

## REACTIONS IN AQUEOUS SOLUTION

### 4.1 PRECIPITATION REACTIONS

2. Use solubility rules (Figure 4.2).
- (a)  $\text{BaCl}_2$             soluble
- (b)  $\text{Mg}(\text{OH})_2$         insoluble
- (c)  $\text{Cr}_2(\text{CO}_3)_3$       insoluble
- (d)  $\text{K}_3\text{PO}_4$             soluble
4. (a) To precipitate  $\text{Fe}(\text{OH})_3$ , add a soluble hydroxide such as a solution of sodium hydroxide ( $\text{NaOH}$ ).
- (b) To precipitate  $\text{Fe}_2(\text{CO}_3)_3$ , add a soluble carbonate such as a solution of sodium carbonate ( $\text{Na}_2\text{CO}_3$ ).
- (c) To precipitate  $\text{FePO}_4$ , add a soluble phosphate such as a solution of sodium phosphate ( $\text{Na}_3\text{PO}_4$ ).
6. This problem is similar to Example 4.2. Using Figure 4.3 in the textbook as your guide, split up the soluble ionic compounds into their respective ions: these ions are present in solution. From these ions, combine a cation with an anion in solution. Two possible combinations of salts are obtained. Check solubility rules (Figure 4.2) to determine which of the possible salts precipitates.
- (a) Ions present:  $\text{Fe}^{3+}$ ,  $\text{NO}_3^-$ ,  $\text{Na}^+$ ,  $\text{OH}^-$   
 Possible precipitates:  $\text{Fe}(\text{OH})_3$ ,  $\text{NaNO}_3$   
 Solubility:  $\text{Fe}(\text{OH})_3$  is insoluble;  $\text{NaNO}_3$  is soluble and remains in solution.  
 Net ionic equation:  $\text{Fe}^{3+}(\text{aq}) + 3\text{OH}^-(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s})$
- (b) Ions present:  $\text{Mg}^{2+}$ ,  $\text{SO}_4^{2-}$ ,  $\text{Ba}^{2+}$ ,  $\text{OH}^-$   
 Possible precipitates:  $\text{Mg}(\text{OH})_2$ ,  $\text{BaSO}_4$   
 Solubility:  $\text{Mg}(\text{OH})_2$  and  $\text{BaSO}_4$  are both insoluble and would both precipitate.  
 Net ionic equations:  $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$   
 $\text{Mg}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Mg}(\text{OH})_2(\text{s})$   
 The net ionic equation can also be written as:  
 $\text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{Mg}(\text{OH})_2(\text{s})$

8. This problem is similar to Example 4.2. Using Figure 4.3 in the textbook as your guide, split up the soluble ionic compounds into their respective ions: these ions are present in solution. From these ions, combine a cation with an anion in solution. Two possible combinations of salts are obtained. Check solubility rules (Figure 4.2) to determine which of the possible salts precipitates.

(a) Ions present:  $\text{Ag}^+$ ,  $\text{NO}_3^-$ ,  $\text{Na}^+$ ,  $\text{Cl}^-$

Possible precipitates:  $\text{AgCl}$ ,  $\text{NaNO}_3$

Solubility:  $\text{NaNO}_3$  is soluble and remains in solution;  $\text{AgCl}$  is insoluble and will precipitate.

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

(b) Ions present:  $\text{Co}^{2+}$ ,  $\text{NO}_3^-$ ,  $\text{Na}^+$ ,  $\text{OH}^-$

Possible precipitates:  $\text{Co}(\text{OH})_2$ ,  $\text{NaNO}_3$

Solubility:  $\text{NaNO}_3$  is soluble and remains in solution;  $\text{Co}(\text{OH})_2$  is insoluble and will precipitate.

Net ionic equation:  $\text{Co}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Co}(\text{OH})_2(\text{s})$

(c) Ions present:  $\text{NH}_4^+$ ,  $\text{PO}_4^{3-}$ ,  $\text{K}^+$ ,  $\text{OH}^-$ ,

Possible precipitates:  $\text{NH}_4\text{OH}$ ,  $\text{K}_3\text{PO}_4$

Solubility: **No reaction.** Both compounds are soluble, thus no precipitate forms.

(d) Ions present:  $\text{Cu}^{2+}$ ,  $\text{SO}_4^{2-}$ ,  $\text{Na}^+$ ,  $\text{CO}_3^{2-}$

Possible precipitates:  $\text{CuCO}_3$ ,  $\text{Na}_2\text{SO}_4$

Solubility:  $\text{Na}_2\text{SO}_4$  is soluble and remains in solution;  $\text{CuCO}_3$  is insoluble and will precipitate.

Net ionic equation:  $\text{Cu}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CuCO}_3(\text{s})$

(e) Ions present:  $\text{Li}^+$ ,  $\text{SO}_4^{2-}$ ,  $\text{Ba}^{2+}$ ,  $\text{OH}^-$

Possible precipitates:  $\text{LiOH}$ ,  $\text{BaSO}_4$

Solubility:  $\text{LiOH}$  is soluble and remains in solution;  $\text{BaSO}_4$  is insoluble and will precipitate.

Net ionic equation:  $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$

10. This problem is similar to Example 4.2. Using Figure 4.3 in the textbook as your guide, split up the soluble ionic compounds into their respective ions: these ions are present in solution. From these ions, combine a cation with an anion in solution. Two possible combinations of salts are obtained. Check solubility rules (Figure 4.2) to determine which of the possible salts precipitates.

(a) Ions present:  $\text{Na}^+$ ,  $\text{PO}_4^{3-}$ ,  $\text{Ba}^{2+}$ ,  $\text{Cl}^-$

Possible precipitates:  $\text{Ba}_3(\text{PO}_4)_2$ ,  $\text{NaCl}$

Solubility:  $\text{NaCl}$  is soluble and remains in solution;  $\text{Ba}_3(\text{PO}_4)_2$  is insoluble and will precipitate.

Net ionic equation:  $3\text{Ba}^{2+}(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq}) \rightarrow \text{Ba}_3(\text{PO}_4)_2(\text{s})$

(b) Ions present:  $\text{Zn}^{2+}$ ,  $\text{SO}_4^{2-}$ ,  $\text{K}^+$ ,  $\text{OH}^-$

Possible precipitates:  $\text{Zn}(\text{OH})_2$ ,  $\text{K}_2\text{SO}_4$

Solubility:  $\text{K}_2\text{SO}_4$  is soluble and remains in solution;  $\text{Zn}(\text{OH})_2$  is insoluble and will precipitate.

Net ionic equation:  $\text{Zn}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s})$

(c) Ions present:  $\text{NH}_4^+$ ,  $\text{SO}_4^{2-}$ ,  $\text{Na}^+$ ,  $\text{Cl}^-$

Possible precipitates:  $\text{NH}_4\text{Cl}$ ,  $\text{Na}_2\text{SO}_4$

Solubility: **No reaction.** Both compounds are soluble, so no precipitate forms.

(d) Ions present:  $\text{Co}^{3+}$ ,  $\text{NO}_3^-$ ,  $\text{Na}^+$ ,  $\text{PO}_4^{3-}$

Possible precipitates:  $\text{CoPO}_4$ ,  $\text{NaNO}_3$

Solubility:  $\text{NaNO}_3$  is soluble and remains in solution;  $\text{CoPO}_4$  is insoluble and will precipitate.

Net ionic equation:  $\text{Co}^{3+}(\text{aq}) + \text{PO}_4^{3-}(\text{aq}) \rightarrow \text{CoPO}_4(\text{s})$

#### 4.1 PRECIPITATION REACTIONS: Stoichiometry

12. These problems are similar to Example 4.3.

(a) Strategy (From Figure 4.6): V of parent compound,  $\text{Na}_2\text{CO}_3 \rightarrow$  mol of parent compound,  $\text{Na}_2\text{CO}_3 \rightarrow$  mol of reacting ion,  $\text{CO}_3^{2-} \rightarrow$  mol of other reacting ion,  $\text{Fe}^{3+} \rightarrow$  mol of parent compound,  $\text{Fe}(\text{NO}_3)_3 \rightarrow$  M of solution of parent compound,  $\text{Fe}(\text{NO}_3)_3$

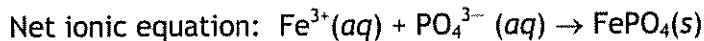
Net ionic equation:  $2\text{Fe}^{3+}(\text{aq}) + 3\text{CO}_3^{2-}(\text{aq}) \rightarrow \text{Fe}_2(\text{CO}_3)_3(\text{s})$

$$\begin{aligned} \text{mol Fe}(\text{NO}_3)_3 &= 12.54 \text{ mL Na}_2\text{CO}_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.1488 \text{ mol Na}_2\text{CO}_3}{1 \text{ L Na}_2\text{CO}_3} \times \frac{1 \text{ mol CO}_3^{2-}}{1 \text{ mol Na}_2\text{CO}_3} \\ &\times \frac{2 \text{ mol Fe}^{3+}}{3 \text{ mol CO}_3^{2-}} \times \frac{1 \text{ mol Fe}(\text{NO}_3)_3}{1 \text{ mol Fe}^{3+}} = 0.001244 \text{ mol Fe}(\text{NO}_3)_3 \end{aligned}$$

The molarity of the  $\text{Fe}(\text{NO}_3)_3$  solution is:

$$M = \frac{\text{moles}}{\text{volume (L)}} = \frac{0.001244 \text{ mol Fe}(\text{NO}_3)_3}{0.02500 \text{ L}} = \underline{0.04976 \text{ M}}$$

- (b) Strategy (From Figure 4.6): mass of parent compound,  $K_3PO_4 \rightarrow$  mol of parent compound,  $K_3PO_4 \rightarrow$  mol of reacting ion,  $PO_4^{3-} \rightarrow$  mol of other reacting ion,  $Fe^{3+} \rightarrow$  mol of parent compound,  $Fe(NO_3)_3 \rightarrow$  M of parent compound,  $Fe(NO_3)_3$

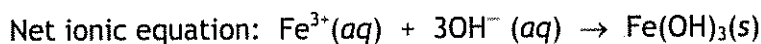


$$\begin{aligned} \text{mol } Fe(NO_3)_3 &= 7.58 \text{ g } K_3PO_4 \times \frac{1 \text{ mol } K_3PO_4}{212.27 \text{ g } K_3PO_4} \times \frac{1 \text{ mol } PO_4^{3-}}{1 \text{ mol } K_3PO_4} \\ &\times \frac{1 \text{ mol } Fe^{3+}}{1 \text{ mol } PO_4^{3-}} \times \frac{1 \text{ mol } Fe(NO_3)_3}{1 \text{ mol } Fe^{3+}} = 0.03571 \text{ mol } Fe(NO_3)_3 \end{aligned}$$

The molarity of the  $Fe(NO_3)_3$  solution is:

$$M = \frac{\text{moles}}{\text{volume (L)}} = \frac{0.03571 \text{ mol } Fe(NO_3)_3}{0.02500 \text{ L}} = \underline{1.43 \text{ M}}$$

- (c) Strategy (From Figure 4.6): V of parent compound,  $Sr(OH)_2 \rightarrow$  mol of parent compound,  $Sr(OH)_2 \rightarrow$  mol of reacting ion,  $OH^- \rightarrow$  mol of other reacting ion,  $Fe^{3+} \rightarrow$  mol of parent compound,  $Fe(NO_3)_3 \rightarrow$  M of parent compound,  $Fe(NO_3)_3$

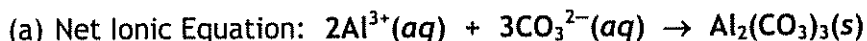


$$\begin{aligned} \text{mol } Fe(NO_3)_3 &= 10.00 \text{ mL } Sr(OH)_2 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.1573 \text{ mol } Sr(OH)_2}{1 \text{ L } Sr(OH)_2} \times \frac{2 \text{ mol } OH^-}{1 \text{ mol } Sr(OH)_2} \\ &\times \frac{1 \text{ mol } Fe^{3+}}{3 \text{ mol } OH^-} \times \frac{1 \text{ mol } Fe(NO_3)_3}{1 \text{ mol } Fe^{3+}} = 0.001049 \text{ mol } Fe(NO_3)_3 \end{aligned}$$

The molarity of the  $Fe(NO_3)_3$  solution is:

$$M = \frac{\text{moles}}{\text{volume (L)}} = \frac{0.001049 \text{ mol } Fe(NO_3)_3}{0.02500 \text{ L}} = \underline{0.04196 \text{ M}}$$

14. These problems are similar to Example 4.3.



- (b) Strategy (From Figure 4.6): V of parent compound,  $Na_2CO_3 \rightarrow$  mol of parent compound,  $Na_2CO_3 \rightarrow$  mol of reacting ion,  $CO_3^{2-} \rightarrow$  mol of other reacting ion,  $Al^{3+} \rightarrow$  mol of parent compound,  $AlCl_3 \rightarrow$  M of  $AlCl_3$  solution

$$\begin{aligned} \text{mol AlCl}_3 &= 35.5 \text{ mL Na}_2\text{CO}_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.137 \text{ mol Na}_2\text{CO}_3}{1 \text{ L Na}_2\text{CO}_3} \times \frac{1 \text{ mol CO}_3^{2-}}{1 \text{ mol Na}_2\text{CO}_3} \\ &\times \frac{2 \text{ mol Al}^{3+}}{3 \text{ mol CO}_3^{2-}} \times \frac{1 \text{ mol AlCl}_3}{1 \text{ mol Al}^{3+}} = 0.00324 \text{ mol AlCl}_3 \end{aligned}$$

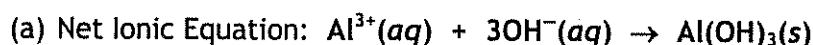
The volume of AlCl<sub>3</sub> required is 30.0 mL or 0.0300 L

$$M = \frac{\text{moles}}{\text{volume (L)}} = \frac{0.00324 \text{ mol AlCl}_3}{0.0300 \text{ L}} = \underline{0.108 \text{ M}}$$

- (c) Strategy (From Figure 4.6): V of parent compound, Na<sub>2</sub>CO<sub>3</sub> → mol of parent compound, Na<sub>2</sub>CO<sub>3</sub> → mol of reacting ion, CO<sub>3</sub><sup>2-</sup> → mol of precipitate, Al<sub>2</sub>(CO<sub>3</sub>)<sub>3</sub> → mass of precipitate, Al<sub>2</sub>(CO<sub>3</sub>)<sub>3</sub>

$$\begin{aligned} \text{mass Al}_2(\text{CO}_3)_3 &= 35.5 \text{ mL Na}_2\text{CO}_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.137 \text{ mol Na}_2\text{CO}_3}{1 \text{ L Na}_2\text{CO}_3} \times \frac{1 \text{ mol CO}_3^{2-}}{1 \text{ mol Na}_2\text{CO}_3} \\ &\times \frac{1 \text{ mol Al}_2(\text{CO}_3)_3}{3 \text{ mol CO}_3^{2-}} \times \frac{233.99 \text{ g Al}_2(\text{CO}_3)_3}{1 \text{ mol Al}_2(\text{CO}_3)_3} = \underline{0.379 \text{ g Al}_2(\text{CO}_3)_3} \end{aligned}$$

16. This is a limiting reactant problem. A review of Example 4.3c will help you.



- (b) To find the mass of precipitate formed, determine first the limiting reactant by identifying the reactant that will produce fewer moles of precipitate.

mol precipitate if Al<sup>3+</sup> is the limiting reactant:

Strategy (from Figure 4.6): mass of parent compound, Al<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub> → mol of parent compound, Al<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub> → mol of reacting ion, Al<sup>3+</sup> → mol of precipitate, Al(OH)<sub>3</sub>

$$\begin{aligned} 2.76 \text{ g Al}_2(\text{SO}_4)_3 &\times \frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{342.17 \text{ g Al}_2(\text{SO}_4)_3} \times \frac{2 \text{ mol Al}^{3+}}{1 \text{ mol Al}_2(\text{SO}_4)_3} \times \frac{1 \text{ mol Al}(\text{OH})_3}{1 \text{ mol Al}^{3+}} \\ &= 0.0161 \text{ mol Al}(\text{OH})_3 \end{aligned}$$

mol precipitate if OH<sup>-</sup> is the limiting reactant:

Strategy (from Figure 4.6): V of parent compound, NaOH → mol of parent compound, NaOH → mol of reacting ion, OH<sup>-</sup> → mol of precipitate, Al(OH)<sub>3</sub>

$$\begin{aligned} 85.0 \text{ mL NaOH} &\times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.2500 \text{ mol NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol OH}^{-}}{1 \text{ mol NaOH}} \times \frac{1 \text{ mol Al}(\text{OH})_3}{3 \text{ mol OH}^{-}} \\ &= 0.00708 \text{ mol Al}(\text{OH})_3 \end{aligned}$$

Since 0.00708 mol  $\text{Al}(\text{OH})_3$  is less than 0.0161 mol  $\text{Al}(\text{OH})_3$ ,  $\text{OH}^-$  is the limiting reactant. Hence, 0.00708 mol of precipitate forms. The mass of this precipitate is

$$0.00708 \text{ mol Al}(\text{OH})_3 \times \frac{78.00 \text{ g Al}(\text{OH})_3}{1 \text{ mol Al}(\text{OH})_3} = \underline{0.552 \text{ g Al}(\text{OH})_3}$$

(c) Since  $\text{Al}^{3+}$  will form more precipitate,  $\text{Al}^{3+}$  is the ion in excess.

$$\begin{aligned} \text{mol Al}^{3+} \text{ originally available: } & 2.76 \text{ g Al}_2(\text{SO}_4)_3 \times \frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{342.17 \text{ g Al}_2(\text{SO}_4)_3} \times \frac{2 \text{ mol Al}^{3+}}{1 \text{ mol Al}_2(\text{SO}_4)_3} \\ & = 0.0161 \text{ mol Al}^{3+} \text{ available} \end{aligned}$$

$$\begin{aligned} \text{mol Al}^{3+} \text{ reacted: } & 85.0 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.2500 \text{ mol NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} \times \frac{1 \text{ mol Al}^{3+}}{3 \text{ mol OH}^-} \\ & = 0.00708 \text{ mol Al}^{3+} \text{ reacted or consumed} \end{aligned}$$

$$\begin{aligned} \text{excess Al}^{3+} &= 0.0161 \text{ mol Al}^{3+} \text{ available} - 0.00708 \text{ mol Al}^{3+} \text{ consumed} \\ &= 0.0090 \text{ mol Al}^{3+} \end{aligned}$$

Assuming volume is additive, total solution volume is 0.235 L or 235 mL (125 mL + 85.0 mL = 210 mL)

$$\text{Molarity} = \frac{0.0090 \text{ mol Al}^{3+}}{0.210 \text{ L}} = \underline{0.043 \text{ M}}$$

## 4.2 ACID-BASE REACTIONS

18. Acids and bases NOT listed in Table 4.1 are weak.

(a)  $\text{H}_2\text{S}$  is a weak acid

(c)  $\text{C}_5\text{H}_5\text{N}$  is a weak base

(b)  $\text{H}_2\text{SO}_4$  is a strong acid

(d)  $\text{Al}(\text{OH})_3$  is a weak base

20. Acids NOT listed in Table 4.1 are weak. As shown in Figure 4.10, for strong acids the reacting species is  $\text{H}^+$  while for weak acids the reacting species is the acid molecule.

(a) Hypochlorous acid,  $\text{HClO}$  is a weak acid; the reacting species is  $\text{HClO}$ .

(b) Formic acid,  $\text{HCHO}_2$  is a weak acid; the reacting species is  $\text{HCHO}_2$ .

(c) Acetic acid,  $\text{HC}_2\text{H}_3\text{O}_2$  is a weak acid; the reacting species is  $\text{HC}_2\text{H}_3\text{O}_2$ .

(d) Hydrobromic acid,  $\text{HBr}$  is a strong acid; the reacting species is  $\text{H}^+$ .

(e) Sulfurous acid,  $\text{H}_2\text{SO}_3$  is a weak acid; the reacting species is  $\text{H}_2\text{SO}_3$ .

22. As shown in Table 4.1, the hydroxides of Group I and Group II metals are the only strong bases. As shown in Figure 4.10, for strong bases the reacting species is  $\text{OH}^-$  while for weak bases the reacting species is the base molecule.
- (a) Toluidine,  $\text{C}_7\text{H}_9\text{N}$  is a weak base; the reacting species is  $\text{C}_7\text{H}_9\text{N}$ .
- (b) Strontium hydroxide,  $\text{Sr}(\text{OH})_2$  is a strong base; the reacting species is  $\text{OH}^-$ .
- (c) Indol,  $\text{C}_8\text{H}_6\text{NH}$  is a weak base; the reacting species is  $\text{C}_8\text{H}_6\text{NH}$ .
- (d) Aqueous ammonia,  $\text{NH}_3$  is a weak base; the reacting species is  $\text{NH}_3$ .
24. This problem is similar to Example 4.4. Use Tables 4.1 and 4.2 and Figure 4.10 as guide. Classify the acid and the base as weak or strong, and then identify the reacting species. Write the reaction for the two reacting species.
- (a) Acetic acid,  $\text{HC}_2\text{H}_3\text{O}_2$  is a weak acid; the reacting species is  $\text{HC}_2\text{H}_3\text{O}_2$ .
- Strontium hydroxide,  $\text{Sr}(\text{OH})_2$  is a strong base; the reacting species is  $\text{OH}^-$ .
- Net Ionic Equation:  $\text{HC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O} + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$
- (b) Diethylamine,  $(\text{C}_2\text{H}_5)_2\text{NH}$  is a weak base; the reacting species is  $(\text{C}_2\text{H}_5)_2\text{NH}$ .
- Sulfuric acid,  $\text{H}_2\text{SO}_4$  is a strong acid; the reacting species is  $\text{H}^+$ .
- Net Ionic Equation:  $(\text{C}_2\text{H}_5)_2\text{NH}(\text{aq}) + \text{H}^+(\text{aq}) \rightarrow (\text{C}_2\text{H}_5)_2\text{NH}_2^+(\text{aq})$
- (c) Hydrofluoric acid (HF) is a weak acid; the reacting species is HF.
- Sodium hydroxide (NaOH) is a strong base; the reacting species is  $\text{OH}^-$ .
- Net Ionic Equation:  $\text{HF}(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O} + \text{F}^-(\text{aq})$
26. The following prototype net ionic equation,  $\text{OH}^-(\text{aq}) + \text{HB}(\text{aq}) \rightarrow \text{B}^-(\text{aq}) + \text{H}_2\text{O}$ , would only be correct when the reactants are a weak acid and a strong base. Use Tables 4.1 and 4.2 and Figure 4.10 as guide.
- (a) The equation is **NOT CORRECT**.
- Hydrochloric acid (HCl) is a strong acid; the reacting species is  $\text{H}^+$ .
- Pyridine ( $\text{C}_5\text{H}_5\text{N}$ ) is a weak base; the reacting species is  $\text{C}_5\text{H}_5\text{N}$ .
- Correct equation:  $\text{H}^+(\text{aq}) + \text{C}_5\text{H}_5\text{N}(\text{aq}) \rightarrow \text{C}_5\text{H}_5\text{NH}^+(\text{aq})$
- (b) The equation is **NOT CORRECT**.
- Sulfuric acid ( $\text{H}_2\text{SO}_4$ ) is a strong acid; the reacting species is  $\text{H}^+$ .
- Rubidium hydroxide (RbOH) is a strong base; the reacting species is  $\text{OH}^-$ .
- Correct Equation:  $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}$

(c) The equation is **CORRECT**.

Hydrofluoric acid (HF) is a weak acid; the reacting species is HF.

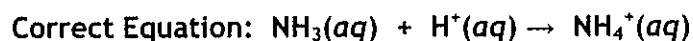
Potassium hydroxide (KOH) is a strong base; the reacting species is OH<sup>-</sup>.



(d) The equation is **NOT CORRECT**.

Hydroiodic acid (HI) is a strong acid; the reacting species is H<sup>+</sup>.

Ammonia (NH<sub>3</sub>) is a weak base; the reacting species is NH<sub>3</sub>.



(e) The equation is **CORRECT**.

Hydrocyanic acid (HCN) is a weak acid; the reacting species is HCN.

Strontium hydroxide (Sr(OH)<sub>2</sub>) is a strong base; the reacting species is OH<sup>-</sup>.



## 4.2 ACID-BASE REACTIONS: Acid-Base Titrations

28. This problem is similar to Example 4.5. Use the flowchart for solution stoichiometry (Figure 4.6) as a guide.

Strategy: V and M NH<sub>3</sub> → mol NH<sub>3</sub> → mol H<sup>+</sup> → mol H<sub>2</sub>SO<sub>4</sub> → V H<sub>2</sub>SO<sub>4</sub>

The stoichiometric ratio of reacting species is 1 mol NH<sub>3</sub> to 1 mol H<sup>+</sup>.

Convert mol H<sup>+</sup> to mol H<sub>2</sub>SO<sub>4</sub> (the ratio is 1:2). Finally, use molarity of H<sub>2</sub>SO<sub>4</sub> to calculate the volume of H<sub>2</sub>SO<sub>4</sub>.

Net ionic equation: NH<sub>3</sub>(aq) + H<sup>+</sup>(aq) → NH<sub>4</sub><sup>+</sup>(aq)

$$\begin{aligned} \text{mol H}_2\text{SO}_4 &: 38.00 \text{ mL NH}_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.189 \text{ mol NH}_3}{1 \text{ L}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol NH}_3} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol H}^+} \\ &= 0.00359 \text{ mol H}_2\text{SO}_4 \end{aligned}$$

volume (mL) of H<sub>2</sub>SO<sub>4</sub> required:

$$V = n \div M = \frac{0.00359 \text{ mol}}{0.2315 \text{ M}} = \underline{0.0155 \text{ L or } 15.5 \text{ mL H}_2\text{SO}_4 \text{ solution}}$$

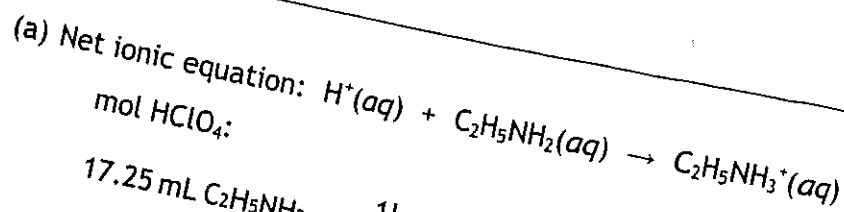
30. These problems are similar to Example 4.5. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide.

Strategy: mol base → mol H<sup>+</sup> → mol HClO<sub>4</sub> → M HClO<sub>4</sub> solution

Calculate mol of the given base; use the mole ratio of reacting base to calculate mol H<sup>+</sup>.

Convert mol H<sup>+</sup> to mol HClO<sub>4</sub> then use the volume of HClO<sub>4</sub> to calculate its molarity.



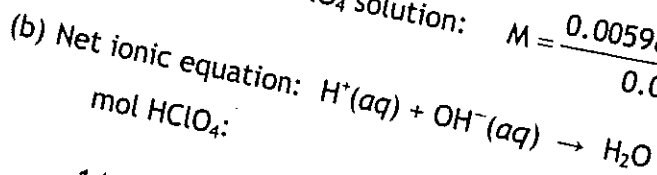


mol  $HClO_4$ :

$$17.25 \text{ mL } C_2H_5NH_2 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.3471 \text{ mol } C_2H_5NH_2}{1 \text{ L}} \times \frac{1 \text{ mol } H^+}{1 \text{ mol } C_2H_5NH_2} \times \frac{1 \text{ mol } HClO_4}{1 \text{ mol } H^+}$$

$$= 0.005987 \text{ mol } HClO_4$$

molarity of  $HClO_4$  solution:  $M = \frac{0.005987 \text{ mol } HClO_4}{0.02500 \text{ L}} = \underline{0.2395 \text{ M}}$



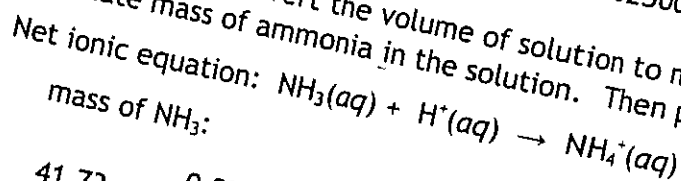
mol  $HClO_4$ :

$$14.17 \text{ g } Sr(OH)_2 \times \frac{1 \text{ mol } Sr(OH)_2}{121.64 \text{ g } Sr(OH)_2} \times \frac{2 \text{ mol } OH^-}{1 \text{ mol } Sr(OH)_2} \times \frac{1 \text{ mol } H^+}{1 \text{ mol } OH^-} \times \frac{1 \text{ mol } HClO_4}{1 \text{ mol } H^+}$$

$$= 0.2330 \text{ mol } HClO_4$$

molarity of  $HClO_4$  solution:  $M = \frac{0.2330 \text{ mol } HClO_4}{0.02500 \text{ L}} = \underline{9.320 \text{ M}}$

(c) Use density to convert the volume of solution to mass then use mass percent to calculate mass of ammonia in the solution. Then proceed as above.



mass of  $NH_3$ :

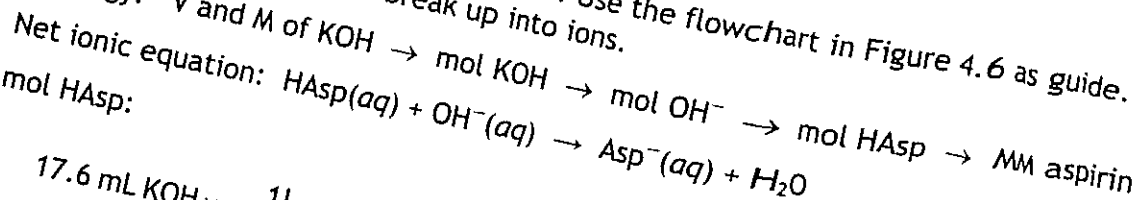
$$41.73 \text{ mL} \times \frac{0.9295 \text{ g solution}}{1 \text{ mL}} \times \frac{18 \text{ g } NH_3}{100 \text{ g solution}} \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{1 \text{ mol } H^+}{1 \text{ mol } NH_3} \times \frac{1 \text{ mol } HClO_4}{1 \text{ mol } H^+}$$

$$= 0.41 \text{ mol } HClO_4$$

molarity of  $HClO_4$  solution:  $M = \frac{0.41 \text{ mol } HClO_4}{0.02500 \text{ L}} = \underline{16 \text{ M}}$

32. This problem is similar to Example 4.5c. Use the flowchart in Figure 4.6 as guide. HAsp is a weak acid so it does not break up into ions.

Strategy:  $V$  and  $M$  of  $KOH \rightarrow$  mol  $KOH \rightarrow$  mol  $OH^- \rightarrow$  mol HAsp  $\rightarrow$  MM aspirin



mol HAsp:

$$17.6 \text{ mL } KOH \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.315 \text{ mol } KOH}{1 \text{ L}} \times \frac{1 \text{ mol } OH^-}{1 \text{ mol } KOH} \times \frac{1 \text{ mol } HAsp}{1 \text{ mol } OH^-} = 0.00554 \text{ mol HAsp}$$

Chapter 4

Molar mass of aspirin:

$$MM_{\text{aspirin}} = \text{mass} \div n = \frac{1.00 \text{ g HAsp}}{0.00554 \text{ mol HAsp}} = 1.80 \times 10^2 \text{ g/mol}$$

34. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide.  $\text{HC}_2\text{H}_3\text{O}_2$  is a weak acid and does not break up into ions.

Strategy: V and M of  $\text{Ba}(\text{OH})_2 \rightarrow \text{mol Ba}(\text{OH})_2 \rightarrow \text{mol OH}^- \rightarrow \text{mol HC}_2\text{H}_3\text{O}_2 \rightarrow \text{mass HC}_2\text{H}_3\text{O}_2$

Net ionic equation:  $\text{OH}^-(\text{aq}) + \text{HC}_2\text{H}_3\text{O}_2(\text{aq}) \rightarrow \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + \text{H}_2\text{O}$

mass  $\text{HC}_2\text{H}_3\text{O}_2$ :

$$37.50 \text{ mL Ba}(\text{OH})_2 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.1250 \text{ mol Ba}(\text{OH})_2}{1 \text{ L Ba}(\text{OH})_2} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ba}(\text{OH})_2}$$

$$\times \frac{1 \text{ mol HC}_2\text{H}_3\text{O}_2}{1 \text{ mol OH}^-} \times \frac{60.053 \text{ g HC}_2\text{H}_3\text{O}_2}{1 \text{ mol HC}_2\text{H}_3\text{O}_2} = 0.5630 \text{ g HC}_2\text{H}_3\text{O}_2$$

$$\begin{aligned} \% \text{ HC}_2\text{H}_3\text{O}_2 &= \frac{\text{mass of HC}_2\text{H}_3\text{O}_2}{\text{mass of sample}} \times 100 \\ &= \frac{0.5630 \text{ g HC}_2\text{H}_3\text{O}_2}{10.00 \text{ g sample}} \times 100 \\ &= \underline{5.630\%} \end{aligned}$$

Yes, the sample can be considered vinegar because percent acetic acid is at least 5.0%.

36. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide. Note that  $\text{C}_6\text{H}_8\text{O}_6$  is a weak acid and does not break up into ions.

Strategy: V and M  $\text{KOH} \rightarrow \text{mol KOH} \rightarrow \text{mol OH}^- \rightarrow \text{mol C}_6\text{H}_8\text{O}_6 \rightarrow \text{mass C}_6\text{H}_8\text{O}_6$

Net ionic equation:  $\text{C}_6\text{H}_8\text{O}_6(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{C}_6\text{H}_7\text{O}_6^-(\text{aq}) + \text{H}_2\text{O}$

mass vitamin C,  $\text{C}_6\text{H}_8\text{O}_6$ :

$$5.94 \text{ mL KOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.450 \text{ mol KOH}}{1 \text{ L KOH}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol KOH}} \times \frac{1 \text{ mol C}_6\text{H}_8\text{O}_6}{1 \text{ mol OH}^-} \times \frac{176.12 \text{ g C}_6\text{H}_8\text{O}_6}{1 \text{ mol C}_6\text{H}_8\text{O}_6}$$

$$= 0.471 \text{ g C}_6\text{H}_8\text{O}_6 \text{ (vitamin C)}$$

$$\% \text{ C}_6\text{H}_8\text{O}_6 = \frac{\text{mass of C}_6\text{H}_8\text{O}_6}{\text{mass of sample}} \times 100 = \frac{0.471 \text{ g C}_6\text{H}_8\text{O}_6}{0.634 \text{ g sample}} \times 100 = \underline{74.3\%}$$

38. To find the number of moles of  $\text{OH}^-$  required to neutralize 1 mole of lactic acid ( $\text{C}_3\text{H}_6\text{O}_3$ ), calculate the number of moles of hydroxide ions and lactic acid. Determine the mole ratio of lactic acid to NaOH. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide. Note that  $\text{C}_3\text{H}_6\text{O}_3$  is a weak acid and does not break up into ions.

Strategy: mass  $\text{C}_3\text{H}_6\text{O}_3 \rightarrow \text{mol C}_3\text{H}_6\text{O}_3$   
 $V \text{ NaOH} \rightarrow \text{mol NaOH} \rightarrow \text{mol OH}^-$

$$\text{mol C}_3\text{H}_6\text{O}_3: 0.100 \text{ g C}_3\text{H}_6\text{O}_3 \times \frac{1 \text{ mol C}_3\text{H}_6\text{O}_3}{90.08 \text{ g C}_3\text{H}_6\text{O}_3} = 1.11 \times 10^{-3} \text{ mol C}_3\text{H}_6\text{O}_3$$

$$\text{mol OH}^-: 12.95 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.0857 \text{ mol NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} = 1.11 \times 10^{-3} \text{ mol OH}^-$$

Since  $1.11 \times 10^{-3} \text{ mol OH}^-$  is required to neutralize  $1.11 \times 10^{-3} \text{ mol}$  lactic acid, then 1 mole  $\text{OH}^-$  is required to neutralize 1 mole lactic acid.

### 4.3 OXIDATION-REDUCTION REACTIONS: Oxidation Number

40. This problem is similar to Example 4.6. Apply the rules for oxidation numbers.

- (a)  $\text{CH}_4$  rule #4: oxidation no. H = +1  
 rule #5: oxidation no. C:  $x + 4(+1) = 0$ ; solve for  $x \Rightarrow x = -4$ ;  
 oxidation no. C = -4
- (b)  $\text{CO}_3^{2-}$  rule #6: oxidation no. O = -2  
 rule #5: oxidation no. C:  $x + 3(-2) = -2$ ; solve for  $x \Rightarrow x = +4$ ;  
 oxidation no. C = +4
- (c)  $\text{IO}_4^-$  rule #6: oxidation no. O = -2  
 rule #5: oxidation no. I:  $x + 4(-2) = -1$ ; solve for  $x \Rightarrow x = +7$ ;  
 oxidation no. I = +7
- (d)  $\text{N}_2\text{H}_4$  rule #4: oxidation no. H = +1  
 rule #5: oxidation no. N:  $x(2) + 4(+1) = 0$ ; solve for  $x \Rightarrow x = -2$ ;  
 oxidation no. N = -2

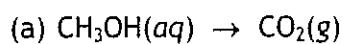
42. These problems are similar to Example 4.6. Apply the rules for oxidation numbers.

- (a)  $\text{HIO}_3$  rule #4: oxidation no. H = +1  
 rule #5: oxidation no. I:  $1(+1) + x + 3(-2) = 0$ ; solve for  $x \Rightarrow x = +5$ ;  
 oxidation no. I = +5  
 rule #6: oxidation no. O = -2

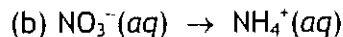
- (b)  $\text{NaMnO}_4$  rule #3: oxidation no. Na = +1  
 rule #5: oxidation no. Mn:  $1(+1) + x + 4(-2) = 0$ ; solve for  $x \Rightarrow x = +7$ ;  
 oxidation no. Mn = +7  
 rule #6: oxidation no. O = -2
- (c)  $\text{SnO}_2$  rule #5: oxidation no. Sn:  $x + 2(-2) = 0$ ; solve for  $x \Rightarrow x = +4$ ;  
 oxidation no. Sn = +4  
 rule #6: oxidation no. O = -2
- (d)  $\text{NOF}$  rule #5: oxidation no. N:  $x + 1(-2) + 1(-1) = 0$ ; solve for  $x \Rightarrow x = +3$ ;  
 oxidation no. N = +3  
 rule #6: oxidation no. O = -2  
 rule #3: oxidation no. F = -1
- (e)  $\text{NaO}_2$  rule #3: oxidation no. Na = +1  
 rule #5: oxidation no. O:  $1(+1) + 2x = 0$ ; solve for  $x \Rightarrow x = -\frac{1}{2}$ ;  
 oxidation no. O =  $-\frac{1}{2}$

### 4.3 OXIDATION-REDUCTION REACTIONS: Balancing Half-Equations

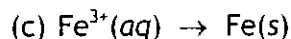
44. To know if a half-equation is an oxidation or a reduction reaction, determine the change in oxidation number. The half-equation is an oxidation reaction if there is an increase in oxidation number; otherwise it is a reduction reaction.



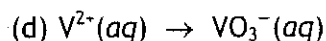
C: oxidation number *increased* from -2 to +4, so this is oxidation.



N: oxidation number *decreased* from +5 to -3, so this is reduction.

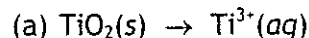


Fe: oxidation number *decreased* from +3 to 0, so this is reduction.

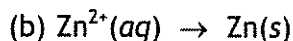
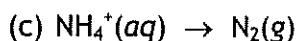
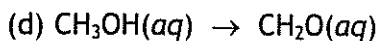


V: oxidation number *increased* from +2 to +5, so this is oxidation.

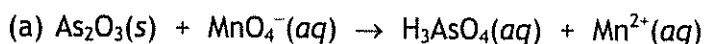
46. As described in #44, the half-reaction is oxidation if there is an increase in the oxidation number in any of the atoms; otherwise it is reduction.



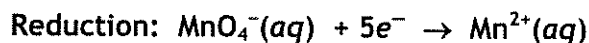
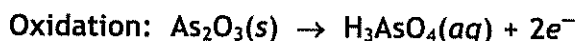
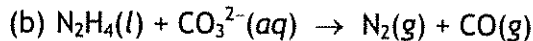
Ti: oxidation number *decreased* from +4 to +3, so this is reduction.

Zn: oxidation number *decreased* from +2 to 0, so this is **reduction**.N: oxidation number *increased* from -3 to 0, so this is **oxidation**.C: oxidation number *increased* from -2 to 0, so this is **oxidation**.

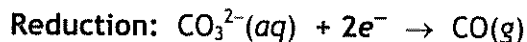
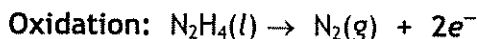
48. Write unbalanced half-reactions. Assign oxidation number to each element. Identify the element oxidized and reduced. If the oxidation number of an element increased, then that element is oxidized. If the oxidation number of an element decreased, then that element is reduced. The species or reactant (ion or molecule) that has the oxidized atom (lost electrons) is called the reducing agent. On the other hand, the species or reactant (ion or molecule) that has the reduced atom (accepted or gained electrons) is called the oxidizing agent.



unbalanced half-reactions:

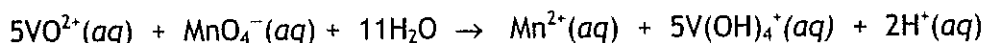
**Mn is reduced:** oxidation number *decreased* from +7 to +2 $\text{MnO}_4^-$  is the species reduced;  $\text{MnO}_4^-$  is the oxidizing agent**As is oxidized:** oxidation number *increased* from +3 to +5 $\text{As}_2\text{O}_3$  is the species oxidized;  $\text{As}_2\text{O}_3$  is the reducing agent

unbalanced half-reactions:

**C is reduced:** oxidation number *decreased* from +4 to +2 $\text{CO}_3^{2-}$  is the species reduced;  $\text{CO}_3^{2-}$  is the oxidizing agent**N is oxidized:** oxidation number *increased* from -2 to 0 $\text{N}_2\text{H}_4$  is the species oxidized;  $\text{N}_2\text{H}_4$  is the reducing agent

## 4.3 OXIDATION-REDUCTION REACTIONS: Stoichiometry

50. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide to determine the mass of vanadium. Use the stoichiometric ratio from the following balanced equation in the calculation.

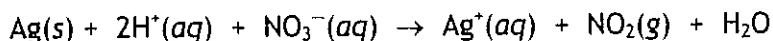


Strategy: V and M  $\text{KMnO}_4 \rightarrow \text{mol KMnO}_4 \rightarrow \text{mol MnO}_4^{-} \rightarrow \text{mol VO}^{2+} \rightarrow \text{mol V} \rightarrow \text{mass V}$

$$\begin{aligned} \text{mass vanadium} &= 26.45 \text{ mL KMnO}_4 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.02250 \text{ mol KMnO}_4}{1 \text{ L}} \times \frac{1 \text{ mol MnO}_4^{-}}{1 \text{ mol KMnO}_4} \\ &\quad \times \frac{5 \text{ mol VO}^{2+}}{1 \text{ mol MnO}_4^{-}} \times \frac{1 \text{ mol V}}{1 \text{ mol VO}^{2+}} \times \frac{50.94 \text{ g V}}{1 \text{ mol V}} = 0.1516 \text{ g vanadium} \end{aligned}$$

$$\% \text{ vanadium in the ore} = \frac{\text{mass of vanadium}}{\text{mass of sample}} \times 100 = \frac{0.1516 \text{ g}}{0.5000 \text{ g}} \times 100 = \underline{\underline{30.32\%}}$$

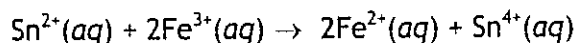
52. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide to determine the mass of silver. The stoichiometric ratio is obtained from the following balanced equation.



Strategy: V and M  $\text{HNO}_3 \rightarrow \text{mol HNO}_3 \rightarrow \text{mol H}^{+} \rightarrow \text{mol Ag} \rightarrow \text{mass Ag}$

$$\begin{aligned} \text{mass Ag} &= 42.50 \text{ mL HNO}_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{12.0 \text{ mol HNO}_3}{1 \text{ L}} \times \frac{1 \text{ mol H}^{+}}{1 \text{ mol HNO}_3} \times \frac{1 \text{ mol Ag}}{2 \text{ mol H}^{+}} \times \frac{107.9 \text{ g}}{1 \text{ mol Ag}} \\ &= \underline{\underline{27.5 \text{ g Ag}}} \end{aligned}$$

54. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide using the stoichiometric ratio from the following balanced equation.



Strategy: mass sample  $\rightarrow \text{mass Fe} \rightarrow \text{mol Fe} \rightarrow \text{mol Sn}^{2+} \rightarrow \text{mol SnCl}_2 \rightarrow \text{M SnCl}_2$

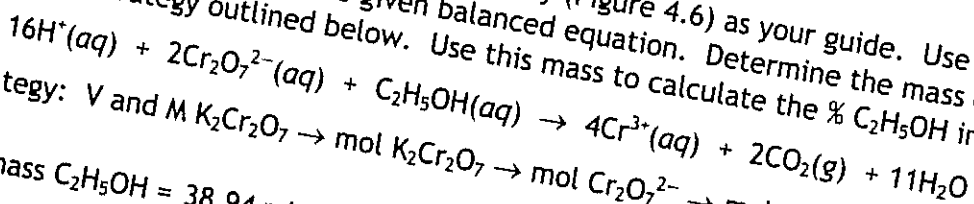
$$\text{mass of Fe} = 0.250 \text{ g sample} \times \frac{92.50 \text{ g Fe}}{100 \text{ g sample}} = 0.231 \text{ g Fe}$$

$$\begin{aligned} \text{mol SnCl}_2 &= 0.231 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}^{3+}}{1 \text{ mol Fe}} \times \frac{1 \text{ mol Sn}^{2+}}{2 \text{ mol Fe}^{3+}} \times \frac{1 \text{ mol SnCl}_2}{1 \text{ mol Sn}^{2+}} \\ &= 2.07 \times 10^{-3} \text{ mol SnCl}_2 \end{aligned}$$

$$\text{vol of solution} = 22.0 \text{ mL} = 0.022 \text{ L}$$

$$M \text{ of SnCl}_2 \text{ solution} = \frac{2.07 \times 10^{-3} \text{ mol SnCl}_2}{0.022 \text{ L solution}} = \underline{0.0941 \text{ M}}$$

56. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide. Use the stoichiometric ratio from the given balanced equation. Determine the mass of  $\text{C}_2\text{H}_5\text{OH}$  using the strategy outlined below. Use this mass to calculate the %  $\text{C}_2\text{H}_5\text{OH}$  in blood.



Strategy: V and M  $\text{K}_2\text{Cr}_2\text{O}_7 \rightarrow \text{mol K}_2\text{Cr}_2\text{O}_7 \rightarrow \text{mol Cr}_2\text{O}_7^{2-} \rightarrow \text{mol C}_2\text{H}_5\text{OH} \rightarrow \text{mass C}_2\text{H}_5\text{OH}$

$$\text{mass C}_2\text{H}_5\text{OH} = 38.94 \text{ mL K}_2\text{Cr}_2\text{O}_7 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.0723 \text{ mol K}_2\text{Cr}_2\text{O}_7}{1 \text{ L}} \times \frac{1 \text{ mol Cr}_2\text{O}_7^{2-}}{1 \text{ mol K}_2\text{Cr}_2\text{O}_7}$$

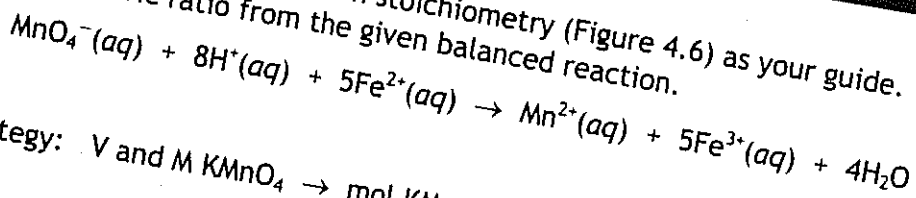
$$\times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{2 \text{ mol Cr}_2\text{O}_7^{2-}} \times \frac{46.07 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 0.0649 \text{ g C}_2\text{H}_5\text{OH}$$

$$\text{mass \%} = \frac{0.0649 \text{ g C}_2\text{H}_5\text{OH}}{50.0 \text{ g blood}} \times 100 = \underline{0.130 \%}$$

**Yes, the person is legally drunk** because the %  $\text{C}_2\text{H}_5\text{OH}$  is more than 0.10% by mass in the blood sample.

### UNCLASSIFIED PROBLEMS

58. Use the flowchart for solution stoichiometry (Figure 4.6) as your guide. Use the stoichiometric ratio from the given balanced reaction.



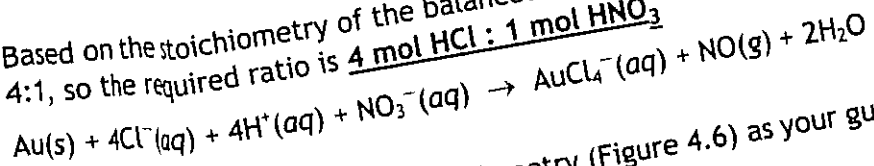
Strategy: V and M  $\text{KMnO}_4 \rightarrow \text{mol KMnO}_4 \rightarrow \text{mol MnO}_4^- \rightarrow \text{mol Fe}^{2+} \rightarrow \text{mass Fe}^{2+}$

$$\text{mass Fe}^{2+} = 32.3 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.00210 \text{ mol KMnO}_4}{1 \text{ L}} \times \frac{1 \text{ mol MnO}_4^-}{1 \text{ mol KMnO}_4}$$

$$\times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} \times \frac{1 \text{ mol Fe}}{1 \text{ mol Fe}^{2+}} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 0.0189 \text{ g Fe}$$

$$\text{mass \%} = \frac{0.0189 \text{ g Fe}}{5.00 \text{ g hemoglobin}} \times 100 = \underline{0.378 \%}$$

60. (a) Based on the stoichiometry of the balanced equation below, the ratio of  $\text{Cl}^-$  to  $\text{NO}_3^-$  is 4:1, so the required ratio is 4 mol HCl : 1 mol  $\text{HNO}_3$



- (b) Use the flowchart for solution stoichiometry (Figure 4.6) as your guide and the stoichiometric ratio from the balanced reaction above.

Strategy: mass Au  $\rightarrow$  mol Au  $\rightarrow$  mol  $\text{Cl}^-$   $\rightarrow$  mol HCl  $\rightarrow$  vol HCl  
 mass Au  $\rightarrow$  mol Au  $\rightarrow$  mol  $\text{NO}_3^-$   $\rightarrow$  mol  $\text{HNO}_3$   $\rightarrow$  vol  $\text{HNO}_3$

$$\text{vol HCl} = 25.0 \text{ g Au} \times \frac{1 \text{ mol Au}}{197 \text{ g Au}} \times \frac{4 \text{ mol Cl}^-}{1 \text{ mol Au}} \times \frac{1 \text{ mol HCl}}{1 \text{ mol Cl}^-} \times \frac{1 \text{ L HCl}}{12 \text{ mol HCl}}$$

$$= 0.042 \text{ L HCl or } \underline{42 \text{ mL HCl}}$$

$$\text{vol HNO}_3 = 25.0 \text{ g Au} \times \frac{1 \text{ mol Au}}{197 \text{ g Au}} \times \frac{1 \text{ mol NO}_3^-}{1 \text{ mol Au}} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol NO}_3^-} \times \frac{1 \text{ L HNO}_3}{16 \text{ mol HNO}_3}$$

$$= 0.0079 \text{ L HNO}_3 \text{ or } \underline{7.9 \text{ mL HNO}_3}$$

62. Balanced net ionic equation (given):  
 $\text{H}_3\text{PO}_4(aq) + 3\text{OH}^-(aq) \rightarrow 3\text{H}_2\text{O} + \text{PO}_4^{3-}(aq)$

Use the flowchart for solution stoichiometry (Figure 4.6) as your guide to calculate for the mass of NaOH that has reacted and subtract this calculated mass from the mass of NaOH originally available (20.0 g).

mass of NaOH that has reacted with  $\text{H}_3\text{PO}_4$

Strategy: mL  $\text{H}_3\text{PO}_4$  solution  $\rightarrow$  mass  $\text{H}_3\text{PO}_4$  solution  $\rightarrow$  mass  $\text{H}_3\text{PO}_4$  (pure)  
 mass  $\text{H}_3\text{PO}_4$   $\rightarrow$  mol  $\text{H}_3\text{PO}_4$   $\rightarrow$  mol  $\text{OH}^-$   $\rightarrow$  mol NaOH  $\rightarrow$  mass NaOH

$$\text{mass NaOH} = 10 \text{ mL H}_3\text{PO}_4 \text{ solution} \times \frac{1.69 \text{ g H}_3\text{PO}_4 \text{ solution}}{1 \text{ mL H}_3\text{PO}_4 \text{ solution}} \times \frac{91.7 \text{ g H}_3\text{PO}_4}{100 \text{ g H}_3\text{PO}_4 \text{ solution}}$$

$$\times \frac{1 \text{ mol H}_3\text{PO}_4}{97.99 \text{ g H}_3\text{PO}_4} \times \frac{3 \text{ mol OH}^-}{1 \text{ mol H}_3\text{PO}_4} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol OH}^-} \times \frac{40.02 \text{ g NaOH}}{1 \text{ mol NaOH}} = 19.0 \text{ g NaOH}$$

$$\text{mass of NaOH left unreacted} = 20.0 \text{ g NaOH} - 19.0 \text{ g NaOH} = \underline{1.0 \text{ g NaOH unreacted}}$$



## CONCEPTUAL QUESTIONS

64. (a) Type: **SA/WB**Hydrochloric acid (HCl) is a strong acid; the reacting species is  $H^+$ .Ethylamine ( $CH_3CH_2NH_2$ ) is a weak base; the reacting species is  $CH_3CH_2NH_2$ .Net Ionic Equation:  $H^+(aq) + CH_3CH_2NH_2(aq) \rightarrow CH_3CH_2NH_3^+(aq)$ (b) Type: **WA/SB**

Hydrofluoric acid (HF) is a weak acid; the reacting species is HF.

Calcium hydroxide ( $Ca(OH)_2$ ) is a strong base; the reacting species is  $OH^-$ .Net Ionic Equation:  $HF(aq) + OH^-(aq) \rightarrow F^-(aq) + H_2O$ (c) Type: **PPT**Net Ionic Equation:  $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$ (d) Type: **PPT**Net Ionic Equation:  $2Ag^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2Cl^-(aq) \rightarrow BaSO_4(s) + 2AgCl(s)$ (e) **NR** $Mg(NO_3)_2(aq) + NaCl(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + Na^+(aq) + Cl^-(aq)$ 

All ions remain in solution and do not combine to form insoluble salts.

66. (a) This picture represents no precipitation, thus equation (1) matches.

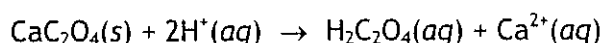
**Equation (1)** involves soluble ions. $2Na^+(aq) + SO_4^{2-}(aq) \rightarrow$  no reaction(b) This picture represents a precipitation reaction where cations and anions combine to form a solid compound. The ratio of cations to anions is 1:1. It matches **equation (3)** because this reaction forms an insoluble compound with a cation to anion ratio of 1:1. $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$ (c) This picture represents a precipitation reaction where cations and anions combine to form a solid compound. The ratio of cations to anions is 1:2. It matches **equation (2)** because this equation forms an insoluble compound with a cation to anion ratio of 1:2. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$ 

68. (a) **weak electrolyte.** The figure shows that only some of the species dissociate, thus it is a weak electrolyte.
- (b) **nonelectrolyte.** The figure shows that none of the species dissociate, thus it is a nonelectrolyte.
- (c) **strong electrolyte.** The figure shows that all of the species dissociate, thus it is a strong electrolyte.
- (d) **weak electrolyte.** The figure shows that most of the species did not dissociate, thus it is a weak electrolyte.

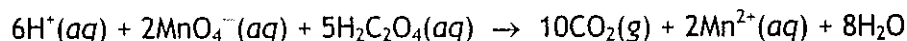
70. (a) **False**. The calculated molarity is the same because a change in the amount of water will not change the number of moles of HCl.
- (b) **True**. Using Ba(OH)<sub>2</sub> instead of NaOH will result in a different volume of base added during the titration, but the number of moles of HCl will be the same.
- (c) **False**. The number of moles of HCl is the same.
- (d) **False**. The number of moles of HCl is the same.
- (e) **True**.

### CHALLENGE PROBLEMS

71. The calcium oxalate is dissolved in strong acid following the following reaction:



The H<sub>2</sub>C<sub>2</sub>O<sub>4</sub> formed is reacted with MnO<sub>4</sub><sup>-</sup> following the balanced reaction below:



Following the flowchart for solution stoichiometry (Figure 4.6) and the strategy below, the amount of CaC<sub>2</sub>O<sub>4</sub> present in the urine sample can be calculated. The stoichiometric ratios in the two balanced reactions above are used in the calculation.

Strategy: M and V KMnO<sub>4</sub> → mol KMnO<sub>4</sub> → mol MnO<sub>4</sub><sup>-</sup> → mol H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>  
 mol H<sub>2</sub>C<sub>2</sub>O<sub>4</sub> → mol CaC<sub>2</sub>O<sub>4</sub> → mass CaC<sub>2</sub>O<sub>4</sub>

$$\begin{aligned} \text{mass CaC}_2\text{O}_4 &= 26.2 \text{ mL KMnO}_4 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.0946 \text{ mol KMnO}_4}{1 \text{ L}} \times \frac{1 \text{ mol MnO}_4^-}{1 \text{ mol KMnO}_4} \\ &\quad \times \frac{5 \text{ mol H}_2\text{C}_2\text{O}_4}{2 \text{ mol MnO}_4^-} \times \frac{1 \text{ mol CaC}_2\text{O}_4}{1 \text{ mol H}_2\text{C}_2\text{O}_4} \times \frac{128.10 \text{ g CaC}_2\text{O}_4}{1 \text{ mol CaC}_2\text{O}_4} = \underline{0.794 \text{ g CaC}_2\text{O}_4} \end{aligned}$$

The urine sample contains **0.794 g CaC<sub>2</sub>O<sub>4</sub>**.

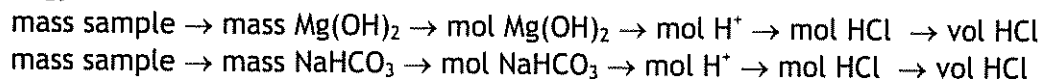
Now, calculate how much Ca<sup>2+</sup> is present in the urine sample.

$$0.794 \text{ g CaC}_2\text{O}_4 \times \frac{1 \text{ mol CaC}_2\text{O}_4}{128.10 \text{ g CaC}_2\text{O}_4} \times \frac{1 \text{ mole Ca}^{2+}}{1 \text{ mol CaC}_2\text{O}_4} \times \frac{40.08 \text{ g Ca}^{2+}}{1 \text{ mole Ca}^{2+}} = \underline{0.248 \text{ g Ca}^{2+}}$$

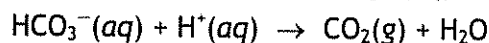
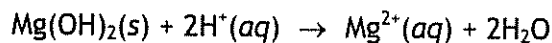
The urine sample contains 0.248 g Ca<sup>2+</sup> or 248 mg Ca<sup>2+</sup>. **Yes, this amount is within the normal range.**

72. Determine the amount of  $\text{Mg}(\text{OH})_2$  and  $\text{NaHCO}_3$  present in the antacid tablet and calculate the volume of acid each could neutralize. Add the two acid volumes. (The  $\text{NaCl}$  in the tablet cannot neutralize any acid.)

Strategy:



The balanced net ionic equations are:



The tablet contains 41.0%  $\text{Mg}(\text{OH})_2$  and 36.2%  $\text{NaHCO}_3$

$$41.0\% \text{ Mg}(\text{OH})_2 = \frac{41.0 \text{ g Mg}(\text{OH})_2}{100 \text{ g tablet}} \quad 36.2\% \text{ NaHCO}_3 = \frac{36.2 \text{ g NaHCO}_3}{100 \text{ g tablet}}$$

Volume of stomach acid (HCl) neutralized by the  $\text{Mg}(\text{OH})_2$  in the tablet:

$$\begin{aligned} 330 \text{ mg tablet} &\times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{41.0 \text{ g Mg}(\text{OH})_2}{100 \text{ g tablet}} \times \frac{1 \text{ mol Mg}(\text{OH})_2}{58.32 \text{ g Mg}(\text{OH})_2} \times \frac{2 \text{ mol H}^+}{1 \text{ mol Mg}(\text{OH})_2} \\ &\times \frac{1 \text{ mol HCl}}{1 \text{ mol H}^+} \times \frac{1 \text{ L HCl}}{0.020 \text{ mol HCl}} = 0.23 \text{ L HCl} \end{aligned}$$

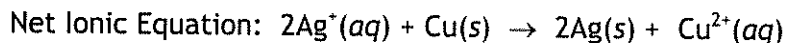
Volume of stomach acid (HCl) neutralized by the  $\text{NaHCO}_3$  in the tablet:

$$\begin{aligned} 330 \text{ mg tablet} &\times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{36.2 \text{ g NaHCO}_3}{100 \text{ g tablet}} \times \frac{1 \text{ mol NaHCO}_3}{84.00 \text{ g NaHCO}_3} \times \frac{1 \text{ mol HCO}_3^-}{1 \text{ mol NaHCO}_3} \\ &\times \frac{1 \text{ mol H}^+}{1 \text{ mol HCO}_3^-} \times \frac{1 \text{ mol HCl}}{1 \text{ mol H}^+} \times \frac{1 \text{ L HCl}}{0.020 \text{ mol HCl}} = 0.071 \text{ L HCl} \end{aligned}$$

Total volume of stomach acid (HCl) neutralized by the tablet:

$$\text{volume} = 0.23 \text{ L} + 0.071 \text{ L} = \underline{0.30 \text{ L}}$$

73. Strategy: mass Cu lost  $\rightarrow$  mol Cu lost  $\rightarrow$  mol Ag coat  $\rightarrow$  mass Ag coat



original mass of Cu = 2.00 g

mass of Cu remaining = 2.00 g - Y (where Y = mass of Cu lost)

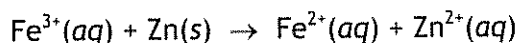
$$\text{mass of Ag coat} = Y \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Cu}} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 3.40Y$$

mass of coated strip = mass of Cu remaining + mass of Ag coat = 4.18 g

$$\begin{aligned}
 \text{substitution yields:} \quad (2.00 \text{ g} - Y) + (3.40Y) &= 4.18 \text{ g} \\
 2.00 \text{ g} - Y + 3.40Y &= 4.18 \text{ g} \\
 -Y + 3.40Y &= 4.18 \text{ g} - 2.00 \text{ g} \\
 2.40Y &= 2.18 \text{ g} \\
 Y &= (2.18/2.40) \text{ g} = 0.908 \text{ g}
 \end{aligned}$$

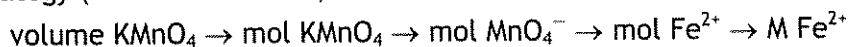
$$\begin{aligned}
 \text{Therefore: mass of Cu in strip} &= 2.00 - 0.908 = \underline{1.09 \text{ g Cu}} \\
 \text{mass of Ag coat in strip} &= 3.40Y = 3.40(0.908) = \underline{3.09 \text{ g Ag}}
 \end{aligned}$$

74. Calculate moles of  $\text{Fe}^{2+}$  from the first permanganate titration; then calculate the total moles of Fe (both  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$ ) from the second titration. The second titration gives the total moles Fe because all  $\text{Fe}^{3+}$  ions present are converted to  $\text{Fe}^{2+}$  according to the following reaction:

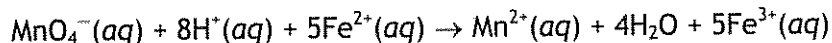


Finally, calculate moles of  $\text{Fe}^{3+}$  from the difference in moles of  $\text{Fe}^{2+}$  obtained from the first and second titrations.

Strategy (for each titration):



Balanced net ionic equation: (step-by-step method for balancing is shown in problem #68 above)



First titration (mol  $\text{Fe}^{2+}$  originally present):

$$35.0 \text{ mL KMnO}_4 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.0280 \text{ mol KMnO}_4}{1 \text{ L}} \times \frac{1 \text{ mol MnO}_4^-}{1 \text{ mol KMnO}_4} \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} = 0.00490 \text{ mol Fe}^{2+}$$

Second titration (total mol of Fe present) :

$$48.0 \text{ mL KMnO}_4 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.0280 \text{ mol KMnO}_4}{1 \text{ L}} \times \frac{1 \text{ mol MnO}_4^-}{1 \text{ mol KMnO}_4} \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} = 0.00672 \text{ mol Fe}^{2+}$$

$$\text{mol Fe}^{3+} = \text{total mol Fe} - \text{mol Fe}^{2+} = 0.00672 - 0.00490 = 0.00182 \text{ mol Fe}^{3+}$$

$$\text{concentration of Fe}^{2+} = \frac{0.00490 \text{ mol Fe}^{2+}}{0.05000 \text{ L}} = \underline{0.0980 \text{ M Fe}^{2+}}$$

$$\text{concentration of Fe}^{3+} = \frac{0.00182 \text{ mol Fe}^{3+}}{0.05000 \text{ L}} = \underline{0.0364 \text{ M Fe}^{3+}}$$

## 75. Balanced Net Ionic Equations:



Strategy (for each acid):

mass acid  $\rightarrow$  mol acid  $\rightarrow$  mol  $\text{OH}^-$   $\rightarrow$  mol NaOH  $\rightarrow$  V NaOH

volume of 0.615 M NaOH required to neutralize 0.0930 g oxalic acid:

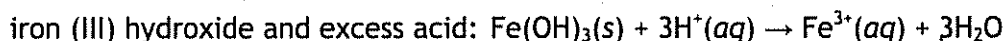
$$0.930\text{g H}_2\text{C}_2\text{O}_4 \times \frac{1\text{mol H}_2\text{C}_2\text{O}_4}{90.04\text{g H}_2\text{C}_2\text{O}_4} \times \frac{2\text{mol OH}^-}{1\text{mol H}_2\text{C}_2\text{O}_4} \times \frac{1\text{mol NaOH}}{1\text{mol OH}^-} \times \frac{1\text{L NaOH}}{0.615\text{mol NaOH}} = 0.0336\text{ L NaOH}$$

volume of 0.615 M NaOH required to neutralize 0.0930 g citric acid:

$$0.930\text{g H}_3\text{C}_6\text{H}_5\text{O}_7 \times \frac{1\text{mol H}_3\text{C}_6\text{H}_5\text{O}_7}{192.12\text{g H}_3\text{C}_6\text{H}_5\text{O}_7} \times \frac{3\text{mol OH}^-}{1\text{mol H}_3\text{C}_6\text{H}_5\text{O}_7} \times \frac{1\text{mol NaOH}}{1\text{mol OH}^-} \times \frac{1\text{L NaOH}}{0.615\text{mol NaOH}} \\ = 0.0236\text{ L NaOH}$$

The unknown is oxalic acid because complete neutralization of the unknown acid required 0.0336 L NaOH or 33.6 mL of 0.615M NaOH, which is also the amount required for 0.930 g oxalic acid.

## 76. Balanced Net Ionic Equations:



total mol  $\text{H}^+$  (added to  $\text{Fe}(\text{OH})_3$ ):

$$625\text{ mL HCl} \times \frac{1\text{L}}{1000\text{ mL}} \times \frac{0.280\text{ mol HCl}}{1\text{L}} \times \frac{1\text{mol H}^+}{1\text{mol HCl}} = 0.175\text{ mol H}^+$$

mol  $\text{H}^+$  in excess (titrated with NaOH):

$$238.2\text{ mL NaOH} \times \frac{1\text{L}}{1000\text{ mL}} \times \frac{0.113\text{ mol NaOH}}{1\text{L}} \times \frac{1\text{mol OH}^-}{1\text{mol NaOH}} \times \frac{1\text{mol H}^+}{1\text{mol OH}^-} = 0.0269\text{ mol H}^+$$

Hence, mol of  $\text{H}^+$  that reacted with  $\text{Fe}(\text{OH})_3$

$$= \text{total mol H}^+ - \text{mol H}^+ \text{ in excess}$$

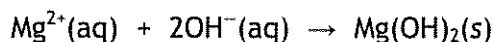
$$= 0.175\text{ mol H}^+ - 0.0269\text{ mol H}^+ = 0.148\text{ mol H}^+$$

mass of iron (III) hydroxide added originally

$$= 0.148\text{ mol H}^+ \times \frac{1\text{mole Fe}(\text{OH})_3}{3\text{mol H}^+} \times \frac{106.87\text{g Fe}(\text{OH})_3}{1\text{mol Fe}(\text{OH})_3} = \underline{5.27\text{ g Fe}(\text{OH})_3}$$

77. mass of  $\text{MgCl}_2$ :

Among the ions in solution ( $\text{Na}^+$ ,  $\text{NO}_3^-$ ,  $\text{Mg}^{2+}$ ,  $\text{Cl}^-$ , and  $\text{Ba}^{2+}$ ) only  $\text{Mg}^{2+}$  forms a precipitate when reacted with  $\text{KOH}$  (see Figure 4.2) as shown by the reaction below.



The mass of the precipitate ( $\text{Mg}(\text{OH})_2$ ) from 100.0 g sample is given as 13.47 g which can be converted to the mass of  $\text{MgCl}_2$ .

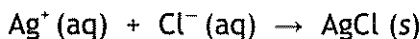
$$13.47 \text{ g Mg}(\text{OH})_2 \times \frac{1 \text{ mol Mg}(\text{OH})_2}{58.316 \text{ g Mg}(\text{OH})_2} \times \frac{1 \text{ mol Mg}^{2+}}{1 \text{ mol Mg}(\text{OH})_2} \times \frac{1 \text{ mol MgCl}_2}{1 \text{ mol Mg}^{2+}} \times \frac{95.2 \text{ g MgCl}_2}{1 \text{ mol MgCl}_2} = 21.99 \text{ g MgCl}_2$$

Thus, for the original 300.0 g sample, the mass of  $\text{MgCl}_2$  is **65.97 g** as shown:

$$300.0 \text{ g sample} \times \frac{21.99 \text{ g MgCl}_2}{100.0 \text{ g sample}} = 65.97 \text{ g MgCl}_2$$

mass of  $\text{BaCl}_2$ :

When the sample was treated with  $\text{AgNO}_3$ , the  $\text{Cl}^-$  in the sample reacted and formed 195.8 g of  $\text{AgCl}$  as described by the following net ionic equation.



The total  $\text{Cl}^-$  in the 200.0 g sample can be calculated from the amount of  $\text{AgCl}$  formed:

$$\text{total Cl}^- = 195.8 \text{ g AgCl} \times \frac{35.45 \text{ g Cl}^-}{143.35 \text{ g AgCl}} = 48.42 \text{ g Cl}^-$$

This 48.42 g  $\text{Cl}^-$  comes from two sources:  $\text{MgCl}_2$  and  $\text{BaCl}_2$ .

$$\text{total Cl}^- = \text{Cl}^- \text{ from MgCl}_2 + \text{Cl}^- \text{ from BaCl}_2 = 48.42 \text{ g}$$

$$\text{mass of Cl}^- \text{ from MgCl}_2: 200 \text{ g sample} \times \frac{21.99 \text{ g MgCl}_2}{100 \text{ g sample}} \times \frac{70.90 \text{ g Cl}^-}{95.2 \text{ g MgCl}_2} = 32.75 \text{ g Cl}^-$$

mass of  $\text{Cl}^-$  from  $\text{BaCl}_2$  can be obtained by difference:

$$\begin{aligned} \text{Cl}^- \text{ from BaCl}_2 &= \text{Cl}^- \text{ total} - \text{Cl}^- \text{ from MgCl}_2 \\ &= 48.42 \text{ g} - 32.75 \text{ g} = 15.67 \text{ g Cl}^- \text{ from BaCl}_2 \end{aligned}$$

$$\text{mass of BaCl}_2 \text{ in the 200.0 g sample: } 15.67 \text{ g Cl}^- \times \frac{208.20 \text{ g BaCl}_2}{70.90 \text{ g Cl}^-} = 46.02 \text{ g BaCl}_2$$

$$\text{mass of BaCl}_2 \text{ in the 300.0 g sample: } 300.0 \text{ g sample} \times \frac{46.02 \text{ g BaCl}_2}{200.0 \text{ g sample}} = \mathbf{69.03 \text{ g BaCl}_2}$$

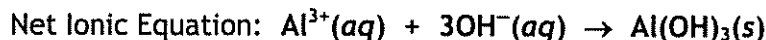
mass of  $\text{NaNO}_3$  in the 300.0 g sample is obtained by difference.

$$\text{mass of sample} = \text{mass of NaNO}_3 + \text{mass of MgCl}_2 + \text{mass of BaCl}_2$$

$$\text{mass of NaNO}_3 = 300.0 \text{ g} - [65.97 \text{ g MgCl}_2 + 69.03 \text{ g BaCl}_2] = \mathbf{165.0 \text{ g NaNO}_3}$$

The 300.0 g sample contained **165.0 g  $\text{NaNO}_3$** , **65.97 g  $\text{MgCl}_2$**  and **69.03 g  $\text{BaCl}_2$** .

78. This is a limiting reactant problem. A review of Example 4.3c will help you.



(a) To find the mass of  $\text{Al}(\text{OH})_3$  formed, determine first the limiting reactant by identifying the reactant that will produce fewer moles of precipitate.

mol precipitate if  $\text{Al}^{3+}$  is the limiting reactant:

Strategy (From Figure 4.6): M and V of parent compound,  $\text{Al}(\text{NO}_3)_3 \rightarrow$  mol of parent compound,  $\text{Al}(\text{NO}_3)_3 \rightarrow$  mol of reacting ion,  $\text{Al}^{3+} \rightarrow$  mol of precipitate,  $\text{Al}(\text{OH})_3$

$$\begin{aligned} 85.00 \text{ mL Al}(\text{NO}_3)_3 &\times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol Al}(\text{NO}_3)_3}{1 \text{ L Al}(\text{NO}_3)_3} \times \frac{1 \text{ mol Al}^{3+}}{1 \text{ mol Al}(\text{NO}_3)_3} \times \frac{1 \text{ mol Al}(\text{OH})_3}{1 \text{ mol Al}^{3+}} \\ &= 0.0213 \text{ mol Al}(\text{OH})_3 \end{aligned}$$

mol precipitate if  $\text{OH}^{-}$  is the limiting reactant:

Strategy (From Figure 4.6): M and V of parent compound,  $\text{Ba}(\text{OH})_2 \rightarrow$  mol of parent compound,  $\text{Ba}(\text{OH})_2 \rightarrow$  mol of reacting ion,  $\text{OH}^{-} \rightarrow$  mol of precipitate,  $\text{Al}(\text{OH})_3$

$$\begin{aligned} 85.0 \text{ mL Ba}(\text{OH})_2 &\times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol Ba}(\text{OH})_2}{1 \text{ L Ba}(\text{OH})_2} \times \frac{2 \text{ mol OH}^{-}}{1 \text{ mol Ba}(\text{OH})_2} \times \frac{1 \text{ mol Al}(\text{OH})_3}{3 \text{ mol OH}^{-}} \\ &= 0.0142 \text{ mol Al}(\text{OH})_3 \end{aligned}$$

Since 0.0142 mol is less than 0.0213 mol  $\text{Al}(\text{OH})_3$ ,  $\text{OH}^{-}$  is the limiting reactant. Hence, 0.0142 mol of precipitate forms. The mass of this precipitate is

$$0.0142 \text{ mol Al}(\text{OH})_3 \times \frac{78.00 \text{ g Al}(\text{OH})_3}{1 \text{ mol Al}(\text{OH})_3} = \underline{1.11 \text{ g Al}(\text{OH})_3}$$

(b) The ions present in solution before the reaction are  $\text{Ba}^{2+}$ ,  $\text{OH}^{-}$ ,  $\text{Al}^{3+}$  and  $\text{NO}_3^{-}$ . After the reaction, all  $\text{OH}^{-}$  are used up but  $\text{Ba}^{2+}$ ,  $\text{Al}^{3+}$  and  $\text{NO}_3^{-}$  will remain in the solution. Assuming the volume is additive, the final volume is 170.0 mL (see below).

$$\begin{aligned} \text{Total volume} &= \text{Ba}(\text{OH})_2 \text{ solution volume} + \text{Al}(\text{OH})_3 \text{ solution volume} \\ &= 85.0 \text{ mL} + 85.00 \text{ mL} \\ &= 170.0 \text{ mL or } 0.1700 \text{ L} \end{aligned}$$

$[\text{OH}^{-}] = \underline{0}$  because  $\text{OH}^{-}$  is the limiting reactant, it is completely used up.

$$[\text{Ba}^{2+}] = \underline{0.125 \text{ M}} \text{ (solution below)}$$

$\text{Ba}^{2+}$  is a spectator ion, thus its original number of moles is not changed.

$$[\text{Ba}^{2+}] = \frac{\text{moles Ba}^{2+}}{\text{total volume (L)}} \\ = \frac{85.0 \text{ mL Ba(OH)}_2 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.2500 \text{ mol Ba(OH)}_2}{1 \text{ L Ba(OH)}_2} \times \frac{1 \text{ mol Ba}^{2+}}{1 \text{ mol Ba(OH)}_2}}{0.1700 \text{ L}} = 0.125 \text{ M}$$

$$[\text{NO}_3^-] = \underline{0.375 \text{ M}} \text{ (solution below)}$$

$\text{NO}_3^-$  is also a spectator ion, thus its original number of moles is not changed.

$$[\text{NO}_3^-] = \frac{\text{moles NO}_3^-}{\text{total volume (L)}} \\ = \frac{85.00 \text{ mL Al(NO}_3)_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol Al(NO}_3)_3}{1 \text{ L Al(NO}_3)_3} \times \frac{3 \text{ mol NO}_3^-}{1 \text{ mol Al(NO}_3)_3}}{0.1700 \text{ L}} = 0.375 \text{ M}$$

$$[\text{Al}^{3+}] = \underline{0.0415 \text{ M}} \text{ (solution below)}$$

$\text{Al}^{3+}$  is the reactant present in excess. Some  $\text{Al}^{3+}$  are consumed in the reaction but some remain unreacted. The unreacted  $\text{Al}^{3+}$  is dissolved in the solution.

mol  $\text{Al}^{3+}$  originally available:

$$85.00 \text{ mL Al(NO}_3)_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol Al(NO}_3)_3}{1 \text{ L Al(NO}_3)_3} \times \frac{1 \text{ mol Al}^{3+}}{1 \text{ mol Al(NO}_3)_3} = 0.02125 \text{ mol} \\ = 0.02125 \text{ mol Al}^{3+} \text{ originally available}$$

mol  $\text{Al}^{3+}$  reacted:

$$85.0 \text{ mL Ba(OH)}_2 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol Ba(OH)}_2}{1 \text{ L Ba(OH)}_2} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ba(OH)}_2} \times \frac{1 \text{ mol Al}^{3+}}{3 \text{ mol OH}^-} \\ = 0.0142 \text{ mol Al}^{3+} \text{ reacted or consumed}$$

$$\text{excess Al}^{3+} \text{ (unreacted)} = 0.02125 \text{ mol Al}^{3+} \text{ available} - 0.0142 \text{ mol Al}^{3+} \text{ consumed} \\ = 0.00705 \text{ mol Al}^{3+}$$

$$[\text{Al}^{3+}] = \frac{\text{moles Al}^{3+}}{\text{total volume (L)}} = \frac{0.00705 \text{ mol Al}^{3+}}{0.1700 \text{ L}} = 0.0415 \text{ M}$$

Ions in solution:  $[\text{OH}^-] = 0$ ;  $[\text{Ba}^{2+}] = 0.125 \text{ M}$ ;  $[\text{NO}_3^-] = 0.375 \text{ M}$ ;  $[\text{Al}^{3+}] = 0.0415 \text{ M}$