\text{ L solution}}{0.120\text{ mol }CuSO_4} = 1.2\underline{5}0 = 1.25\text{ L solution}$

4.66 $0.102\text{ mol }HClO_4 \times \dfrac{1\text{ L solution}}{0.126\text{ mol }HClO_4} = 0.80\underline{9}52\text{ L} = 0.810\text{ L soln}$

4.67 $0.0353\text{ g KOH} \times \dfrac{1\text{ mol KOH}}{56.10\text{ g KOH}} \times \dfrac{1\text{ L solution}}{0.0176\text{ mol KOH}} = 0.035\underline{7}51\text{ L} = 35.8\text{ mL}$

4.68 $0.949\text{ g }H_2SO_4 \times \dfrac{1\text{ mol }H_2SO_4}{98.09\text{ g }H_2SO_4} \times \dfrac{1\text{ L soln}}{0.215\text{ mol }H_2SO_4}$

$$= 0.044\underline{9}99\text{ L} \quad (45.0\text{ mL soln})$$

4.69 From the molarity, one L of heme solution is equivalent to 0.0019 mol of heme solute. Before starting the calculation, note that 150 mL of solution is equivalent to 150×10^{-3} L of solution:

$$150 \times 10^{-3} \text{ L soln} \times \frac{0.0019 \text{ mol heme}}{1 \text{ L solution}} = 2.\underline{8}50 \times 10^{-4} = 2.9 \times 10^{-4} \text{ mol heme}$$

4.70 From the molarity, one L of insulin solution is equivalent to 0.0048 mol of insulin solute. Before starting the calculation, note that 28 mL of solution is equivalent to 28×10^{-3} L of solution:

$$28 \times 10^{-3} \text{ L soln} \times \frac{0.0048 \text{ mol insulin}}{1 \text{ L soln}} = 1.\underline{3}44 \times 10^{-4} = 1.3 \times 10^{-4} \text{ mol insulin}$$

4.71 Multiply the volume of solution by molarity to convert it to moles; then convert to mass of solute by multiplying by the molar mass:

$$100.0 \times 10^{-3} \text{ L soln} \times \frac{0.025 \text{ mol } Na_2Cr_2O_7}{1 \text{ L solution}} \times \frac{262.0 \text{ g } Na_2Cr_2O_7}{1 \text{ mol } Na_2Cr_2O_7}$$

$$= 0.6\underline{5}50 = 0.66 \text{ g } Na_2Cr_2O_7$$

4.72 Multiply the desired volume of solution by the molarity to convert it to moles; then convert to mass of solute by multiplying by the molar mass:

$$0.250 \text{ L soln} \times \frac{0.20 \text{ mol } Na_2SO_4}{1 \text{ L solution}} \times \frac{142.05 \text{ g } Na_2SO_4}{1 \text{ mol } Na_2SO_4} = 7.\underline{1}0 = 7.1 \text{ g } Na_2SO_4$$

Weigh out 7.1 g of pure Na_2SO_4, and place it in a 250-mL volumetric flask. Add enough water to fill the flask to the mark on the neck.

4.73 Use the rearranged version of the dilution formula to calculate the initial volume of 15.8 M HNO_3 required:

$$V_i = \frac{M_f V_f}{M_i} = \frac{0.12 \text{ M} \times 1000 \text{ mL}}{15.8 \text{ M}} = 7.\underline{5}9 = 7.6 \text{ mL}$$

4.74 Use the rearranged version of the dilution formula to calculate the initial volume of 12.4 M HCl required:

$$V_i = \frac{M_f V_f}{M_i} = \frac{0.50 \text{ M} \times 1500 \text{ mL}}{12.4 \text{ M}} = 6\underline{0}.4 = 60. \text{ mL}$$

4.75 The initial concentration of KCl (molar mass, 74.55 g/mol) is

$$3.50 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \times \frac{1}{0.0100 \text{ L}} = 4.6\underline{9}4 \text{ M}$$

Using the dilution factor, $M_1V_1 = M_2V_2$, with V_2 = 10.0 mL + 60.0 mL = 70.0 mL, after the solutions are mixed, the concentration of KCl is

$$M_2 = \frac{M_1V_1}{V_2} = \frac{4.694 \text{ M} \times 10.0 \text{ mL}}{70.0 \text{ mL}} = 0.67\underline{0}69 \text{ M KCl}$$

For $CaCl_2$, the concentration is

$$M_2 = \frac{M_1V_1}{V_2} = \frac{0.500 \text{ M} \times 10.0 \text{ mL}}{70.0 \text{ mL}} = 0.42\underline{8}57 \text{ M } CaCl_2$$

Therefore, the concentrations of the ions are 0.671 M K^+ and 0.429 M Ca^{2+}. For Cl^-, it is 0.67$\underline{0}$69 M + 2 x 0.42$\underline{8}$57 M = 1.52$\underline{7}$8 M = 1.528 M

4.76 Using the dilution factor, $M_1V_1 = M_2V_2$, with V_2 = 50.0 mL + 25.0 mL = 75.0 mL, after the solutions are mixed, the concentration of $NaClO_3$ is

$$M_2 = \frac{M_1V_1}{V_2} = \frac{0.20 \text{ M} \times 50.0 \text{ mL}}{75.0 \text{ mL}} = 0.1\underline{3}33 \text{ M } NaClO_3$$

For Na_2SO_4, the concentration is

$$M_2 = \frac{M_1V_1}{V_2} = \frac{0.20 \text{ M} \times 25.0 \text{ mL}}{75.0 \text{ mL}} = 0.06\underline{6}66 \text{ M } Na_2SO_4$$

Therefore, the concentrations of the ions are 0.13 M ClO_3^- and 0.067 M SO_4^{2-}. For Na^+, it is 0.1$\underline{3}$33 M + 2 x 0.06$\underline{6}$66 M = 0.2$\underline{6}$66 = 0.27 M.

4.77 Use the appropriate conversion factors to convert the mass of $BaSO_4$ to the mass of Ba^{2+} ions:

$$0.513 \text{ g } BaSO_4 \times \frac{1 \text{ mol } BaSO_4}{233.40 \text{ g } BaSO_4} \times \frac{1 \text{ mol } Ba^{2+}}{1 \text{ mol } BaSO_4} \times \frac{137.33 \text{ g } Ba^{2+}}{1 \text{ mol } Ba^{2+}}$$

$$= 0.30\underline{1}84 \text{ g } Ba^{2+}$$

(continued)

Then calculate the percentage of barium in the 458 mg (0.458 g) compound:

$$\frac{0.30184 \text{ g } Ba^{2+}}{0.458 \text{ g}} \times 100\% = 65.\underline{9}039 = 65.9\% \; Ba^{2+}$$

4.78 Use the appropriate conversion factors to convert the mass of AgI to the mass of I^- ions:

$$2.185 \text{ g AgI} \times \frac{1 \text{ mol AgI}}{234.77 \text{ g AgI}} \times \frac{1 \text{ mol } I^-}{1 \text{ mol AgI}} \times \frac{126.90 \text{ g } I^-}{1 \text{ mol } I^-} = 1.18\underline{1}05 \text{ g } I^-$$

Then calculate the percentage of iodine in the 1.545 g compound:

$$\frac{1.18105 \text{ g } I^-}{1.545 \text{ g}} \times 100\% = 76.4\underline{4}3 = 76.44\% \; I^-$$

4.79 a. The mass of chloride ion in the AgCl from the copper chloride compound is:

$$86.00 \text{ mg AgCl} \times \frac{35.45 \text{ mg } Cl^-}{143.32 \text{ mg AgCl}} = 21.2\underline{7}1 \text{ mg } Cl^-$$

The percentage of chlorine in the 59.40 mg sample is:

$$\frac{21.2\underline{7}1 \text{ mg } Cl^-}{59.40 \text{ mg sample}} \times 100\% = 35.8\underline{0}9 = 35.81\% \; Cl^-$$

b. Of the various approaches, it is as easy to calculate the theoretical percentage of Cl^- in both CuCl and $CuCl_2$ as it is to use another approach:

$$\text{CuCl: } \frac{35.45 \text{ mg } Cl^-}{99.00 \text{ mg CuCl}} \times 100\% = 35.8\underline{0}8\%$$

$$CuCl_2\text{: } \frac{70.90 \text{ mg } Cl^-}{134.45 \text{ mg } CuCl_2} \times 100\% = 52.7\underline{3}3\%$$

The compound is obviously CuCl.

4.80 a. The mass of chloride ion in the AgCl from the gold chloride compound is:

$$100.3 \text{ mg AgCl} \times \frac{35.45 \text{ mg Cl}^-}{143.32 \text{ mg AgCl}} = 24.8\underline{0}9 \text{ mg Cl}^-$$

The percentage of chloride in the 162.7 mg sample is:

$$\frac{24.8\underline{0}9 \text{ mg Cl}^-}{162.7 \text{ mg sample}} \times 100\% = 15.2\underline{4}8 = 15.25\% \text{ Cl}^-$$

b. Of the various approaches, it is as easy to calculate the theoretical percentage of Cl^- in both AuCl and $AuCl_3$ as it is to use another approach:

$$\text{AuCl: } \frac{35.45 \text{ mg Cl}^-}{232.4 \text{ mg AuCl}} \times 100\% = 15.2\underline{5}3\%$$

$$\text{AuCl}_3\text{: } \frac{106.35 \text{ mg Cl}^-}{303.32 \text{ mg AuCl}_3} \times 100\% = 35.06\underline{2}0\%$$

The compound is obviously AuCl.

4.81 First, calculate the moles of chlorine in the compound:

$$0.3048 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol AgCl}} = 0.00212\underline{6}7 \text{ mol Cl}^-$$

Then, calculate the g Fe^{x+} from the g Cl^-:

$$\text{g Fe}^{x+} = 0.1348 \text{ g comp} - \left(0.0021267 \text{ mol Cl}^- \times \frac{35.45 \text{ g Cl}^-}{1 \text{ mol Cl}^-}\right)$$

$$= 0.0594\underline{0}8 \text{ g Fe}^{x+}$$

Now, calculate the moles of Fe^{x+} using the molar mass:

$$0.0594\underline{0}8 \text{ g Fe}^{x+} \times \frac{1 \text{ mol Fe}^{x+}}{55.85 \text{ g Fe}^{x+}} = 0.00106\underline{3}7 \text{ mol Fe}^{x+}$$

Finally, divide the mole numbers by the smallest mole number:

$$\text{For Cl: } \frac{0.002127 \text{ mol Cl}^-}{0.0010637 \text{ mol}} = 2.00; \text{ for Fe}^{x+}\text{: } \frac{0.0010637 \text{ mol Fe}^{x+}}{0.0010637 \text{ mol}} = 1.00$$

Thus, the formula is $FeCl_2$.

4.82 First, calculate the moles of Ba^{2+} in the compound:

$$2.012 \text{ g BaCrO}_4 \times \frac{1 \text{ mol BaCrO}_4}{253.33 \text{ g BaCrO}_4} \times \frac{1 \text{ mol Ba}^{2+}}{1 \text{ mol BaCrO}_4} = 0.00794\underline{2}2 \text{ mol Ba}^{2+}$$

Next, calculate the g O from the g Ba^{2+}:

$$\text{g O} = 1.345 \text{ g comp} - \left(0.0079422 \text{ mol Ba}^{2+} \times \frac{137.33 \text{ g Ba}^{2+}}{1 \text{ mol Ba}^{2+}}\right)$$

$$= 0.254\underline{3}0 \text{ mol O}$$

Now, calculate the moles of O using the molar mass:

$$0.25430 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.0158\underline{9}4 \text{ mol O}$$

Finally, divide the mole numbers by the smallest mole number:

$$\text{For O: } \frac{0.015894 \text{ mol O}}{0.0079422 \text{ mol}} = 2.00; \text{ for Ba}^{2+}\text{: } \frac{0.0079422 \text{ mol Ba}^{2+}}{0.0079422 \text{ mol}} = 1.00$$

Thus, the formula is BaO_2 (barium peroxide).

4.83 Using molarity, convert the volume of Na_2CO_3 to moles of Na_2CO_3; then use the equation to convert to moles of HNO_3, and finally to volume:

$$2HNO_3 + Na_2CO_3 \rightarrow 2NaNO_3 + H_2O + CO_2$$

$$44.8 \times 10^{-3} \text{ L Na}_2\text{CO}_3 \times \frac{0.150 \text{ mol Na}_2\text{CO}_3}{1 \text{ L soln}} \times \frac{2 \text{ mol HNO}_3}{1 \text{ mol Na}_2\text{CO}_3}$$

$$\times \frac{1 \text{ L HNO}_3}{0.250 \text{ mol HNO}_3} = 0.053\underline{7}6 \text{ L} = 0.0538 \text{ L} = 53.8 \text{ mL}$$

4.84 As in the previous problem, use molarity to convert to volume; then use the equation to convert to moles of the other reactant, and finally to volume:

$$Na_2CO_3 + Ca(OH)_2 \rightarrow CaCO_3 + 2NaOH$$

$$49.8 \times 10^{-3} \text{ L Ca(OH)}_2 \times \frac{0.150 \text{ mol Ca(OH)}_2}{1 \text{ L soln}} \times \frac{1 \text{ mol Na}_2\text{CO}_3}{1 \text{ mol Ca(OH)}_2}$$

$$\times \frac{1 \text{ L Na}_2\text{CO}_3}{0.350 \text{ mol Na}_2\text{CO}_3} = 0.021\underline{3}4 \text{ L (21.3 mL) Na}_2\text{CO}_3$$

4.85 The reaction is $H_2SO_4 + 2NaHCO_3 \rightarrow Na_2SO_4 + 2H_2O + CO_2$.

$$8.20 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaHCO}_3} \times \frac{1 \text{ L soln}}{0.150 \text{ mol H}_2\text{SO}_4}$$

$$= 0.32\underline{5}3 \text{ L (325 mL) soln}$$

4.86 The reaction is:

$$10FeSO_4 + 2KMnO_4 + 8H_2SO_4 \rightarrow 5Fe_2(SO_4)_3 + 2MnSO_4 + K_2SO_4 + 4H_2O$$

$$3.36 \text{ g FeSO}_4 \times \frac{1 \text{ mol FeSO}_4}{151.92 \text{ g FeSO}_4} \times \frac{2 \text{ mol KMnO}_4}{10 \text{ mol FeSO}_4} \times \frac{1 \text{ L soln}}{0.250 \text{ mol KMnO}_4}$$

$$= 0.017\underline{6}9 \text{ L (17.7 mL) soln}$$

4.87 First, find the mass of H_2O_2 required to react with $KMnO_4$.

$$5H_2O_2 + 2KMnO_4 + 3H_2SO_4 \rightarrow 5O_2 + 2MnSO_4 + K_2SO_4 + 8H_2O$$

$$51.7 \times 10^{-3} \text{ L KMnO}_4 \times \frac{0.145 \text{ mol KMnO}_4}{1 \text{ L soln}} \times \frac{5 \text{ mol H}_2\text{O}_2}{2 \text{ mol KMnO}_4} \times \frac{34.02 \text{ g H}_2\text{O}_2}{1 \text{ mol H}_2\text{O}_2}$$

$$= 0.63\underline{7}5 \text{ g H}_2\text{O}_2$$

Percent H_2O_2 = (mass H_2O_2 ÷ mass sample) x 100%

$$= (0.6375 \text{ g} \div 20.0 \text{ g}) \times 100\% = 3.1\underline{8}7 = 3.19\%$$

4.88 First, find the mass of Fe^{2+} required to react with the $K_2Cr_2O_7$.

$$6FeSO_4 + K_2Cr_2O_7 + 7H_2SO_4 \rightarrow 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + 7H_2O + K_2SO_4$$

$$43.7 \times 10^{-3} \text{ L KMnO}_4 \times \frac{0.150 \text{ mol K}_2\text{Cr}_2\text{O}_7}{1 \text{ L soln}} \times \frac{6 \text{ mol FeSO}_4}{1 \text{ mol K}_2\text{Cr}_2\text{O}_7}$$

$$\times \frac{1 \text{ mol Fe}^{2+}}{1 \text{ mol FeSO}_4} \times \frac{55.85 \text{ g Fe}^{2+}}{1 \text{ mol Fe}^{2+}} = 2.1\underline{9}6 \text{ g Fe}^{2+}$$

(continued)

$2.196 \text{ g } Fe^{2+}$ in reaction = $2.196 \text{ g } Fe^{2+}$ in ore, so

Percent Fe = (mass Fe^{2+} ÷ mass ore) x 100% = (2.196 g ÷ 3.33 g) x 100%

= 65.$\underline{9}$6 = 66.0%

■ Solutions to General Problems

4.89 For the reaction of magnesium metal and hydrobromic acid, the equations are as follows.

Molecular equation: $Mg(s) + 2HBr(aq) \rightarrow H_2(g) + MgBr_2(aq)$

Ionic equation: $Mg(s) + 2H^+(aq) \rightarrow H_2(g) + Mg^{2+}(aq)$

4.90 For the reaction of aluminum metal and perchloric acid, the equations are as follows.

Molecular equation: $2Al(s) + 6HClO_4(aq) \rightarrow 3H_2(g) + 2Al(ClO_4)_3(aq)$

Ionic equation: $2Al(s) + 6H^+(aq) \rightarrow 3H_2(g) + 2Al^{3+}(aq)$

4.91 For the reaction of nickel(II) sulfate and lithium hydroxide, the equations are as follows.

Molecular equation: $NiSO_4(aq) + 2LiOH(aq) \rightarrow Ni(OH)_2(s) + Li_2SO_4(aq)$

Ionic equation: $Ni^{2+}(aq) + 2OH^-(aq) \rightarrow Ni(OH)_2(s)$

4.92 For the reaction of potassium sulfate and barium bromide, the equations are as follows.

Molecular equation: $BaBr_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KBr(aq)$

Ionic equation: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

4.93 a. Molecular equation: $LiOH(aq) + HCN(aq) \rightarrow LiCN(aq) + H_2O(l)$

Ionic equation: $OH^-(aq) + HCN(aq) \rightarrow CN^-(aq) + H_2O(l)$

b. Molecular equation: $Li_2CO_3(aq) + 2HNO_3(aq) \rightarrow 2LiNO_3(aq) + CO_2(g) + H_2O(l)$

Ionic equation: $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

(continued)

c. Molecular equation: $LiCl(aq) + AgNO_3(aq) \rightarrow LiNO_3(aq) + AgCl(s)$

Ionic equation: $Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

d. Molecular equation: $MgSO_4(aq) + LiCl(aq) \rightarrow NR$

(Li_2SO_4 and $MgCl_2$ are soluble.)

4.94 a. Molecular equation: $Al(OH)_3(s) + 3HNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3H_2O(l)$

Ionic equation: $Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$

b. Molecular equation: $FeS(s) + 2HClO_4(aq) \rightarrow Fe(ClO_4)_2(aq) + H_2S(g)$

Ionic equation: $FeS(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2S(g)$

c. Molecular equation: $CaCl_2(aq) + NaNO_3(aq) \rightarrow NR$

[$Ca(NO_3)_2$ and NaCl are soluble.]

d. Molecular equation: $MgSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + BaSO_4(s)$

Ionic equation: $SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$

4.95 a. Molecular equation: $Sr(OH)_2(aq) + 2HC_2H_3O_2(aq) \rightarrow$

$Sr(C_2H_3O_2)_2(aq) + 2H_2O(l)$

Ionic equation: $HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

b. Molecular equation: $NH_4I(aq) + CsCl(aq) \rightarrow NR$

(NH_4Cl and CsI are soluble.)

c. Molecular equation: $NaNO_3(aq) + CsCl(aq) \rightarrow NR$

(NaCl and $CsNO_3$ are soluble.)

d. Molecular equation: $NH_4I(aq) + AgNO_3(aq) \rightarrow NH_4NO_3(aq) + AgI(s)$

Ionic equation: $I^-(aq) + Ag^+(aq) \rightarrow AgI(s)$

4.96 a. Molecular equation: $2HClO_4(aq) + BaCO_3(s) \rightarrow$

$Ba(ClO_4)_2(aq) + H_2CO_3(aq)$ [$CO_2(g) + H_2O(l)$]

Ionic equation: $2H^+(aq) + BaCO_3(s) \rightarrow Ba^{2+}(aq) + CO_2(g) + H_2O(l)$

(continued)

b. Molecular equation: $H_2CO_3(aq) + Sr(OH)_2(aq) \rightarrow 2H_2O(l) + SrCO_3(s)$

Ionic equation: $H_2CO_3(aq) + Sr^{2+}(aq) + 2OH^-(aq) \rightarrow 2H_2O(l) + SrCO_3(s)$

c. Molecular equation: $2K_3PO_4(aq) + 3MgCl_2(aq) \rightarrow 6KCl(aq) + Mg_3(PO_4)_2(s)$

Ionic equation: $2PO_4^{3-}(aq) + 3Mg^{2+}(aq) \rightarrow Mg_3(PO_4)_2(s)$

d. Molecular equation: $FeSO_4(aq) + MgCl_2(aq) \rightarrow NR$

($FeCl_2$ and $MgSO_4$ are soluble.)

4.97 For each preparation, the compound to be prepared is given first, followed by the compound from which it is to be prepared. Then the method of preparation is given, followed by the molecular equation for the preparation reaction. Steps such as evaporation, etc., are not given in the molecular equation.

a. To prepare $CuCl_2$ from $CuSO_4$, add a solution of $BaCl_2$ to a solution of the $CuSO_4$, precipitating $BaSO_4$. The $BaSO_4$ can be filtered off, leaving aqueous $CuCl_2$, which can be obtained in solid form by evaporation. Molecular equation:

$$CuSO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + CuCl_2(aq)$$

b. To prepare $Ca(C_2H_3O_2)_2$ from $CaCO_3$, add a solution of acetic acid, $HC_2H_3O_2$, to the solid $CaCO_3$, forming CO_2, H_2O, and aqueous $Ca(C_2H_3O_2)_2$. The aqueous $Ca(C_2H_3O_2)_2$ can be converted to the solid form by evaporation, which also removes the CO_2 and H_2O products. Molecular equation:

$$CaCO_3(s) + 2HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + CO_2(g) + H_2O(l)$$

c. To prepare $NaNO_3$ from Na_2SO_3, add a solution of nitric acid, HNO_3, to the solid Na_2SO_3, forming SO_2, H_2O, and aqueous $NaNO_3$. The aqueous $NaNO_3$ can be converted to the solid by evaporation, which also removes the SO_2 and H_2O products. Molecular equation:

$$Na_2SO_3(s) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + SO_2(g) + H_2O(l)$$

d. To prepare $MgCl_2$ from $Mg(OH)_2$, add a solution of hydrochloric acid (HCl) to the solid $Mg(OH)_2$, forming H_2O and aqueous $MgCl_2$. The aqueous $MgCl_2$ can be converted to the solid form by evaporation. Molecular equation:

$$Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$$

4.98 For each preparation, the compound to be prepared is given first and followed by the compound from which it is to be prepared. Then, the method of preparation is given, followed by the molecular equation for the preparation reaction. Steps such as evaporation, etc., are not given in the molecular equation.

a. To prepare $MgCl_2$ from $MgCO_3$, add a solution of hydrochloric acid (HCl) to the solid $MgCO_3$, forming CO_2, H_2O, and aqueous $MgCl_2$. The aqueous $MgCl_2$ can be converted to the solid form by evaporation. Molecular equation:

$$MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$$

b. To prepare $NaNO_3$ from NaCl, add a solution of $AgNO_3$ to a solution of the NaCl, precipitating AgCl. The AgCl can be filtered off, leaving aqueous $NaNO_3$, which can be obtained in solid form by evaporation. Molecular equation:

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

c. To prepare $Al(OH)_3$ from $Al(NO_3)_3$, add a solution of NaOH to a solution of $Al(NO_3)_3$, precipitating $Al(OH)_3$. The $Al(OH)_3$ can be filtered off and dried to remove any water adhering to it. Molecular equation:

$$Al(NO_3)_3(aq) + 3NaOH(aq) \rightarrow Al(OH)_3(s) + 3NaNO_3(aq)$$

d. To prepare HCl from H_2SO_4, add a solution of $BaCl_2$ to the solution of H_2SO_4, precipitating $BaSO_4$. The $BaSO_4$ can be filtered off, leaving the desired solution of aqueous HCl. Molecular equation:

$$H_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$$

4.99	a. Decomposition	b. Decomposition	c. Combination	d. Displacement
4.100	a. Combination	b. Displacement	c. Decomposition	d. Combination

4.101 a. $Pb(NO_3)_2 + H_2SO_4$ [$\rightarrow PbSO_4(s) + HNO_3(aq)$]

$Pb(NO_3)_2 + MgSO_4$ [$\rightarrow PbSO_4(s) + Mg(NO_3)_2(aq)$]

$Pb(NO_3)_2 + Ba(OH)_2$ [$\rightarrow Pb(OH)_2(s) + Ba(NO_3)_2(aq)$]

b. $Ba(OH)_2 + MgSO_4$ [$\rightarrow BaSO_4(s) + Mg(OH)_2(s)$]

c. $Ba(OH)_2 + H_2SO_4$ [$\rightarrow BaSO_4(s) + H_2O(l)$]

4.102 a. $AgNO_3 + H_3PO_4$ [$\rightarrow Ag_3PO_4(s) + HNO_3(aq)$]

$AgNO_3 + Sr(OH)_2$ [$\rightarrow AgOH(s) + Sr(NO_3)_2(aq)$]

(continued)

b. $Sr(OH)_2 + CuSO_4$ $[\rightarrow\ SrSO_4(s) + Cu(OH)_2(s)\ ]$

c. $Sr(OH)_2 + H_3PO_4$ $[\rightarrow\ Sr_3(PO_4)_2(s) + H_2O(l)\ ]$

4.103 Divide the mass of $CaCl_2$ by its molar mass and volume to find molarity:

$$4.50 \text{ g } CaCl_2 \times \frac{1 \text{ mol } CaCl_2}{110.98 \text{ g } CaCl_2} \times \frac{1}{1.000 \text{ L soln}} = 0.040\underline{5}4 = 0.0405 \text{ M } CaCl_2$$

The $CaCl_2$ dissolves to form Ca^{2+} and $2Cl^-$ ions. Therefore, the molarities of the ions are 0.0405 M Ca^{2+} and 2 x 0.04054, or 0.0811, M Cl^- ions.

4.104 Divide the mass of $Fe_2(SO_4)_3$ by its molar mass and volume to find molarity:

$$3.45 \text{ g } Fe_2(SO_4)_3 \times \frac{1 \text{ mol } Fe_2(SO_4)_3}{399.91 \text{ g } Fe_2(SO_4)_3} \times \frac{1}{1.000 \text{ L soln}} = 0.0086\underline{2}7$$

$$= 0.00863 \text{ M } Fe_2(SO_4)_3$$

The $Fe_2(SO_4)_3$ dissolves to form $2Fe^{3+}$ and $3SO_4^{2-}$ ions. Therefore, the molarities of the ions are 2 x 0.00863, or 0.0173, M Fe^{3+} and 3 x 0.00863, or 0.0259, M SO_4^{2-}.

4.105 Divide the mass of $K_2Cr_2O_7$ by its molar mass and volume to find molarity. Then calculate the volume needed to prepare 1.00 L of a 0.100 M solution.

$$89.3 \text{ g } K_2Cr_2O_7 \times \frac{1 \text{ mol } K_2Cr_2O_7}{294.20 \text{ g } K_2Cr_2O_7} = 0.30\underline{3}5 \text{ mol } K_2Cr_2O_7$$

$$\text{Molarity} = \frac{0.3035 \text{ mol } K_2Cr_2O_7}{1.00 \text{ L}} = 0.30\underline{3}5 \text{ M}$$

$$V_i = \frac{M_f \times V_f}{M_i} = \frac{0.100 \text{ M} \times 1.00 \text{ L}}{0.3035 \text{ M}} = 0.32\underline{9}4 \text{ L (329 mL)}$$

4.106 Divide the mass by the molar mass and then by volume to find molarity. Then, calculate the volume needed to prepare 1.00 L of 0.1150 M.

$$71.2 \text{ g } H_2C_2O_4 \times \frac{1 \text{ mol } H_2C_2O_4}{90.04 \text{ g } H_2C_2O_4} = 0.79\underline{0}7 \text{ mol } H_2C_2O_4$$

$$\text{Molarity} = \frac{0.7907 \text{ mol } H_2C_2O_4}{1.00 \text{ L}} = 0.79\underline{0}7 = 0.791 \text{ M}$$

$$V_i = \frac{M_f \times V_f}{M_i} = \frac{0.150 \text{ M} \times 1.00 \text{ L}}{0.7907 \text{ M}} = 0.18\underline{9}6 \text{ L (190. mL)}$$

Place 190. mL of the 0.791 M solution in a 1 L volumetric flask, and dilute to 1.00 L.

4.107 Assume a volume of 1.000 L (1000 cm^3) for the 6.00 percent NaBr solution, and convert to moles and then to molarity.

$$1000 \text{ cm}^3 \times \frac{1.046 \text{ g soln}}{1 \text{ cm}^3} \times \frac{6.00 \text{ g NaBr}}{100 \text{ g soln}} \times \frac{1 \text{ mol NaBr}}{102.89 \text{ g NaBr}} = 0.60\underline{9}9 \text{ mol}$$

$$\text{Molarity NaBr} = \frac{0.6099 \text{ mol NaBr}}{1.000 \text{ L}} = 0.60\underline{9}9 = 0.610 \text{ M}$$

4.108 Assume a volume of 1.000 L (1000 mL) for the 4.00 percent NH_3 solution, and convert to moles and then to molarity.

$$1000 \text{ mL} \times \frac{0.979 \text{ g soln}}{1 \text{ mL}} \times \frac{4.00 \text{ g } NH_3}{100 \text{ g soln}} \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} = 2.2\underline{9}9 \text{ mol } NH_3$$

$$\text{Molarity } NH_3 = \frac{2.299 \text{ mol } NH_3}{1.000 \text{ L}} = 2.2\underline{9}9 = 2.30 \text{ M}$$

4.109 First, calculate the moles of $BaCl_2$:

$$1.128 \text{ g } BaSO_4 \times \frac{1 \text{ mol } BaSO_4}{233.40 \text{ g } BaSO_4} \times \frac{1 \text{ mol } BaCl_2}{1 \text{ mol } BaSO_4} = 0.00483\underline{2}9 \text{ mol } BaCl_2$$

Then, calculate the molarity from the moles and volume (0.0500 L):

$$\text{Molarity} = \frac{0.0048329 \text{ mol } BaCl_2}{0.0500 \text{ L}} = 0.096\underline{6}58 = 0.0967 \text{ M}$$

4.110 First, calculate the moles of $CaCl_2$:

$$1.437 \text{ g } CaC_2O_4 \times \frac{1 \text{ mol } CaC_2O_4}{128.10 \text{ g } CaC_2O_4} \times \frac{1 \text{ mol } CaCl_2}{1 \text{ mol } CaC_2O_4}$$

$$= 0.0112\underline{1}77 \text{ mol } CaCl_2$$

Then, calculate the molarity from the moles and volume (0.0500 L):

$$\text{Molarity} = \frac{0.0112177 \text{ mol } CaCl_2}{0.0500 \text{ L}} = 0.22\underline{4}35 = 0.224 \text{ M}$$

4.111 First, calculate the g of thallium(I) sulfate:

$$0.2122 \text{ g TlI} \times \frac{1 \text{ mol TlI}}{331.28 \text{ g TlI}} \times \frac{1 \text{ mol } Tl_2SO_4}{2 \text{ mol TlI}} \times \frac{504.83 \text{ g } Tl_2SO_4}{1 \text{ mol } Tl_2SO_4}$$

$$= 0.161\underline{6}8 \text{ g } Tl_2SO_4$$

Then, calculate the percent Tl_2SO_4 in the rat poison:

$$\text{Percent } Ti_2SO_4 = \frac{0.16168 \text{ g}}{0.7590 \text{ g}} \times 100\% = 21.3\underline{0}1 = 21.30\%$$

4.112 First, calculate the g of $CaCO_3$:

$$0.6332 \text{ g } CaC_2O_4 \times \frac{1 \text{ mol } CaC_2O_4}{128.10 \text{ g } CaC_2O_4} \times \frac{1 \text{ mol } CaCO_3}{1 \text{ mol } CaC_2O_4} \times \frac{100.1 \text{ g } CaCO_3}{1 \text{ mol } CaCO_3}$$

$$= 0.494\underline{7}9 \text{ g } CaCO_3$$

Then, calculate the percent $CaCO_3$ in the antacid:

$$\text{Percent } CaCO_3 = \frac{0.49479 \text{ g}}{0.680 \text{ g}} \times 100\% = 72.\underline{7}64 = 72.8\%$$

4.113 The mass of copper(II) ion and mass of sulfate ion in the 98.77 mg sample is:

$$0.09877 \text{ g} \times 0.3250 = 0.0321\underline{0}0 \text{ g } Cu^{2+} \text{ ion}$$

$$0.11666 \text{ g } BaSO_4 \times \frac{96.07 \text{ g } SO_4^{2-}}{233.40 \text{ g } BaSO_4} = 0.0480\underline{1}9 = 0.04802 \text{ g } SO_4^{2-}$$

The mass of water left $= 0.09877 \text{ g} - (0.03210 \text{ g } Cu^{2+} + 0.04802 \text{ g } SO_4^{2-})$

$= 0.01865 \text{ g } H_2O$

The moles of water left $= 0.01865 \text{ g} \div 18.02 \text{ g/mol} = 1.035 \times 10^{-3}$ mol

The moles of Cu^{2+} $= 0.03210 \text{ g} \div 63.55 \text{ g/mol} = 5.05\underline{1}1 \times 10^{-4}$ mol

The ratio of water to Cu^{2+}, or $CuSO_4$ $= 1.035 \times 10^{-3} \div 5.0511 \times 10^{-4} = 2.05$ or 2

The formula is thus $CuSO_4 \bullet 2H_2O$

4.114 The mass of copper(II) ion and mass of sulfate ion in the 85.42 mg sample is:

$$0.08542 \text{ g} \times 0.2976 = 0.0254\underline{2}1 \text{ g } Cu^{2+} \text{ ion}$$

$$0.09333 \text{ g } BaSO_4 \times \frac{96.07 \text{ g } SO_4^{2-}}{233.40 \text{ g } BaSO_4} = 0.0384\underline{1}6 = 0.03842 \text{ g } SO_4^{2-}$$

The mass of water left $= 0.08542 \text{ g} - (0.025421 \text{ g } Cu^{2+} + 0.038416 \text{ g } SO_4^{2-})$
$= 0.021583 \text{ g } H_2O$

The moles of water left $= 0.021583 \div 18.02 \text{ g/mol} = 1.1977 \times 10^{-3}$ mol

The moles of Cu^{2+} $= 0.025421 \text{ g} \div 63.55 \text{ g/mol} = 4.0002 \times 10^{-4}$ mol

The ratio of water to Cu^{2+}, or $CuSO_4$ $= 1.1977 \times 10^{-3} \div 4.000 \times 10^{-4} = 2.99$ or 3

The formula is $CuSO_4 \bullet 3H_2O$

4.115 For these calculations, the relative numbers of moles of gold and chlorine must be determined. These can be found from the masses of the two elements in the sample:

Total mass = mass of Au + mass of Cl = 328 mg

The mass of chlorine in the precipitated AgCl = the mass of chlorine in the compound of gold and chlorine. The mass of Cl in the 0.464 g of AgCl is

$$0.464 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol Cl}}{1 \text{ mol AgCl}} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}}$$

$$= 0.11\underline{4}76 \text{ g Cl} \ \ (114.76 \text{ mg Cl})$$

$$\text{Mass Percent Cl} = \frac{\text{mass Cl}}{\text{mass comp}} \times 100\% = \frac{114.76 \text{ mg}}{328 \text{ mg}} \times 100\% = 35.0\% \text{ Cl}$$

To find the empirical formula, convert each mass to moles:

$$\text{Mass Au} = 328 \text{ mg} - 11\underline{4}.76 \text{ mg Cl} = 21\underline{3}.23 \text{ mg Au}$$

$$0.11476 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.0032\underline{3}8 \text{ mol Cl}$$

$$0.21323 \text{ g Au} \times \frac{1 \text{ mol Au}}{196.97 \text{ g Au}} = 0.0010\underline{8}2 \text{ mol Au}$$

Divide both numbers of moles by the smaller number (0.001082) to find the integers:

Integer for Cl: 0.003238 mol ÷ 0.001082 mol = 2.99, or 3

Integer for Au: 0.001082 mol ÷ 0.001082 mol = 1.00, or 1

The empirical formula is $AuCl_3$.

4.116 To find the mass percentages of Cl and Sc, find the mass of the chlorine in the precipitated AgCl, which is equal to the mass of the chlorine in the compound of scandium and chlorine. The mass of Cl in the 167.4 mg (0.1674 g) of AgCl is

$$0.1674 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol Cl}}{1 \text{ mol AgCl}} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 0.041\underline{4}1 \text{ g Cl}$$

(continued)

The mass of Sc equals the difference between the sample mass and the mass of Cl:

Mass Sc = 0.0589 g - 0.04141 g = 0.01749 g Sc

The mass percentage of each element is found by dividing the mass of the element by the mass of the compound and multiplying by 100 percent: Mass % = (mass elem. ÷ mass comp.) x 100%.

Mass % Cl = (0.04141 g ÷ 0.0589 g) x 100% = 70.305 = 70.3%

Mass % Sc = (0.01749 g ÷ 0.0589 g) x 100% = 29.69 = 29.7%

To find the empirical formula, convert each mass to moles:

$$0.04141 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.0011680 \text{ mol Cl}$$

$$0.01749 \text{ g Sc} \times \frac{1 \text{ mol Sc}}{44.96 \text{ g Sc}} = 0.0003890 \text{ mol Sc}$$

Divide both numbers of moles by the smaller number (0.0003890) to find the integers:

Integer for Cl = 0.0011680 mol ÷ 0.0003890 mol = 3.00, or 3

Integer for Sc = 0.0003890 mol ÷ 0.0003890 mol = 1.00, or 1

The empirical formula is $ScCl_3$.

4.117 From the equations NH_3 + HCl → NH_4Cl and NaOH + HCl → NaCl + H_2O, we write

Mol NH_3 = mol HCl(NH_3)

Mol NaOH = mol HCl(NaOH)

We can calculate the mol NaOH and the sum [mol HCl(NH_3) + mol HCl(NaOH)] from the titration data. Because the sum equals mol NH_3 plus mol NaOH, we can calculate the unknown mol of NH_3 from the difference: Mol NH_3 = sum - mol NaOH.

$$\text{Mol HCl (NaOH)} + \text{mol HCl (NH}_3\text{)} = 0.0463 \text{ L} \times \frac{0.0213 \text{ mol HCl}}{1.000 \text{ L}}$$

$$= 0.009862 \text{ mol HCl}$$

(continued)

$$\text{Mol NaOH} = 0.0443\ \text{L} \times \frac{0.128\ \text{mol NaOH}}{1.000\ \text{L}} = 0.0056\underline{7}0\ \text{mol NaOH}$$

$$\text{Mol HCl(NH}_3\text{)} = 0.009862\ \text{mol} - 0.005670\ \text{mol} = 0.004192\ \text{mol}$$

$$\text{Mol NH}_3 = \text{mol HCl(NH}_3\text{)} = 0.004192\ \text{mol NH}_3$$

Because all the N in the $(NH_4)_2SO_4$ was liberated as and titrated as NH_3, the amount of N in the fertilizer is equal to the amount of N in the NH_3. Thus, the moles of NH_3 can be used to calculate the mass percentage of N in the fertilizer:

$$0.004192\ \text{mol NH}_3 \times \frac{1\ \text{mol N}}{1\ \text{mol NH}_3} \times \frac{14.01\ \text{g N}}{1\ \text{mol N}} = 0.058\underline{7}3\ \text{g N}$$

$$\text{Percent N} = \frac{\text{mass N}}{\text{mass fert.}} \times 100\% = \frac{0.05873\ \text{g N}}{0.608\ \text{g}} \times 100\% = 9.6\underline{5}9 = 9.66\%$$

4.118 From $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ and $HCl + NaOH \rightarrow NaCl + H_2O$, we write

$$\text{Mol NaHCO}_3 = \text{mol HCl(NaHCO}_3\text{)}$$

$$\text{Mol NaOH} = \text{mol HCl(NaOH)}$$

We can calculate the mol NaOH and the sum [mol HCl($NaHCO_3$) + mol HCl(NaOH)] from the titration data. Because the sum equals mol $NaHCO_3$ plus mol NaOH, we can calculate the unknown mol of $NaHCO_3$ from the difference: Mol $NaHCO_3$ = sum - mol NaOH.

$$\text{Mol HCl (NaOH)} + \text{mol HCl (NaHCO}_3\text{)} = 0.0500\ \text{L} \times \frac{0.190\ \text{mol HCl}}{1.000\ \text{L}}$$

$$= 0.0095\underline{0}0\ \text{mol HCl}$$

$$\text{Mol NaOH} = 0.0471\ \text{L} \times \frac{0.128\ \text{mol NaOH}}{1.000\ \text{L}} = 0.0060\underline{2}9\ \text{mol NaOH}$$

$$\text{Mol HCl(NaHCO}_3\text{)} = 0.009500\ \text{mol} - 0.006029\ \text{mol} = 0.003471\ \text{mol}$$

$$\text{Mol NaHCO}_3 = \text{mol HCl(NaHCO}_3\text{)} = 0.003471\ \text{mol}$$

Because all the $NaHCO_3$ in the tablet was titrated as $NaHCO_3$, the moles of $NaHCO_3$ can be used to calculate the mass percentage of $NaHCO_3$ in the tablet:

(continued)

$$0.003471 \text{ mol NaHCO}_3 \times \frac{84.01 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} = 0.29\underline{1}6 \text{ g NaHCO}_3$$

$$\text{Percent NaHCO}_3 = \frac{\text{mass NaHCO}_3}{\text{mass tab.}} \times 100\% = \frac{0.2916 \text{ g N}}{0.500 \text{ g}} \times 100\%$$

$$= 58.\underline{3}2 = 58.3\%$$

■ Solutions to Cumulative-Skills Problems

4.119 For this reaction, the formulas are listed first, followed by the molecular and net ionic equations, the names of the products, and the molecular equation for another reaction giving the same precipitate.

Lead(II) nitrate is $Pb(NO_3)_2$, and cesium sulfate is Cs_2SO_4.

Molecular equation: $Pb(NO_3)_2(aq) + Cs_2SO_4(aq) \rightarrow PbSO_4(s) + 2CsNO_3(aq)$

Net ionic equation: $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

$PbSO_4$ is lead(II) sulfate, and $CsNO_3$ is cesium nitrate.

Molecular equation: $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$

4.120 For this reaction, the formulas are listed first, followed by the molecular and net ionic equations, the names of the products, and the molecular equation for another reaction giving the same precipitate.

Silver nitrate is $AgNO_3$, and strontium chloride is $SrCl_2$.

Molecular equation: $2AgNO_3(aq) + SrCl_2(aq) \rightarrow 2AgCl(s) + Sr(NO_3)_2(aq)$

Net ionic equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

AgCl is silver(I) chloride, and $Sr(NO_3)_2$ is strontium nitrate.

Molecular equation: $AgClO_4(aq) + NaCl(aq) \rightarrow AgCl(s) + NaClO_4(aq)$

4.121 Net ionic equation: $2Br^-(aq) + Cl_2(g) \rightarrow 2Cl^-(aq) + Br_2(l)$

Molecular equation: $CaBr_2(aq) + Cl_2(g) \rightarrow CaCl_2(aq) + Br_2(l)$

Masses: 40.0 g 14.2 g 22.2 g (40.0 + 14.2 - 22.2) g

Combining the three known masses gives the unknown mass of Br_2. Now, use a ratio of the known masses of $CaBr_2$ to Br_2 to convert pounds of Br_2 to grams of $CaBr_2$:

$$10.0 \text{ lb } Br_2 \times \frac{40.0 \text{ g } CaBr_2}{(40.0 + 14.2 - 22.2) \text{ g } Br_2} \times \frac{453.6 \text{ g}}{1 \text{ lb}}$$

$$= 56\underline{7}0 = 5.67 \times 10^3 \text{ g } CaBr_2$$

4.122 Net ionic equation: $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$

Molecular equation: $BaS(aq) + Na_2CO_3(aq) \rightarrow Na_2S(aq) + BaCO_3(s)$

Masses: 33.9 g 21.2 g 15.6 g (33.9 + 21.2 - 15.6) g

Combining the three known masses gives the unknown mass of $BaCO_3$. Now, use a ratio of the known masses of BaS to $BaCO_3$ to ultimately obtain grams of BaS:

$$10.0 \text{ ton } BaCO_3 \times \frac{33.9 \text{ g BaS}}{(33.9 + 21.2 - 15.62) \text{ g } BaCO_3} = 8.5\underline{8}2 \text{ ton BaS}$$

$$8.5\underline{8}2 \text{ ton BaS} \times \frac{2000 \text{ lb}}{1 \text{ ton}} \times \frac{453.6 \text{ g}}{1 \text{ lb}} = 7.7\underline{8}5 \times 10^6 = 7.79 \times 10^6 \text{ g BaS}$$

4.123 Molecular equation: $Hg(NO_3)_2 + H_2S \rightarrow HgS(s) + 2HNO_3(aq)$

Net ionic equation: $Hg^{2+} + H_2S \rightarrow HgS(s) + 2H^+(aq)$

The acid formed is nitric acid, a strong acid. The other product is mercury(II) sulfide.

Mass HNO_3 = (81.15 g + 8.52 g) - 58.16 g = $31.5\underline{1}$ g

Mass of solution = $550.\underline{0}$ g H_2O + $31.5\underline{1}$ g HNO_3 = $581.\underline{5}1$ = 581.5 g

4.124 Molecular equation: $Hg(NO_3)_2 + H_2S \rightarrow HgS(s) + 2HNO_3(aq)$

Net ionic equation: $Hg^{2+} + H_2S \rightarrow HgS(s) + 2H^+(aq)$

The acid formed is nitric acid, a strong acid. The other product is mercury(II) sulfide.

Mass HNO_3 = (65.65 g + 4.26 g) - 54.16 g = $15.7\underline{5}$ g

Mass of solution = $490.\underline{0}$ g H_2O + $15.7\underline{5}$ g HNO_3 = $505.\underline{7}5$ = 505.8 g

4.125 Let the number of Fe^{3+} ions equals y; then the number of Fe^{2+} ions equals (7 - y). Since Fe_7S_8 is neutral, the number of positive charges must equal the number of negative charges. If the signs are omitted, then:

Total charge on both Fe^{2+} and Fe^{3+} = total charge on all eight sulfide ions

$$3y + 2(7 - y) = 8 \times 2$$

$$y + 14 = 16$$

$$y = 2$$

So the ratio of Fe^{2+} to Fe^{3+} = 5/2

4.126 To define the problem in terms of percentages, use 100 X_2O_3 oxides. Then the number of X^{3+} ions equals 100, and the sum of the X^{2+} and X^{5+} ions also equals 100. Let the number of X^{2+} ions equals y; then the number of X^{5+} ions equals (100 - y). Since X_2O_3 is neutral, the number of positive charges must be equal the number of negative charges. If the signs are omitted, then;

Total charge on X^{2+}, X^{3+}, and X^{5+} = total charge of all 300 oxide ions

$$2y + (100 \times 3) + 5(100 - y) = 300 \times 2$$

$$-3y + 800 = 600$$

$$y = (200 \div 3) = 66.67$$

The percentage of X^{2+} in the 100 X_2O_3 's = (66.67 ÷ 200) x 100 = 33.3%

4.127 Use the density, formula weight, and percentage to convert to molarity. Then, combine the 0.200 mol with mol/L to obtain the volume in liters.

$$\frac{0.807 \text{ g soln}}{1 \text{ mL}} \times \frac{0.940 \text{ g ethanol}}{1.00 \text{ g soln}} \times \frac{1 \text{ mol ethanol}}{46.07 \text{ g ethanol}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= \frac{16.\underline{4}65 \text{ mol ethanol}}{1 \text{ L ethanol}}$$

$$\text{L ethanol} = 0.200 \text{ mol ethanol} \times \frac{1 \text{ L ethanol}}{16.465 \text{ mol ethanol}} = 0.012\underline{1}46 = 0.0121 \text{ L}$$

4.128 Use the density, formula weight, and percentage to convert to molarity. Then, combine the 0.350 mol with mol/L to obtain the volume in liters.

$$\frac{1.072 \text{ g soln}}{1 \text{ mL}} \times \frac{0.560 \text{ g glycol}}{1.00 \text{ g soln}} \times \frac{1 \text{ mol glycol}}{62.07 \text{ g ethanol}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= \frac{9.6\underline{7}2 \text{ mol glycol}}{1 \text{ L glycol}}$$

$$\text{L glycol} = 0.350 \text{ mol glycol} \times \frac{1 \text{ L glycol}}{9.6\underline{7}2 \text{ mol glycol}} = 0.036\underline{1}8 = 0.0362 \text{ L}$$

4.129 Convert the 2.183 g of Ag to mol AgI, which is chemically equivalent to moles of KI. Use that to calculate the molarity of the KI.

$$2.183 \text{ g AgI} \times \frac{1 \text{ mol AgI}}{234.77 \text{ g AgI}} = 9.29\underline{8}4 \times 10^{-3} \text{ mol AgI (eq to mol KI)}$$

$$\text{Molarity} = \frac{9.2984 \times 10^{-3} \text{ mol KI}}{0.0100 \text{ L}} = 0.92\underline{9}8 = 0.930 \text{ M}$$

4.130 Convert the 5.719 g of $BaSO_4$ to mol $BaSO_4$, which is chemically equivalent to moles of Na_2SO_4. Use that to calculate the molarity of the Na_2SO_4.

$$5.719 \text{ g } BaSO_4 \times \frac{1 \text{ mol } BaSO_4}{233.40 \text{ g } BaSO_4} = 0.0245\underline{0}2 \text{ mol } BaSO_4 \text{ (eq. to mol } Na_2SO_4\text{)}$$

$$\text{Molarity} = \frac{0.024502 \text{ mol } Na_2SO_4}{0.0250 \text{ L}} = 0.98\underline{0}1 = 0.980 \text{ M}$$

4.131 Convert the 6.026 g of $BaSO_4$ to mol $BaSO_4$; then, from the equation, deduce that three mol $BaSO_4$ is equivalent to one mol $M_2(SO_4)_3$ and is equivalent to two mol of M. Use that with 1.200 g of the metal M to calculate the atomic weight of M.

$$6.026 \text{ g } BaSO_4 \times \frac{1 \text{ mol } BaSO_4}{233.40 \text{ g } BaSO_4} \times \frac{2 \text{ mol M}}{3 \text{ mol } BaSO_4} = 0.0172\underline{1}2 \text{ mol M}$$

$$\text{Atomic wt of M in g/mol} = \frac{1.200 \text{ g M}}{0.017212 \text{ mol M}} = 69.7\underline{1}9 \text{ g/mol (gallium)}$$

4.132 Convert the 7.964 g of AgCl to mol AgCl; then, from the equation, deduce that two mol AgCl is equivalent to one mol MCl_2 and is equivalent to one mol M. Use that with 2.434 g of the metal M to calculate the atomic weight of M.

$$7.964 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol M}}{2 \text{ mol AgCl}} = 0.0277\underline{8}4 \text{ mol M}$$

$$\text{Atomic wt M} = \frac{2.434 \text{ g M}}{0.027784 \text{ mol M}} = 87.6\underline{0}4 = 87.60 \text{ g/mol (strontium)}$$

4.133 Use the density, formula weight, percentage, and volume to convert to mol H_3PO_4. Then, from the equation $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$, deduce that four mol H_3PO_4 is equivalent to one mol of P_4O_{10}, and use that to convert to mol P_4O_{10}.

$$15\underline{0}0 \text{ mL} \times \frac{1.025 \text{ g soln}}{1 \text{ mL}} \times \frac{0.0500 \text{ g } H_3PO_4}{1 \text{ g soln}} \times \frac{1 \text{ mol } H_3PO_4}{98.00 \text{ g } H_3PO_4}$$

$$= 0.78\underline{4}4 \text{ mol } H_3PO_4$$

$$0.7844 \text{ mol } H_3PO_4 \times \frac{1 \text{ mol } P_4O_{10}}{4 \text{ mol } H_3PO_4} = 0.19\underline{6}1 \text{ mol } P_4O_{10}$$

$$\text{Mass } P_4O_{10} = 0.1961 \text{ mol } P_4O_{10} \times \frac{283.92 \text{ g } P_4O_{10}}{1 \text{ mol } P_4O_{10}}$$

$$= 55.\underline{6}77 = 55.7 \text{ g } P_4O_{10}$$

4.134 Use the density, formula weight, percentage, and volume to convert to mol $FeCl_3$, which is chemically equivalent to mol Fe, from the equation $2Fe + 3Cl_2(g) \rightarrow 2FeCl_3$.

$$\frac{1.067 \text{ g soln}}{1 \text{ mL}} \times \frac{0.0900 \text{ g } FeCl_3}{1 \text{ g soln}} \times \frac{1 \text{ mol } H_3PO_4}{162.20 \text{ g } FeCl_3} \times 2500 \text{ mL}$$

$$= 1.4\underline{8}0 \text{ mol } FeCl_3$$

$$1.480 \text{ mol } FeCl_3 = 1.480 \text{ mol Fe}$$

$$\text{Mass Fe} = 1.480 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 82.\underline{6}6 = 82.7 \text{ g Fe}$$

4.135 Convert the 0.1068 g of hydrogen to mol H_2; then, deduce from the equation that three mol H_2 is equivalent to two mol Al. Use the moles of Al to calculate mass Al and the percentage Al.

$$0.1068 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol Al}}{3 \text{ mol } H_2} = 0.0353\underline{1}74 \text{ mol Al}$$

$$\text{Percent Al} = \frac{0.0353174 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}}}{1.118 \text{ g alloy}} \times 100\% = 85.2\underline{2}9 = 85.23\%$$

4.136 Convert the 0.1352 g of hydrogen to mol H_2; then, deduce from the equation that three mol H_2 is equivalent to two mol Fe. Use the moles of Fe to calculate mass Fe and the percentage Fe.

$$0.1352 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol Fe}}{3 \text{ mol } H_2} = 0.0447\underline{0}8 \text{ mol Fe}$$

$$\text{Percent Fe} = \frac{0.044708 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}}{2.358 \text{ g alloy}} \times 100\% = 105.\underline{8}9 = 105.9\%$$

4.137 Use the formula weight of $Al_2(SO_4)_3$ to convert to mol $Al_2(SO_4)_3$. Then deduce from the equation that one mol $Al_2(SO_4)_3$ is equivalent to three mol H_2SO_4, and calculate the moles of H_2SO_4 needed. Combine density, percentage, and formula weight to obtain molarity of H_2SO_4. Then, combine molarity and moles to obtain volume.

$$37.4 \text{ g } Al_2(SO_4)_3 \times \frac{1 \text{ mol } Al_2(SO_4)_3}{342.17 \text{ g } Al_2(SO_4)_3} \times \frac{3 \text{ mol } H_2SO_4}{1 \text{ mol } Al_2(SO_4)_3}$$

$$= 0.32\underline{7}9 \text{ mol } H_2SO_4$$

$$\frac{1.104 \text{ g soln}}{1 \text{ mL}} \times \frac{15.0 \text{ g } H_2SO_4}{100 \text{ g soln}} \times \frac{1 \text{ mol } H_2SO_4}{98.09 \text{ g } H_2SO_4} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= 1.6\underline{8}8 \text{ mol } H_2SO_4/\text{L}$$

$$0.3279 \text{ mol } H_2SO_4 \times \frac{1 \text{ L } H_2SO_4}{1.688 \text{ mol } H_2SO_4} = 0.19\underline{4}2 = 0.194 \text{ L (194 mL)}$$

4.138 Use the formula weight of Na_3PO_4 to convert to mol Na_3PO_4. Then, deduce from the equation that one mol Na_3PO_4 is equivalent to three mol NaOH, and calculate the moles of NaOH needed. Combine density, percentage, and formula weight to obtain molarity of NaOH. Then, combine molarity and moles to obtain the volume.

$$26.2 \text{ g } Na_3PO_4 \times \frac{1 \text{ mol } Na_3PO_4}{163.94 \text{ g } Na_3PO_4} \times \frac{3 \text{ mol NaOH}}{1 \text{ mol } Na_3PO_4} = 0.47\underline{9}4 \text{ mol NaOH}$$

$$\frac{1.133 \text{ g soln}}{1 \text{ mL}} \times \frac{0.120 \text{ g NaOH}}{1 \text{ g soln}} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= 3.3\underline{9}9 \text{ mol NaOH}$$

$$0.4794 \text{ mol NaOH} \times \frac{1 \text{ L NaOH}}{3.399 \text{ mol NaOH}} = 0.14\underline{1}04 \text{ L (141 mL)}$$

4.139 The equations for the neutralization are $2HCl + Mg(OH)_2 \rightarrow MgCl_2 + 2H_2O$ and $3HCl + Al(OH)_3 \rightarrow AlCl_3 + 3H_2O$. Calculate the moles of HCl, and set equal to the total moles of hydroxide ion, OH^-.

$$0.0485 \text{ L HCl} \times \frac{0.187 \text{ mol HCl}}{1 \text{ L HCl}} = 0.0090\underline{6}95 \text{ mol HCl}$$

$$0.0090695 \text{ mol HCl} = 2\,[\text{mol } Mg(OH)_2] + 3\,[\text{mol } Al(OH)_3]$$

Rearrange the last equation, and solve for the moles of $Al(OH)_3$.

$$\text{(Eq 1)}\quad \text{mol } Al(OH)_3 = 0.0030231 \text{ mol HCl} - 2/3\,[\text{mol } Mg(OH)_2]$$

The total mass of chloride salts is equal to the sum of the masses of $MgCl_2$ (molar mass = 95.21 g/mol) and $AlCl_3$ (molar mass = 133.33 g/mol).

$$[95.21 \text{ g/mol} \times \text{mol } MgCl_2] + [133.33 \text{ g/mol} \times \text{mol } AlCl_3] = 0.4200 \text{ g}$$

Since the mol $Mg(OH)_2$ equals mol $MgCl_2$, and mol $Al(OH)_3$ equals mol $AlCl_3$, you can substitute these quantities into the last equation and get

$$[95.21 \text{ g/mol} \times \text{mol } Mg(OH)_2] + [133.33 \text{ g/mol} \times \text{mol } Al(OH)_3] = 0.4200 \text{ g}$$

Substitute Eq 1 into this equation for the mol of $Al(OH)_3$:

$$[95.21 \text{ g/mol} \times \text{mol } Mg(OH)_2]$$
$$+ [133.33 \text{ g/mol} \times (0.0030231 \text{ mol} - 2/3 \text{ mol } Mg(OH)_2)] = 0.4200 \text{ g}$$

(continued)

Rearrange the equation, and solve for the moles of $Mg(OH)_2$.

$$6.323 \text{ mol } Mg(OH)_2 + 0.4030\underline{7}0 = 0.4200$$

$$\text{mol } Mg(OH)_2 = 0.016\underline{9}3 \div 6.323 = 0.0026\underline{7}28 \text{ mol}$$

Calculate the mass of $Mg(OH)_2$ in the antacid tablet (molar mass = 58.33 g/mol).

$$0.0026728 \text{ mol } Mg(OH)_2 \times \frac{58.33 \text{ g } Mg(OH)_2}{1 \text{ mol } Mg(OH)_2} = 0.15\underline{5}90 \text{ g } Mg(OH)_2$$

Use Eq 1 to find the moles and mass of $Al(OH)_3$ (molar mass 78.00 g/mol) in a similar fashion.

$$\text{Mol } Al(OH)_3 = 0.0030231 - 2/3[0.0026\underline{7}28 \text{ mol } Mg(OH)_2] = 0.0012\underline{4}12 \text{ mol}$$

$$0.0012\underline{4}12 \text{ mol } Al(OH)_3 \times \frac{78.00 \text{ g } Al(OH)_3}{1 \text{ mol } Al(OH)_3} = 0.096\underline{6}81 \text{ g } Al(OH)_3$$

The percent $Mg(OH)_2$ in the antacid is

$$\text{Percent } Mg(OH)_2 = [0.15590 \text{ g} \div (0.15590 + 0.096\underline{6}81) \text{ g}] \times 100\%$$

$$= 61.\underline{7}2 = 61.7\%$$

4.140 The equations for the neutralization are $2HCl + MgCO_3 \rightarrow MgCl_2 + H_2O + CO_2$ and $2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$. Calculate the moles of HCl, and set equal to the total moles of carbonate ion, CO_3^{2-}.

$$0.04133 \text{ L HCl} \times \frac{0.08750 \text{ mol HCl}}{1 \text{ L HCl}} = 0.00361\underline{6}4 \text{ mol HCl}$$

$$0.0036164 \text{ mol HCl} = 2 [\text{mol } CaCO_3] + 2 [\text{mol } MgCO_3]$$

Rearrange this equation, and solve for the moles of $MgCO_3$.

$$\text{(Eq 1)} \quad \text{mol } MgCO_3 = 0.0018082 \text{ mol HCl} - \text{mol } CaCO_3$$

The total mass of chloride salts equals the sum of the masses of $MgCl_2$ (molar mass = 95.21 g/mol) and $CaCl_2$ (molar mass = 110.98 g/mol).

$$[110.98 \text{ g/mol} \times \text{mol } CaCl_2] + [95.21 \text{ g/mol} \times \text{mol } MgCl_2] = 0.1900 \text{ g}$$

(continued)

Since the mol $CaCO_3$ = mol $CaCl_2$, and mol $MgCO_3$ = mol $MgCl_2$, you can substitute these quantities into the last equation and get

$$[110.98 \text{ g/mol} \times \text{mol } CaCO_3] + [95.21 \text{ g/mol} \times \text{mol } MgCO_3] = 0.1900 \text{ g}$$

Substitute Eq 1 into this equation for the mol of $MgCO_3$:

$$[110.98 \text{ g/mol} \times \text{mol } CaCO_3] + [95.21 \text{ g/mol} \times (0.0018082 \text{ mol} - \text{mol } CaCO_3)] = 0.1900 \text{ g}$$

Rearrange the equation and solve for the moles of $MgCO_3$.

$$15.77 \text{ mol } CaCO_3 + 0.172159 = 0.1900$$

$$\text{mol } CaCO_3 = 0.017\underline{8}41 \div 15.77 = 0.0011\underline{3}13 \text{ mol}$$

Calculate the mass of $CaCO_3$ (molar mass = 100.09 g/mol) in the antacid tablet.

$$0.0011313 \text{ mol } CaCO_3 \times \frac{100.09 \text{ g } CaCO_3}{1 \text{ mol } CaCO_3} = 0.11\underline{3}234 \text{ g } CaCO_3$$

Use Eq 1, and find the moles and mass of $MgCO_3$ (molar mass = 84.32 g/mol) in a similar fashion.

$$\text{mol } MgCO_3 = 0.0018082 \text{ mol} - 0.0011313 \text{ mol } CaCO_3 = 0.0006769 \text{ mol}$$

$$0.0006769 \text{ mol } MgCO_3 \times \frac{84.32 \text{ g } MgCO_3}{1 \text{ mol } MgCO_3} = 0.057076 \text{ g } MgCO_3$$

The percent $CaCO_3$ in the antacid is

$$\text{Percent } CaCO_3 = [0.113234 \text{ g} \div (0.113234 + 0.057076) \text{ g}] \times 100\%$$
$$= 66.\underline{4}87 = 66.5\%$$

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