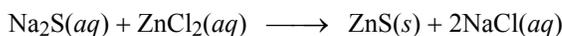


- 4.17** Refer to Table 4.2 of the text to solve this problem. AgCl is insoluble in water. It will precipitate from solution. NaNO₃ is soluble in water and will remain as Na⁺ and NO₃⁻ ions in solution. Diagram **(c)** best represents the mixture.
- 4.18** Refer to Table 4.2 of the text to solve this problem. Mg(OH)₂ is insoluble in water. It will precipitate from solution. KCl is soluble in water and will remain as K⁺ and Cl⁻ ions in solution. Diagram **(b)** best represents the mixture.
- 4.19** Refer to Table 4.2 of the text to solve this problem.
- (a) Ca₃(PO₄)₂ is *insoluble*.
 (b) Mn(OH)₂ is *insoluble*.
 (c) AgClO₃ is *soluble*.
 (d) K₂S is *soluble*.
- 4.20** **Strategy:** Although it is not necessary to memorize the solubilities of compounds, you should keep in mind the following useful rules: all ionic compounds containing alkali metal cations, the ammonium ion, and the nitrate, bicarbonate, and chlorate ions are soluble. For other compounds, refer to Table 4.2 of the text.

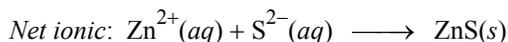
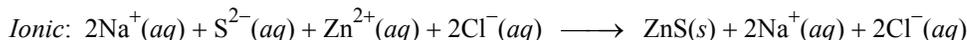
Solution:

- (a) CaCO₃ is **insoluble**. Most carbonate compounds are insoluble.
 (b) ZnSO₄ is **soluble**. Most sulfate compounds are soluble.
 (c) Hg(NO₃)₂ is **soluble**. All nitrate compounds are soluble.
 (d) HgSO₄ is **insoluble**. Most sulfate compounds are soluble, but those containing Ag⁺, Ca²⁺, Ba²⁺, Hg²⁺, and Pb²⁺ are insoluble.
 (e) NH₄ClO₄ is **soluble**. All ammonium compounds are soluble.
- 4.21** (a) Ionic: $2\text{Ag}^+(aq) + 2\text{NO}_3^-(aq) + 2\text{Na}^+(aq) + \text{SO}_4^{2-}(aq) \longrightarrow \text{Ag}_2\text{SO}_4(s) + 2\text{Na}^+(aq) + 2\text{NO}_3^-(aq)$
Net ionic: $2\text{Ag}^+(aq) + \text{SO}_4^{2-}(aq) \longrightarrow \text{Ag}_2\text{SO}_4(s)$
- (b) Ionic: $\text{Ba}^{2+}(aq) + 2\text{Cl}^-(aq) + \text{Zn}^{2+}(aq) + \text{SO}_4^{2-}(aq) \longrightarrow \text{BaSO}_4(s) + \text{Zn}^{2+}(aq) + 2\text{Cl}^-(aq)$
Net ionic: $\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \longrightarrow \text{BaSO}_4(s)$
- (c) Ionic: $2\text{NH}_4^+(aq) + \text{CO}_3^{2-}(aq) + \text{Ca}^{2+}(aq) + 2\text{Cl}^-(aq) \longrightarrow \text{CaCO}_3(s) + 2\text{NH}_4^+(aq) + 2\text{Cl}^-(aq)$
Net ionic: $\text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \longrightarrow \text{CaCO}_3(s)$
- 4.22** (a) **Strategy:** Recall that an *ionic equation* shows dissolved ionic compounds in terms of their free ions. A *net ionic equation* shows only the species that actually take part in the reaction. What happens when ionic compounds dissolve in water? What ions are formed from the dissociation of Na₂S and ZnCl₂? What happens when the cations encounter the anions in solution?

Solution: In solution, Na_2S dissociates into Na^+ and S^{2-} ions and ZnCl_2 dissociates into Zn^{2+} and Cl^- ions. According to Table 4.2 of the text, zinc ions (Zn^{2+}) and sulfide ions (S^{2-}) will form an insoluble compound, zinc sulfide (ZnS), while the other product, NaCl , is soluble and remains in solution. This is a precipitation reaction. The balanced molecular equation is:



The ionic and net ionic equations are:

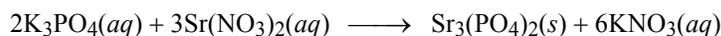


Check: Note that because we balanced the molecular equation first, the net ionic equation is balanced as to the number of atoms on each side, and the number of positive and negative charges on the left-hand side of the equation is the same.

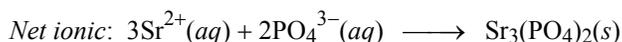
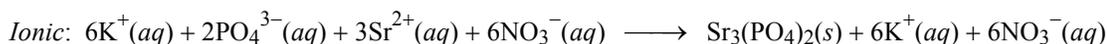
(b)

Strategy: What happens when ionic compounds dissolve in water? What ions are formed from the dissociation of K_3PO_4 and $\text{Sr}(\text{NO}_3)_2$? What happens when the cations encounter the anions in solution?

Solution: In solution, K_3PO_4 dissociates into K^+ and PO_4^{3-} ions and $\text{Sr}(\text{NO}_3)_2$ dissociates into Sr^{2+} and NO_3^- ions. According to Table 4.2 of the text, strontium ions (Sr^{2+}) and phosphate ions (PO_4^{3-}) will form an insoluble compound, strontium phosphate [$\text{Sr}_3(\text{PO}_4)_2$], while the other product, KNO_3 , is soluble and remains in solution. This is a precipitation reaction. The balanced molecular equation is:



The ionic and net ionic equations are:



Check: Note that because we balanced the molecular equation first, the net ionic equation is balanced as to the number of atoms on each side, and the number of positive and negative charges on the left-hand side of the equation is the same.

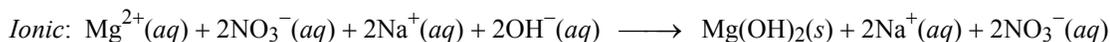
(c)

Strategy: What happens when ionic compounds dissolve in water? What ions are formed from the dissociation of $\text{Mg}(\text{NO}_3)_2$ and NaOH ? What happens when the cations encounter the anions in solution?

Solution: In solution, $\text{Mg}(\text{NO}_3)_2$ dissociates into Mg^{2+} and NO_3^- ions and NaOH dissociates into Na^+ and OH^- ions. According to Table 4.2 of the text, magnesium ions (Mg^{2+}) and hydroxide ions (OH^-) will form an insoluble compound, magnesium hydroxide [$\text{Mg}(\text{OH})_2$], while the other product, NaNO_3 , is soluble and remains in solution. This is a precipitation reaction. The balanced molecular equation is:

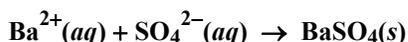


The ionic and net ionic equations are:



Check: Note that because we balanced the molecular equation first, the net ionic equation is balanced as to the number of atoms on each side, and the number of positive and negative charges on the left-hand side of the equation is the same.

- 4.23 (a) Both reactants are soluble ionic compounds. The other possible ion combinations, Na_2SO_4 and $\text{Cu}(\text{NO}_3)_2$, are also soluble.
- (b) Both reactants are soluble. Of the other two possible ion combinations, KCl is soluble, but BaSO_4 is insoluble and will precipitate.



- 4.24 (a) Add chloride ions. KCl is soluble, but AgCl is not.
- (b) Add hydroxide ions. $\text{Ba}(\text{OH})_2$ is soluble, but $\text{Pb}(\text{OH})_2$ is insoluble.
- (c) Add carbonate ions. $(\text{NH}_4)_2\text{CO}_3$ is soluble, but CaCO_3 is insoluble.
- (d) Add sulfate ions. CuSO_4 is soluble, but BaSO_4 is insoluble.

- 4.31 (a) HI dissolves in water to produce H^+ and I^- , so HI is a **Brønsted acid**.
- (b) CH_3COO^- can accept a proton to become acetic acid CH_3COOH , so it is a **Brønsted base**.
- (c) H_2PO_4^- can either accept a proton, H^+ , to become H_3PO_4 and thus behaves as a **Brønsted base**, or can donate a proton in water to yield H^+ and HPO_4^{2-} , thus behaving as a **Brønsted acid**.
- (d) HSO_4^- can either accept a proton, H^+ , to become H_2SO_4 and thus behaves as a **Brønsted base**, or can donate a proton in water to yield H^+ and SO_4^{2-} , thus behaving as a **Brønsted acid**.

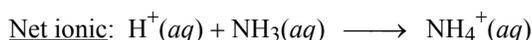
- 4.32 **Strategy:** What are the characteristics of a Brønsted acid? Does it contain at least an H atom? With the exception of ammonia, most Brønsted bases that you will encounter at this stage are anions.

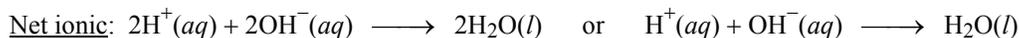
Solution:

- (a) PO_4^{3-} in water can accept a proton to become HPO_4^{2-} , and is thus a **Brønsted base**.
- (b) ClO_2^- in water can accept a proton to become HClO_2 , and is thus a **Brønsted base**.
- (c) NH_4^+ dissolved in water can donate a proton H^+ , thus behaving as a **Brønsted acid**.
- (d) HCO_3^- can either accept a proton to become H_2CO_3 , thus behaving as a **Brønsted base**. Or, HCO_3^- can donate a proton to yield H^+ and CO_3^{2-} , thus behaving as a **Brønsted acid**.

Comment: The HCO_3^- species is said to be *amphoteric* because it possesses both acidic and basic properties.

- 4.33 Recall that strong acids and strong bases are strong electrolytes. They are completely ionized in solution. An ionic equation will show strong acids and strong bases in terms of their free ions. A net ionic equation shows only the species that actually take part in the reaction.

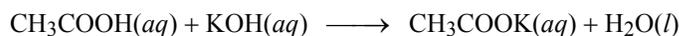




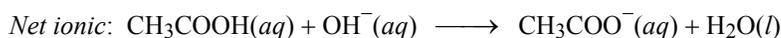
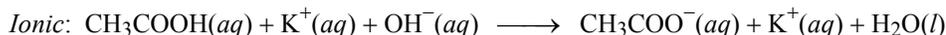
4.34 Strategy: Recall that strong acids and strong bases are strong electrolytes. They are completely ionized in solution. An *ionic equation* will show strong acids and strong bases in terms of their free ions. Weak acids and weak bases are weak electrolytes. They only ionize to a small extent in solution. Weak acids and weak bases are shown as molecules in ionic and net ionic equations. A *net ionic equation* shows only the species that actually take part in the reaction.

(a)

Solution: CH_3COOH is a weak acid. It will be shown as a molecule in the ionic equation. KOH is a strong base. It completely ionizes to K^{+} and OH^{-} ions. Since CH_3COOH is an acid, it donates an H^{+} to the base, OH^{-} , producing water. The other product is the salt, CH_3COOK , which is soluble and remains in solution. The balanced molecular equation is:



The ionic and net ionic equations are:

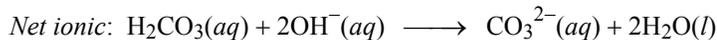
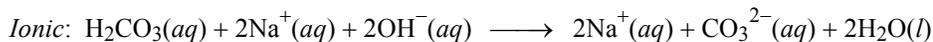


(b)

Solution: H_2CO_3 is a weak acid. It will be shown as a molecule in the ionic equation. NaOH is a strong base. It completely ionizes to Na^{+} and OH^{-} ions. Since H_2CO_3 is an acid, it donates an H^{+} to the base, OH^{-} , producing water. The other product is the salt, Na_2CO_3 , which is soluble and remains in solution. The balanced molecular equation is:



The ionic and net ionic equations are:

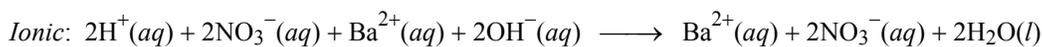


(c)

Solution: HNO_3 is a strong acid. It completely ionizes to H^{+} and NO_3^{-} ions. $\text{Ba}(\text{OH})_2$ is a strong base. It completely ionizes to Ba^{2+} and OH^{-} ions. Since HNO_3 is an acid, it donates an H^{+} to the base, OH^{-} , producing water. The other product is the salt, $\text{Ba}(\text{NO}_3)_2$, which is soluble and remains in solution. The balanced molecular equation is:



The ionic and net ionic equations are:



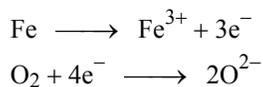
- 4.43 Even though the problem doesn't ask you to assign oxidation numbers, you need to be able to do so in order to determine what is being oxidized or reduced.

	<u>(i) Half Reactions</u>	<u>(ii) Oxidizing Agent</u>	<u>(iii) Reducing Agent</u>
(a)	$\text{Sr} \rightarrow \text{Sr}^{2+} + 2\text{e}^{-}$ $\text{O}_2 + 4\text{e}^{-} \rightarrow 2\text{O}^{2-}$	O_2	Sr
(b)	$\text{Li} \rightarrow \text{Li}^{+} + \text{e}^{-}$ $\text{H}_2 + 2\text{e}^{-} \rightarrow 2\text{H}^{-}$	H_2	Li
(c)	$\text{Cs} \rightarrow \text{Cs}^{+} + \text{e}^{-}$ $\text{Br}_2 + 2\text{e}^{-} \rightarrow 2\text{Br}^{-}$	Br_2	Cs
(d)	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^{-}$ $\text{N}_2 + 6\text{e}^{-} \rightarrow 2\text{N}^{3-}$	N_2	Mg

- 4.44 **Strategy:** In order to break a redox reaction down into an oxidation half-reaction and a reduction half-reaction, you should first assign oxidation numbers to all the atoms in the reaction. In this way, you can determine which element is oxidized (loses electrons) and which element is reduced (gains electrons).

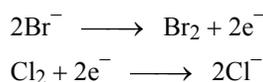
Solution: In each part, the reducing agent is the reactant in the first half-reaction and the oxidizing agent is the reactant in the second half-reaction. The coefficients in each half-reaction have been reduced to smallest whole numbers.

- (a) The product is an ionic compound whose ions are Fe^{3+} and O^{2-} .



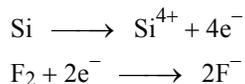
O_2 is the oxidizing agent; Fe is the reducing agent.

- (b) Na^{+} does not change in this reaction. It is a "spectator ion."



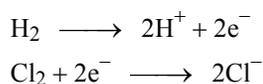
Cl_2 is the oxidizing agent; Br^{-} is the reducing agent.

- (c) Assume SiF_4 is made up of Si^{4+} and F^{-} .



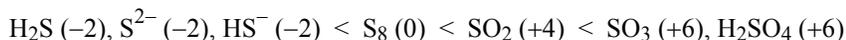
F_2 is the oxidizing agent; Si is the reducing agent.

- (d) Assume HCl is made up of H^{+} and Cl^{-} .



Cl_2 is the oxidizing agent; H_2 is the reducing agent.

- 4.45** The oxidation number for hydrogen is +1 (rule 4), and for oxygen is -2 (rule 3). The oxidation number for sulfur in S_8 is zero (rule 1). Remember that in a neutral molecule, the sum of the oxidation numbers of all the atoms must be zero, and in an ion the sum of oxidation numbers of all elements in the ion must equal the net charge of the ion (rule 6).



The number in parentheses denotes the oxidation number of sulfur.

- 4.46 Strategy:** In general, we follow the rules listed in Section 4.4 of the text for assigning oxidation numbers. Remember that all alkali metals have an oxidation number of +1 in ionic compounds, and in most cases hydrogen has an oxidation number of +1 and oxygen has an oxidation number of -2 in their compounds.

Solution: All the compounds listed are neutral compounds, so the oxidation numbers must sum to zero (Rule 6, Section 4.4 of the text).

Let the oxidation number of P = x .

- | | |
|---|--|
| (a) $x + 1 + (3)(-2) = 0, x = +5$ | (d) $x + (3)(+1) + (4)(-2) = 0, x = +5$ |
| (b) $x + (3)(+1) + (2)(-2) = 0, x = +1$ | (e) $2x + (4)(+1) + (7)(-2) = 0, 2x = 10, x = +5$ |
| (c) $x + (3)(+1) + (3)(-2) = 0, x = +3$ | (f) $3x + (5)(+1) + (10)(-2) = 0, 3x = 15, x = +5$ |

The molecules in part (a), (e), and (f) can be made by strongly heating the compound in part (d). Are these oxidation-reduction reactions?

Check: In each case, does the sum of the oxidation numbers of all the atoms equal the net charge on the species, in this case zero?

- 4.47** See Section 4.4 of the text.

- | | |
|--|--|
| (a) $\underline{Cl}F$: F -1 (rule 5), Cl +1 (rule 6) | (b) $\underline{I}F_7$: F -1 (rule 5), I +7 (rules 5 and 6) |
| (c) $\underline{C}H_4$: H +1 (rule 4), C -4 (rule 6) | (d) \underline{C}_2H_2 : H +1 (rule 4), C -1 (rule 6) |
| (e) \underline{C}_2H_4 : H +1 (rule 4), C -2 (rule 6), | (f) $K_2\underline{Cr}O_4$: K +1 (rule 2), O -2 (rule 3), Cr +6 (rule 6) |
| (g) $K_2\underline{Cr}_2O_7$: K +1 (rule 2), O -2 (rule 3), Cr +6 (rule 6) | |
| (h) $K\underline{Mn}O_4$: K +1 (rule 2), O -2 (rule 3), Mn +7 (rule 6) | |
| (i) $NaH\underline{C}O_3$: Na +1 (rule 2), H +1 (rule 4), O -2 (rule 3), C +4 (rule 6) | |
| (j) \underline{Li}_2 : Li 0 (rule 1) | (k) $Na\underline{I}O_3$: Na +1 (rule 2), O -2 (rule 3), I +5 (rule 6) |
| (l) $K\underline{O}_2$: K +1 (rule 2), O -1/2 (rule 6) | (m) $\underline{P}F_6^-$: F -1 (rule 5), P +5 (rule 6) |
| (n) $K\underline{Au}Cl_4$: K +1 (rule 2), Cl -1 (rule 5), Au +3 (rule 6) | |

- 4.48** All are free elements, so all have an oxidation number of **zero**.

- | | | | | |
|-----------------------------------|------------------------------------|------------------------------------|---------------------------------|------------------------------|
| (a) \underline{C}_s_2O , +1 | (b) $Ca\underline{I}_2$, -1 | (c) \underline{Al}_2O_3 , +3 | (d) $H_3\underline{As}O_3$, +3 | (e) $\underline{Ti}O_2$, +4 |
| (f) $\underline{Mn}O_4^{2-}$, +6 | (g) $\underline{Pt}Cl_4^{2-}$, +2 | (h) $\underline{Pt}Cl_6^{2-}$, +4 | (i) $\underline{Sn}F_2$, +2 | (j) $\underline{Cl}F_3$, +3 |
| (k) $\underline{Sb}F_6^-$, +5 | | | | |

- 4.50 (a) N: -3 (b) O: -1/2 (c) C: -1 (d) C: +4
 (e) C: +3 (f) O: -2 (g) B: +3 (h) W: +6

4.51 If nitric acid is a strong oxidizing agent and zinc is a strong reducing agent, then zinc metal will probably reduce nitric acid when the two react; that is, N will gain electrons and the oxidation number of N must decrease. Since the oxidation number of nitrogen in nitric acid is +5 (verify!), then the nitrogen-containing product must have a smaller oxidation number for nitrogen. The only compound in the list that doesn't have a nitrogen oxidation number less than +5 is N_2O_5 , (what is the oxidation number of N in N_2O_5 ?). This is never a product of the reduction of nitric acid.

4.52 **Strategy:** *Hydrogen displacement:* Any metal above hydrogen in the activity series will displace it from water or from an acid. Metals below hydrogen will *not* react with either water or an acid.

Solution: Only (b) Li and (d) Ca are above hydrogen in the activity series, so they are the only metals in this problem that will react with water.

4.53 In order to work this problem, you need to assign the oxidation numbers to all the elements in the compounds. In each case oxygen has an oxidation number of -2 (rule 3). These oxidation numbers should then be compared to the range of possible oxidation numbers that each element can have. Molecular oxygen is a powerful oxidizing agent. In SO_3 alone, the oxidation number of the element bound to oxygen (S) is at its maximum value (+6); the sulfur cannot be oxidized further. The other elements bound to oxygen in this problem have less than their maximum oxidation number and can undergo further oxidation.

4.54 (a) $Cu(s) + HCl(aq) \rightarrow$ no reaction, since $Cu(s)$ is less reactive than the hydrogen from acids.

(b) $I_2(s) + NaBr(aq) \rightarrow$ no reaction, since $I_2(s)$ is less reactive than $Br_2(l)$.

(c) $Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$, since $Mg(s)$ is more reactive than $Cu(s)$.

Net ionic equation: $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

(d) $Cl_2(g) + 2KBr(aq) \rightarrow Br_2(l) + 2KCl(aq)$, since $Cl_2(g)$ is more reactive than $Br_2(l)$

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

- 4.55 (a) Disproportionation reaction (b) Displacement reaction
 (c) Decomposition reaction (d) Combination reaction

- 4.56 (a) Combination reaction
 (b) Decomposition reaction
 (c) Displacement reaction
 (d) Disproportionation reaction

4.59 First, calculate the moles of KI needed to prepare the solution.

$$\text{mol KI} = \frac{2.80 \text{ mol KI}}{1000 \text{ mL soln}} \times (5.00 \times 10^2 \text{ mL soln}) = 1.40 \text{ mol KI}$$

Converting to grams of KI:

$$1.40 \text{ mol KI} \times \frac{166.0 \text{ g KI}}{1 \text{ mol KI}} = 232 \text{ g KI}$$

- 4.60 Strategy:** How many moles of NaNO_3 does 250 mL of a 0.707 M solution contain? How would you convert moles to grams?

Solution: From the molarity (0.707 M), we can calculate the moles of NaNO_3 needed to prepare 250 mL of solution.

$$\text{Moles NaNO}_3 = \frac{0.707 \text{ mol NaNO}_3}{1000 \text{ mL soln}} \times 250 \text{ mL soln} = 0.1768 \text{ mol}$$

Next, we use the molar mass of NaNO_3 as a conversion factor to convert from moles to grams.

$$\mathcal{M}(\text{NaNO}_3) = 85.00 \text{ g/mol.}$$

$$0.1768 \text{ mol NaNO}_3 \times \frac{85.00 \text{ g NaNO}_3}{1 \text{ mol NaNO}_3} = 15.0 \text{ g NaNO}_3$$

To make the solution, **dissolve 15.0 g of NaNO_3 in enough water to make 250 mL of solution.**

Check: As a ball-park estimate, the mass should be given by [molarity (mol/L) \times volume (L) = moles \times molar mass (g/mol) = grams]. Let's round the molarity to 1 M and the molar mass to 80 g, because we are simply making an estimate. This gives: [1 mol/L \times (1/4)L \times 80 g = 20 g]. This is close to our answer of 15.0 g.

- 4.61** mol = $M \times L$

$$60.0 \text{ mL} = 0.0600 \text{ L}$$

$$\text{mol MgCl}_2 = \frac{0.100 \text{ mol MgCl}_2}{1 \text{ L soln}} \times 0.0600 \text{ L soln} = 6.00 \times 10^{-3} \text{ mol MgCl}_2$$

- 4.62** Since the problem asks for grams of solute (KOH), you should be thinking that you can calculate moles of solute from the molarity and volume of solution. Then, you can convert moles of solute to grams of solute.

$$? \text{ moles KOH solute} = \frac{5.50 \text{ moles solute}}{1000 \text{ mL solution}} \times 35.0 \text{ mL solution} = 0.1925 \text{ mol KOH}$$

The molar mass of KOH is 56.11 g/mol. Use this conversion factor to calculate grams of KOH.

$$? \text{ grams KOH} = 0.1925 \text{ mol KOH} \times \frac{56.108 \text{ g KOH}}{1 \text{ mol KOH}} = 10.8 \text{ g KOH}$$

- 4.63** Molar mass of $\text{C}_2\text{H}_5\text{OH}$ = 46.068 g/mol; molar mass of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ = 342.3 g/mol; molar mass of NaCl = 58.44 g/mol.

$$\text{(a)} \quad ? \text{ mol C}_2\text{H}_5\text{OH} = 29.0 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.068 \text{ g C}_2\text{H}_5\text{OH}} = 0.6295 \text{ mol C}_2\text{H}_5\text{OH}$$

$$\text{Molarity} = \frac{\text{mol solute}}{\text{L of soln}} = \frac{0.6295 \text{ mol C}_2\text{H}_5\text{OH}}{0.545 \text{ L soln}} = 1.16 \text{ M}$$

$$\text{(b)} \quad ? \text{ mol C}_{12}\text{H}_{22}\text{O}_{11} = 15.4 \text{ g C}_{12}\text{H}_{22}\text{O}_{11} \times \frac{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}}{342.3 \text{ g C}_{12}\text{H}_{22}\text{O}_{11}} = 0.04499 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}$$

$$\text{Molarity} = \frac{\text{mol solute}}{\text{L of soln}} = \frac{0.04499 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}}{74.0 \times 10^{-3} \text{ L soln}} = \mathbf{0.608 M}$$

$$(c) \quad ? \text{ mol NaCl} = 9.00 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.154 \text{ mol NaCl}$$

$$\text{Molarity} = \frac{\text{mol solute}}{\text{L of soln}} = \frac{0.154 \text{ mol NaCl}}{86.4 \times 10^{-3} \text{ L soln}} = \mathbf{1.78 M}$$

$$4.64 \quad (a) \quad ? \text{ mol CH}_3\text{OH} = 6.57 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.042 \text{ g CH}_3\text{OH}} = 0.205 \text{ mol CH}_3\text{OH}$$

$$M = \frac{0.205 \text{ mol CH}_3\text{OH}}{0.150 \text{ L}} = \mathbf{1.37 M}$$

$$(b) \quad ? \text{ mol CaCl}_2 = 10.4 \text{ g CaCl}_2 \times \frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2} = 0.09371 \text{ mol CaCl}_2$$

$$M = \frac{0.09371 \text{ mol CaCl}_2}{0.220 \text{ L}} = \mathbf{0.426 M}$$

$$(c) \quad ? \text{ mol C}_{10}\text{H}_8 = 7.82 \text{ g C}_{10}\text{H}_8 \times \frac{1 \text{ mol C}_{10}\text{H}_8}{128.16 \text{ g C}_{10}\text{H}_8} = 0.06102 \text{ mol C}_{10}\text{H}_8$$

$$M = \frac{0.06102 \text{ mol C}_{10}\text{H}_8}{0.0852 \text{ L}} = \mathbf{0.716 M}$$

4.65 First, calculate the moles of each solute. Then, you can calculate the volume (in L) from the molarity and the number of moles of solute.

$$(a) \quad ? \text{ mol NaCl} = 2.14 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.03662 \text{ mol NaCl}$$

$$\text{L soln} = \frac{\text{mol solute}}{\text{Molarity}} = \frac{0.03662 \text{ mol NaCl}}{0.270 \text{ mol/L}} = 0.136 \text{ L} = \mathbf{136 \text{ mL soln}}$$

$$(b) \quad ? \text{ mol C}_2\text{H}_5\text{OH} = 4.30 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.068 \text{ g C}_2\text{H}_5\text{OH}} = 0.09334 \text{ mol C}_2\text{H}_5\text{OH}$$

$$\text{L soln} = \frac{\text{mol solute}}{\text{Molarity}} = \frac{0.09334 \text{ mol C}_2\text{H}_5\text{OH}}{1.50 \text{ mol/L}} = 0.0622 \text{ L} = \mathbf{62.2 \text{ mL soln}}$$

$$(c) \quad ? \text{ mol CH}_3\text{COOH} = 0.85 \text{ g CH}_3\text{COOH} \times \frac{1 \text{ mol CH}_3\text{COOH}}{60.052 \text{ g CH}_3\text{COOH}} = 0.0142 \text{ mol CH}_3\text{COOH}$$

$$\text{L soln} = \frac{\text{mol solute}}{\text{Molarity}} = \frac{0.0142 \text{ mol CH}_3\text{COOH}}{0.30 \text{ mol/L}} = 0.047 \text{ L} = \mathbf{47 \text{ mL soln}}$$

- 4.66** A 250 mL sample of 0.100 M solution contains 0.0250 mol of solute (mol = $M \times L$). The computation in each case is the same:

$$(a) \quad 0.0250 \text{ mol CsI} \times \frac{259.8 \text{ g CsI}}{1 \text{ mol CsI}} = \mathbf{6.50 \text{ g CsI}}$$

$$(b) \quad 0.0250 \text{ mol H}_2\text{SO}_4 \times \frac{98.086 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = \mathbf{2.45 \text{ g H}_2\text{SO}_4}$$

$$(c) \quad 0.0250 \text{ mol Na}_2\text{CO}_3 \times \frac{105.99 \text{ g Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3} = \mathbf{2.65 \text{ g Na}_2\text{CO}_3}$$

$$(d) \quad 0.0250 \text{ mol K}_2\text{Cr}_2\text{O}_7 \times \frac{294.2 \text{ g K}_2\text{Cr}_2\text{O}_7}{1 \text{ mol K}_2\text{Cr}_2\text{O}_7} = \mathbf{7.36 \text{ g K}_2\text{Cr}_2\text{O}_7}$$

$$(e) \quad 0.0250 \text{ mol KMnO}_4 \times \frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} = \mathbf{3.95 \text{ g KMnO}_4}$$

4.69 $M_{\text{initial}}V_{\text{initial}} = M_{\text{final}}V_{\text{final}}$

You can solve the equation algebraically for V_{initial} . Then substitute in the given quantities to solve for the volume of 2.00 M HCl needed to prepare 1.00 L of a 0.646 M HCl solution.

$$V_{\text{initial}} = \frac{M_{\text{final}} \times V_{\text{final}}}{M_{\text{initial}}} = \frac{0.646 \text{ M} \times 1.00 \text{ L}}{2.00 \text{ M}} = \mathbf{0.323 \text{ L} = 323 \text{ mL}}$$

To prepare the 0.646 M solution, you would dilute 323 mL of the 2.00 M HCl solution to a final volume of 1.00 L.

- 4.70 Strategy:** Because the volume of the final solution is greater than the original solution, this is a dilution process. Keep in mind that in a dilution, the concentration of the solution decreases, but the number of moles of the solute remains the same.

Solution: We prepare for the calculation by tabulating our data.

$$\begin{array}{ll} M_i = 0.866 \text{ M} & M_f = ? \\ V_i = 25.0 \text{ mL} & V_f = 500 \text{ mL} \end{array}$$

We substitute the data into Equation (4.3) of the text.

$$\begin{aligned} M_i V_i &= M_f V_f \\ (0.866 \text{ M})(25.0 \text{ mL}) &= M_f(500 \text{ mL}) \\ M_f &= \frac{(0.866 \text{ M})(25.0 \text{ mL})}{500 \text{ mL}} = \mathbf{0.0433 \text{ M}} \end{aligned}$$

4.71 $M_{\text{initial}}V_{\text{initial}} = M_{\text{final}}V_{\text{final}}$

You can solve the equation algebraically for V_{initial} . Then substitute in the given quantities to solve for the volume of 4.00 M HNO₃ needed to prepare 60.0 mL of a 0.200 M HNO₃ solution.

$$V_{\text{initial}} = \frac{M_{\text{final}} \times V_{\text{final}}}{M_{\text{initial}}} = \frac{0.200 \text{ M} \times 60.00 \text{ mL}}{4.00 \text{ M}} = \mathbf{3.00 \text{ mL}}$$

To prepare the 0.200 M solution, you would dilute 3.00 mL of the 4.00 M HNO₃ solution to a final volume of 60.0 mL.

- 4.72** You need to calculate the final volume of the dilute solution. Then, you can subtract 505 mL from this volume to calculate the amount of water that should be added.

$$V_{\text{final}} = \frac{M_{\text{initial}} V_{\text{initial}}}{M_{\text{final}}} = \frac{(0.125 \text{ M})(505 \text{ mL})}{(0.100 \text{ M})} = 631 \text{ mL}$$

$$(631 - 505) \text{ mL} = \mathbf{126 \text{ mL of water}}$$

- 4.73** Moles of KMnO₄ in the first solution:

$$\frac{1.66 \text{ mol}}{1000 \text{ mL soln}} \times 35.2 \text{ mL} = 0.05843 \text{ mol KMnO}_4$$

Moles of KMnO₄ in the second solution:

$$\frac{0.892 \text{ mol}}{1000 \text{ mL soln}} \times 16.7 \text{ mL} = 0.01490 \text{ mol KMnO}_4$$

The total volume is 35.2 mL + 16.7 mL = 51.9 mL. The concentration of the final solution is:

$$M = \frac{(0.05843 + 0.01490) \text{ mol}}{51.9 \times 10^{-3} \text{ L}} = \mathbf{1.41 \text{ M}}$$

- 4.74** Moles of calcium nitrate in the first solution:

$$\frac{0.568 \text{ mol}}{1000 \text{ mL soln}} \times 46.2 \text{ mL soln} = 0.02624 \text{ mol Ca(NO}_3)_2$$

Moles of calcium nitrate in the second solution:

$$\frac{1.396 \text{ mol}}{1000 \text{ mL soln}} \times 80.5 \text{ mL soln} = 0.1124 \text{ mol Ca(NO}_3)_2$$

The volume of the combined solutions = 46.2 mL + 80.5 mL = 126.7 mL. The concentration of the final solution is:

$$M = \frac{(0.02624 + 0.1124) \text{ mol}}{0.1267 \text{ L}} = \mathbf{1.09 \text{ M}}$$

- 4.77** The balanced equation is: $\text{CaCl}_2(aq) + 2\text{AgNO}_3(aq) \longrightarrow \text{Ca(NO}_3)_2(aq) + 2\text{AgCl}(s)$

We need to determine the limiting reagent. Ag⁺ and Cl⁻ combine in a 1:1 mole ratio to produce AgCl. Let's calculate the amount of Ag⁺ and Cl⁻ in solution.

$$\text{mol Ag}^+ = \frac{0.100 \text{ mol Ag}^+}{1000 \text{ mL soln}} \times 15.0 \text{ mL soln} = 1.50 \times 10^{-3} \text{ mol Ag}^+$$

$$\text{mol Cl}^- = \frac{0.150 \text{ mol CaCl}_2}{1000 \text{ mL soln}} \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2} \times 30.0 \text{ mL soln} = 9.00 \times 10^{-3} \text{ mol Cl}^-$$

Since Ag^+ and Cl^- combine in a 1:1 mole ratio, AgNO_3 is the limiting reagent. Only 1.50×10^{-3} mole of AgCl can form. Converting to grams of AgCl :

$$1.50 \times 10^{-3} \text{ mol AgCl} \times \frac{143.35 \text{ g AgCl}}{1 \text{ mol AgCl}} = \mathbf{0.215 \text{ g AgCl}}$$

4.78 Strategy: We want to calculate the mass % of Ba in the original compound. Let's start with the definition of mass %.

$$\text{mass \% Ba} = \frac{\text{mass Ba}}{\text{mass of sample}} \times 100\%$$

want to calculate
need to find
↓
↓
mass % Ba
mass Ba
=
↓
=
mass of sample
× 100%
↑
=
given

The mass of the sample is given in the problem (0.6760 g). Therefore we need to find the mass of Ba in the original sample. We assume the precipitation is quantitative, that is, that all of the barium in the sample has been precipitated as barium sulfate. From the mass of BaSO_4 produced, we can calculate the mass of Ba. There is 1 mole of Ba in 1 mole of BaSO_4 .

Solution: First, we calculate the mass of Ba in 0.4105 g of the BaSO_4 precipitate. The molar mass of BaSO_4 is 233.4 g/mol.

$$\begin{aligned} ? \text{ mass of Ba} &= 0.4105 \text{ g BaSO}_4 \times \frac{1 \text{ mol BaSO}_4}{233.37 \text{ g BaSO}_4} \times \frac{1 \text{ mol Ba}}{1 \text{ mol BaSO}_4} \times \frac{137.3 \text{ g Ba}}{1 \text{ mol Ba}} \\ &= 0.24151 \text{ g Ba} \end{aligned}$$

Next, we calculate the mass percent of Ba in the unknown compound.

$$\mathbf{\% \text{Ba by mass}} = \frac{0.24151 \text{ g}}{0.6760 \text{ g}} \times 100\% = \mathbf{35.73\%}$$

4.79 The net ionic equation is: $\text{Ag}^+(aq) + \text{Cl}^-(aq) \longrightarrow \text{AgCl}(s)$

One mole of Cl^- is required per mole of Ag^+ . First, find the number of moles of Ag^+ .

$$\text{mol Ag}^+ = \frac{0.0113 \text{ mol Ag}^+}{1000 \text{ mL soln}} \times (2.50 \times 10^2 \text{ mL soln}) = 2.825 \times 10^{-3} \text{ mol Ag}^+$$

Now, calculate the mass of NaCl using the mole ratio from the balanced equation.

$$(2.825 \times 10^{-3} \text{ mol Ag}^+) \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+} \times \frac{1 \text{ mol NaCl}}{1 \text{ mol Cl}^-} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = \mathbf{0.165 \text{ g NaCl}}$$

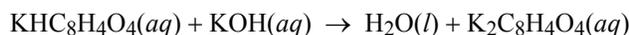
4.80 The net ionic equation is: $\text{Cu}^{2+}(aq) + \text{S}^{2-}(aq) \longrightarrow \text{CuS}(s)$

The answer sought is the molar concentration of Cu^{2+} , that is, moles of Cu^{2+} ions per liter of solution. The factor-label method is used to convert, in order:

g of CuS \rightarrow moles CuS \rightarrow moles Cu^{2+} \rightarrow moles Cu^{2+} per liter soln

$$[\text{Cu}^{2+}] = 0.0177 \text{ g CuS} \times \frac{1 \text{ mol CuS}}{95.62 \text{ g CuS}} \times \frac{1 \text{ mol Cu}^{2+}}{1 \text{ mol CuS}} \times \frac{1}{0.800 \text{ L}} = 2.31 \times 10^{-4} \text{ M}$$

4.85 The reaction between KHP ($\text{KHC}_8\text{H}_4\text{O}_4$) and KOH is:



We know the volume of the KOH solution, and we want to calculate the molarity of the KOH solution.

$$M \text{ of KOH} = \frac{\text{mol KOH}}{\text{L of KOH soln}}$$

want to calculate
need to find
↓
↓
 M of KOH
=
 $\frac{\text{mol KOH}}{\text{L of KOH soln}}$

↑

given

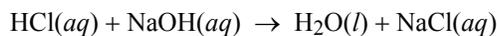
If we can determine the moles of KOH in the solution, we can then calculate the molarity of the solution. From the mass of KHP and its molar mass, we can calculate moles of KHP. Then, using the mole ratio from the balanced equation, we can calculate moles of KOH.

$$? \text{ mol KOH} = 0.4218 \text{ g KHP} \times \frac{1 \text{ mol KHP}}{204.22 \text{ g KHP}} \times \frac{1 \text{ mol KOH}}{1 \text{ mol KHP}} = 2.0654 \times 10^{-3} \text{ mol KOH}$$

From the moles and volume of KOH, we calculate the molarity of the KOH solution.

$$M \text{ of KOH} = \frac{\text{mol KOH}}{\text{L of KOH soln}} = \frac{2.0654 \times 10^{-3} \text{ mol KOH}}{18.68 \times 10^{-3} \text{ L soln}} = 0.1106 \text{ M}$$

4.86 The reaction between HCl and NaOH is:



We know the volume of the NaOH solution, and we want to calculate the molarity of the NaOH solution.

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

want to calculate
need to find
↓
↓
 M of NaOH
=
 $\frac{\text{mol NaOH}}{\text{L of NaOH soln}}$

↑

given

If we can determine the moles of NaOH in the solution, we can then calculate the molarity of the solution. From the volume and molarity of HCl, we can calculate moles of HCl. Then, using the mole ratio from the balanced equation, we can calculate moles of NaOH.

$$? \text{ mol NaOH} = 17.4 \text{ mL HCl} \times \frac{0.312 \text{ mol HCl}}{1000 \text{ mL soln}} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} = 5.429 \times 10^{-3} \text{ mol NaOH}$$

From the moles and volume of NaOH, we calculate the molarity of the NaOH solution.

$$M \text{ of NaOH} = \frac{\text{mol NaOH}}{\text{L of NaOH soln}} = \frac{5.429 \times 10^{-3} \text{ mol NaOH}}{25.0 \times 10^{-3} \text{ L soln}} = \mathbf{0.217 M}$$

- 4.87 (a)** In order to have the correct mole ratio to solve the problem, you must start with a balanced chemical equation.



From the molarity and volume of the HCl solution, you can calculate moles of HCl. Then, using the mole ratio from the balanced equation above, you can calculate moles of NaOH.

$$? \text{ mol NaOH} = 25.00 \text{ mL} \times \frac{2.430 \text{ mol HCl}}{1000 \text{ mL soln}} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} = 6.075 \times 10^{-2} \text{ mol NaOH}$$

Solving for the volume of NaOH:

$$\text{liters of solution} = \frac{\text{moles of solute}}{M}$$

$$\mathbf{\text{volume of NaOH}} = \frac{6.075 \times 10^{-2} \text{ mol NaOH}}{1.420 \text{ mol/L}} = 4.278 \times 10^{-2} \text{ L} = \mathbf{42.78 \text{ mL}}$$

- (b)** This problem is similar to part (a). The difference is that the mole ratio between base and acid is 2:1.



$$? \text{ mol NaOH} = 25.00 \text{ mL} \times \frac{4.500 \text{ mol H}_2\text{SO}_4}{1000 \text{ mL soln}} \times \frac{2 \text{ mol NaOH}}{1 \text{ mol H}_2\text{SO}_4} = 0.2250 \text{ mol NaOH}$$

$$\mathbf{\text{volume of NaOH}} = \frac{0.2250 \text{ mol NaOH}}{1.420 \text{ mol/L}} = 0.1585 \text{ L} = \mathbf{158.5 \text{ mL}}$$

- (c)** This problem is similar to parts (a) and (b). The difference is that the mole ratio between base and acid is 3:1.



$$? \text{ mol NaOH} = 25.00 \text{ mL} \times \frac{1.500 \text{ mol H}_3\text{PO}_4}{1000 \text{ mL soln}} \times \frac{3 \text{ mol NaOH}}{1 \text{ mol H}_3\text{PO}_4} = 0.1125 \text{ mol NaOH}$$

$$\mathbf{\text{volume of NaOH}} = \frac{0.1125 \text{ mol NaOH}}{1.420 \text{ mol/L}} = 0.07923 \text{ L} = \mathbf{79.23 \text{ mL}}$$

- 4.88 Strategy:** We know the molarity of the HCl solution, and we want to calculate the volume of the HCl solution.

$$M \text{ of HCl} = \frac{\text{mol HCl}}{\text{L of HCl soln}}$$

given need to find
 ↓ ↓
 mol HCl
 ↓
 L of HCl soln
 ↓
 want to calculate

If we can determine the moles of HCl, we can then use the definition of molarity to calculate the volume of HCl needed. From the volume and molarity of NaOH or Ba(OH)₂, we can calculate moles of NaOH or Ba(OH)₂. Then, using the mole ratio from the balanced equation, we can calculate moles of HCl.

Solution:

- (a) In order to have the correct mole ratio to solve the problem, you must start with a balanced chemical equation.



$$? \text{ mol HCl} = 10.0 \text{ mL} \times \frac{0.300 \text{ mol NaOH}}{1000 \text{ mL of solution}} \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} = 3.00 \times 10^{-3} \text{ mol HCl}$$

From the molarity and moles of HCl, we calculate volume of HCl required to neutralize the NaOH.

$$\text{liters of solution} = \frac{\text{moles of solute}}{M}$$

$$\text{volume of HCl} = \frac{3.00 \times 10^{-3} \text{ mol HCl}}{0.500 \text{ mol/L}} = 6.00 \times 10^{-3} \text{ L} = 6.00 \text{ mL}$$

- (b) This problem is similar to part (a). The difference is that the mole ratio between acid and base is 2:1.



$$? \text{ mol HCl} = 10.0 \text{ mL} \times \frac{0.200 \text{ mol Ba}(\text{OH})_2}{1000 \text{ mL of solution}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Ba}(\text{OH})_2} = 4.00 \times 10^{-3} \text{ mol HCl}$$

$$\text{volume of HCl} = \frac{4.00 \times 10^{-3} \text{ mol HCl}}{0.500 \text{ mol/L}} = 8.00 \times 10^{-3} \text{ L} = 8.00 \text{ mL}$$

- 4.91** The balanced equation is given in the problem. The mole ratio between Fe²⁺ and Cr₂O₇²⁻ is 6:1.

First, calculate the moles of Fe²⁺ that react with Cr₂O₇²⁻.

$$26.00 \text{ mL soln} \times \frac{0.0250 \text{ mol Cr}_2\text{O}_7^{2-}}{1000 \text{ mL soln}} \times \frac{6 \text{ mol Fe}^{2+}}{1 \text{ mol Cr}_2\text{O}_7^{2-}} = 3.90 \times 10^{-3} \text{ mol Fe}^{2+}$$

The molar concentration of Fe^{2+} is:

$$M = \frac{3.90 \times 10^{-3} \text{ mol Fe}^{2+}}{25.0 \times 10^{-3} \text{ L soln}} = \mathbf{0.156 \text{ M}}$$

- 4.92 Strategy:** We want to calculate the grams of SO_2 in the sample of air. From the molarity and volume of KMnO_4 , we can calculate moles of KMnO_4 . Then, using the mole ratio from the balanced equation, we can calculate moles of SO_2 . How do we convert from moles of SO_2 to grams of SO_2 ?

Solution: The balanced equation is given in the problem.



The moles of KMnO_4 required for the titration are:

$$\frac{0.00800 \text{ mol KMnO}_4}{1000 \text{ mL soln}} \times 7.37 \text{ mL} = 5.896 \times 10^{-5} \text{ mol KMnO}_4$$

We use the mole ratio from the balanced equation and the molar mass of SO_2 as conversion factors to convert to grams of SO_2 .

$$(5.896 \times 10^{-5} \text{ mol KMnO}_4) \times \frac{5 \text{ mol SO}_2}{2 \text{ mol KMnO}_4} \times \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} = \mathbf{9.44 \times 10^{-3} \text{ g SO}_2}$$

- 4.93** The balanced equation is given in problem 4.91. The mole ratio between Fe^{2+} and $\text{Cr}_2\text{O}_7^{2-}$ is 6:1.

First, calculate the moles of $\text{Cr}_2\text{O}_7^{2-}$ that reacted.

$$23.30 \text{ mL soln} \times \frac{0.0194 \text{ mol Cr}_2\text{O}_7^{2-}}{1000 \text{ mL soln}} = 4.52 \times 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-}$$

Use the mole ratio from the balanced equation to calculate the mass of iron that reacted.

$$(4.52 \times 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-}) \times \frac{6 \text{ mol Fe}^{2+}}{1 \text{ mol Cr}_2\text{O}_7^{2-}} \times \frac{55.85 \text{ g Fe}^{2+}}{1 \text{ mol Fe}^{2+}} = 0.1515 \text{ g Fe}^{2+}$$

The percent by mass of iron in the ore is:

$$\frac{0.1515 \text{ g}}{0.2792 \text{ g}} \times 100\% = \mathbf{54.3\%}$$

- 4.94** The balanced equation is given in the problem.



First, calculate the moles of potassium permanganate in 36.44 mL of solution.

$$\frac{0.01652 \text{ mol KMnO}_4}{1000 \text{ mL soln}} \times 36.44 \text{ mL} = 6.0199 \times 10^{-4} \text{ mol KMnO}_4$$

Next, calculate the moles of hydrogen peroxide using the mole ratio from the balanced equation.

$$(6.0199 \times 10^{-4} \text{ mol KMnO}_4) \times \frac{5 \text{ mol H}_2\text{O}_2}{2 \text{ mol KMnO}_4} = 1.505 \times 10^{-3} \text{ mol H}_2\text{O}_2$$

Finally, calculate the molarity of the H₂O₂ solution. The volume of the solution is 0.02500 L.

$$\text{Molarity of H}_2\text{O}_2 = \frac{1.505 \times 10^{-3} \text{ mol H}_2\text{O}_2}{0.02500 \text{ L}} = \mathbf{0.06020 \text{ M}}$$

4.95 First, calculate the moles of KMnO₄ in 24.0 mL of solution.

$$\frac{0.0100 \text{ mol KMnO}_4}{1000 \text{ mL soln}} \times 24.0 \text{ mL} = 2.40 \times 10^{-4} \text{ mol KMnO}_4$$

Next, calculate the mass of oxalic acid needed to react with 2.40×10^{-4} mol KMnO₄. Use the mole ratio from the balanced equation.

$$(2.40 \times 10^{-4} \text{ mol KMnO}_4) \times \frac{5 \text{ mol H}_2\text{C}_2\text{O}_4}{2 \text{ mol KMnO}_4} \times \frac{90.036 \text{ g H}_2\text{C}_2\text{O}_4}{1 \text{ mol H}_2\text{C}_2\text{O}_4} = 0.05402 \text{ g H}_2\text{C}_2\text{O}_4$$

The original sample had a mass of 1.00 g. The mass percent of H₂C₂O₄ in the sample is:

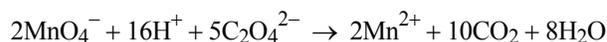
$$\text{mass \%} = \frac{0.05402 \text{ g}}{1.00 \text{ g}} \times 100\% = \mathbf{5.40\% \text{ H}_2\text{C}_2\text{O}_4}$$

4.96 From the reaction of oxalic acid with NaOH, the moles of oxalic acid in 15.0 mL of solution can be determined. Then, using this number of moles and other information given, the volume of the KMnO₄ solution needed to react with a second sample of oxalic acid can be calculated.

First, calculate the moles of oxalic acid in the solution. $\text{H}_2\text{C}_2\text{O}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{C}_2\text{O}_4(aq) + 2\text{H}_2\text{O}(l)$

$$0.0252 \text{ L} \times \frac{0.149 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol H}_2\text{C}_2\text{O}_4}{2 \text{ mol NaOH}} = 1.877 \times 10^{-3} \text{ mol H}_2\text{C}_2\text{O}_4$$

Because we are reacting a second sample of equal volume (15.0 mL), the moles of oxalic acid will also be 1.877×10^{-3} mole in this second sample. The balanced equation for the reaction between oxalic acid and KMnO₄ is:



Let's calculate the moles of KMnO₄ first, and then we will determine the volume of KMnO₄ needed to react with the 15.0 mL sample of oxalic acid.

$$(1.877 \times 10^{-3} \text{ mol H}_2\text{C}_2\text{O}_4) \times \frac{2 \text{ mol KMnO}_4}{5 \text{ mol H}_2\text{C}_2\text{O}_4} = 7.508 \times 10^{-4} \text{ mol KMnO}_4$$

Using Equation (4.2) of the text:

$$M = \frac{n}{V}$$

$$V_{\text{KMnO}_4} = \frac{n}{M} = \frac{7.508 \times 10^{-4} \text{ mol}}{0.122 \text{ mol/L}} = \mathbf{0.00615 \text{ L} = 6.15 \text{ mL}}$$

- 4.97** The balanced equation shows that 2 moles of electrons are lost for each mole of SO_3^{2-} that reacts. The electrons are gained by IO_3^- . We need to find the moles of electrons gained for each mole of IO_3^- that reacts. Then, we can calculate the final oxidation state of iodine.

The number of moles of electrons lost by SO_3^{2-} is:

$$32.5 \text{ mL} \times \frac{0.500 \text{ mol SO}_3^{2-}}{1000 \text{ mL soln}} \times \frac{2 \text{ mol e}^-}{1 \text{ mol SO}_3^{2-}} = 0.0325 \text{ mol e}^- \text{ lost}$$

The number of moles of iodate, IO_3^- , that react is:

$$1.390 \text{ g KIO}_3 \times \frac{1 \text{ mol KIO}_3}{214.0 \text{ g KIO}_3} \times \frac{1 \text{ mol IO}_3^-}{1 \text{ mol KIO}_3} = 6.4953 \times 10^{-3} \text{ mol IO}_3^-$$

6.4953×10^{-3} mole of IO_3^- gain 0.0325 mole of electrons. The number of moles of electrons gained per mole of IO_3^- is:

$$\frac{0.0325 \text{ mol e}^-}{6.4953 \times 10^{-3} \text{ mol IO}_3^-} = 5.00 \text{ mol e}^-/\text{mol IO}_3^-$$

The oxidation number of iodine in IO_3^- is +5. Since 5 moles of electrons are gained per mole of IO_3^- , the final oxidation state of iodine is $+5 - 5 = 0$. The iodine containing product of the reaction is most likely elemental iodine, I_2 .

- 4.98** The balanced equation is:



$$\text{mol MnO}_4^- = \frac{9.56 \times 10^{-4} \text{ mol MnO}_4^-}{1000 \text{ mL of soln}} \times 24.2 \text{ mL} = 2.314 \times 10^{-5} \text{ mol MnO}_4^-$$

Using the mole ratio from the balanced equation, we can calculate the mass of Ca^{2+} in the 10.0 mL sample of blood.

$$(2.314 \times 10^{-5} \text{ mol MnO}_4^-) \times \frac{5 \text{ mol C}_2\text{O}_4^{2-}}{2 \text{ mol MnO}_4^-} \times \frac{1 \text{ mol Ca}^{2+}}{1 \text{ mol C}_2\text{O}_4^{2-}} \times \frac{40.08 \text{ g Ca}^{2+}}{1 \text{ mol Ca}^{2+}} = 2.319 \times 10^{-3} \text{ g Ca}^{2+}$$

Converting to mg/mL:

$$\frac{2.319 \times 10^{-3} \text{ g Ca}^{2+}}{10.0 \text{ mL of blood}} \times \frac{1 \text{ mg}}{0.001 \text{ g}} = \mathbf{0.232 \text{ mg Ca}^{2+}/\text{mL of blood}}$$

4.99 In redox reactions the oxidation numbers of elements change. To test whether an equation represents a redox process, assign the oxidation numbers to each of the elements in the reactants and products. If oxidation numbers change, it is a redox reaction.

- (a) On the left the oxidation number of chlorine in Cl_2 is zero (rule 1). On the right it is -1 in Cl^- (rule 2) and $+1$ in OCl^- (rules 3 and 5). Since chlorine is both oxidized and reduced, this is a disproportionation redox reaction.
- (b) The oxidation numbers of calcium and carbon do not change. This is not a redox reaction; it is a precipitation reaction.
- (c) The oxidation numbers of nitrogen and hydrogen do not change. This is not a redox reaction; it is an acid-base reaction.
- (d) The oxidation numbers of carbon, chlorine, chromium, and oxygen do not change. This is not a redox reaction; it doesn't fit easily into any category, but could be considered as a type of combination reaction.
- (e) The oxidation number of calcium changes from 0 to $+2$, and the oxidation number of fluorine changes from 0 to -1 . This is a combination redox reaction.

The remaining parts (f) through (j) can be worked the same way.

- (f) Redox (g) Precipitation (h) Redox (i) Redox (j) Redox

4.100 First, the gases could be tested to see if they supported combustion. O_2 would support combustion, CO_2 would not. Second, if CO_2 is bubbled through a solution of calcium hydroxide $[\text{Ca}(\text{OH})_2]$, a white precipitate of CaCO_3 forms. No reaction occurs when O_2 is bubbled through a calcium hydroxide solution.

4.101 Choice (d), $0.20\text{ M Mg}(\text{NO}_3)_2$, should be the best conductor of electricity; the total ion concentration in this solution is 0.60 M . The total ion concentrations for solutions (a) and (c) are 0.40 M and 0.50 M , respectively. We can rule out choice (b), because acetic acid is a weak electrolyte.

4.102 Starting with a balanced chemical equation:



From the mass of Mg, you can calculate moles of Mg. Then, using the mole ratio from the balanced equation above, you can calculate moles of HCl reacted.

$$4.47\text{ g Mg} \times \frac{1\text{ mol Mg}}{24.31\text{ g Mg}} \times \frac{2\text{ mol HCl}}{1\text{ mol Mg}} = 0.3677\text{ mol HCl reacted}$$

Next we can calculate the number of moles of HCl in the original solution.

$$\frac{2.00\text{ mol HCl}}{1000\text{ mL soln}} \times (5.00 \times 10^2\text{ mL}) = 1.00\text{ mol HCl}$$

$$\text{Moles HCl remaining} = 1.00\text{ mol} - 0.3677\text{ mol} = 0.6323\text{ mol HCl}$$

$$\text{conc. of HCl after reaction} = \frac{\text{mol HCl}}{\text{L soln}} = \frac{0.6323\text{ mol HCl}}{0.500\text{ L}} = 1.26\text{ mol/L} = \mathbf{1.26\text{ M}}$$

4.103 The balanced equation for the displacement reaction is:



The moles of CuSO_4 that react with 7.89 g of zinc are:

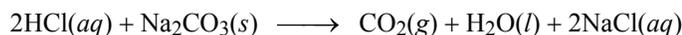
$$7.89 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{1 \text{ mol CuSO}_4}{1 \text{ mol Zn}} = 0.1207 \text{ mol CuSO}_4$$

The volume of the 0.156 M CuSO_4 solution needed to react with 7.89 g Zn is:

$$\text{L of soln} = \frac{\text{mole solute}}{M} = \frac{0.1207 \text{ mol CuSO}_4}{0.156 \text{ mol/L}} = \mathbf{0.774 \text{ L} = 774 \text{ mL}}$$

Would you expect Zn to displace Cu^{2+} from solution, as shown in the equation?

4.104 The balanced equation is:



The mole ratio from the balanced equation is 2 moles HCl : 1 mole Na_2CO_3 . The moles of HCl needed to react with 0.256 g of Na_2CO_3 are:

$$0.256 \text{ g Na}_2\text{CO}_3 \times \frac{1 \text{ mol Na}_2\text{CO}_3}{105.99 \text{ g Na}_2\text{CO}_3} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Na}_2\text{CO}_3} = 4.831 \times 10^{-3} \text{ mol HCl}$$

$$\text{Molarity HCl} = \frac{\text{moles HCl}}{\text{L soln}} = \frac{4.831 \times 10^{-3} \text{ mol HCl}}{0.0283 \text{ L soln}} = 0.171 \text{ mol/L} = \mathbf{0.171 \text{ M}}$$

4.105 The neutralization reaction is: $\text{HA}(aq) + \text{NaOH}(aq) \longrightarrow \text{NaA}(aq) + \text{H}_2\text{O}(l)$

The mole ratio between the acid and NaOH is 1:1. The moles of HA that react with NaOH are:

$$20.27 \text{ mL soln} \times \frac{0.1578 \text{ mol NaOH}}{1000 \text{ mL soln}} \times \frac{1 \text{ mol HA}}{1 \text{ mol NaOH}} = 3.1986 \times 10^{-3} \text{ mol HA}$$

3.664 g of the acid reacted with the base. The molar mass of the acid is:

$$\text{Molar mass} = \frac{3.664 \text{ g HA}}{3.1986 \times 10^{-3} \text{ mol HA}} = \mathbf{1146 \text{ g/mol}}$$

4.106 Starting with a balanced chemical equation:



From the molarity and volume of the NaOH solution, you can calculate moles of NaOH. Then, using the mole ratio from the balanced equation above, you can calculate moles of CH_3COOH .

$$5.75 \text{ mL solution} \times \frac{1.00 \text{ mol NaOH}}{1000 \text{ mL of solution}} \times \frac{1 \text{ mol CH}_3\text{COOH}}{1 \text{ mol NaOH}} = 5.75 \times 10^{-3} \text{ mol CH}_3\text{COOH}$$

$$\text{Molarity CH}_3\text{COOH} = \frac{5.75 \times 10^{-3} \text{ mol CH}_3\text{COOH}}{0.0500 \text{ L}} = \mathbf{0.115 \text{ M}}$$

- 4.107** Let's call the original solution, soln 1; the first dilution, soln 2; and the second dilution, soln 3. Start with the concentration of soln 3, 0.00383 M. From the concentration and volume of soln 3, we can find the concentration of soln 2. Then, from the concentration and volume of soln 2, we can find the concentration of soln 1, the original solution.

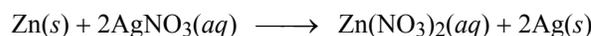
$$M_2V_2 = M_3V_3$$

$$M_2 = \frac{M_3V_3}{V_2} = \frac{(0.00383 \text{ M})(1.000 \times 10^3 \text{ mL})}{25.00 \text{ mL}} = 0.1532 \text{ M}$$

$$M_1V_1 = M_2V_2$$

$$M_1 = \frac{M_2V_2}{V_1} = \frac{(0.1532 \text{ M})(125.0 \text{ mL})}{15.00 \text{ mL}} = \mathbf{1.28 \text{ M}}$$

- 4.108** The balanced equation is:



Let x = mass of Ag produced. We can find the mass of Zn reacted in terms of the amount of Ag produced.

$$x \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}} \times \frac{1 \text{ mol Zn}}{2 \text{ mol Ag}} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 0.303x \text{ g Zn reacted}$$

The mass of Zn remaining will be:

$$2.50 \text{ g} - \text{amount of Zn reacted} = 2.50 \text{ g Zn} - 0.303x \text{ g Zn}$$

The final mass of the strip, 3.37 g, equals the mass of Ag produced + the mass of Zn remaining.

$$3.37 \text{ g} = x \text{ g Ag} + (2.50 \text{ g Zn} - 0.303x \text{ g Zn})$$

$$x = \mathbf{1.25 \text{ g} = \text{mass of Ag produced}}$$

$$\text{mass of Zn remaining} = 3.37 \text{ g} - 1.25 \text{ g} = \mathbf{2.12 \text{ g Zn}}$$

or

$$\text{mass of Zn remaining} = 2.50 \text{ g Zn} - 0.303x \text{ g Zn} = 2.50 \text{ g} - (0.303)(1.25 \text{ g}) = \mathbf{2.12 \text{ g Zn}}$$

- 4.109** The balanced equation is: $\text{Ba}(\text{OH})_2(aq) + \text{Na}_2\text{SO}_4(aq) \longrightarrow \text{BaSO}_4(s) + 2\text{NaOH}(aq)$

moles $\text{Ba}(\text{OH})_2$: $(2.27 \text{ L})(0.0820 \text{ mol/L}) = 0.1861 \text{ mol Ba}(\text{OH})_2$

moles Na_2SO_4 : $(3.06 \text{ L})(0.0664 \text{ mol/L}) = 0.2032 \text{ mol Na}_2\text{SO}_4$

Since the mole ratio between $\text{Ba}(\text{OH})_2$ and Na_2SO_4 is 1:1, $\text{Ba}(\text{OH})_2$ is the limiting reagent. The mass of BaSO_4 formed is:

$$0.1861 \text{ mol Ba}(\text{OH})_2 \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol Ba}(\text{OH})_2} \times \frac{233.37 \text{ g BaSO}_4}{\text{mol BaSO}_4} = \mathbf{43.4 \text{ g BaSO}_4}$$

4.110 The balanced equation is: $\text{HNO}_3(aq) + \text{NaOH}(aq) \longrightarrow \text{NaNO}_3(aq) + \text{H}_2\text{O}(l)$

$$\text{mol HNO}_3 = \frac{0.211 \text{ mol HNO}_3}{1000 \text{ mL soln}} \times 10.7 \text{ mL soln} = 2.258 \times 10^{-3} \text{ mol HNO}_3$$

$$\text{mol NaOH} = \frac{0.258 \text{ mol NaOH}}{1000 \text{ mL soln}} \times 16.3 \text{ mL soln} = 4.205 \times 10^{-3} \text{ mol NaOH}$$

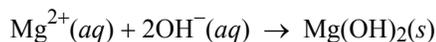
Since the mole ratio from the balanced equation is 1 mole NaOH : 1 mole HNO₃, then 2.258×10^{-3} mol HNO₃ will react with 2.258×10^{-3} mol NaOH.

$$\text{mol NaOH remaining} = (4.205 \times 10^{-3} \text{ mol}) - (2.258 \times 10^{-3} \text{ mol}) = 1.947 \times 10^{-3} \text{ mol NaOH}$$

$$10.7 \text{ mL} + 16.3 \text{ mL} = 27.0 \text{ mL} = 0.0270 \text{ L}$$

$$\text{molarity NaOH} = \frac{1.947 \times 10^{-3} \text{ mol NaOH}}{0.0270 \text{ L}} = \mathbf{0.0721 \text{ M}}$$

4.111 (a) Magnesium hydroxide is insoluble in water. It can be prepared by mixing a solution containing Mg²⁺ ions such as MgCl₂(aq) or Mg(NO₃)₂(aq) with a solution containing hydroxide ions such as NaOH(aq). Mg(OH)₂ will precipitate, which can then be collected by filtration. The net ionic reaction is:



(b) The balanced equation is: $2\text{HCl} + \text{Mg}(\text{OH})_2 \longrightarrow \text{MgCl}_2 + 2\text{H}_2\text{O}$

The moles of Mg(OH)₂ in 10 mL of milk of magnesia are:

$$10 \text{ mL soln} \times \frac{0.080 \text{ g Mg}(\text{OH})_2}{1 \text{ mL soln}} \times \frac{1 \text{ mol Mg}(\text{OH})_2}{58.326 \text{ g Mg}(\text{OH})_2} = 0.0137 \text{ mol Mg}(\text{OH})_2$$

$$\text{Moles of HCl reacted} = 0.0137 \text{ mol Mg}(\text{OH})_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg}(\text{OH})_2} = 0.0274 \text{ mol HCl}$$

$$\text{Volume of HCl} = \frac{\text{mol solute}}{M} = \frac{0.0274 \text{ mol HCl}}{0.035 \text{ mol/L}} = \mathbf{0.78 \text{ L}}$$

4.112 The balanced equations for the two reactions are:



First, let's find the number of moles of excess acid from the reaction with NaOH.

$$0.0334 \text{ L} \times \frac{0.500 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} = 8.35 \times 10^{-3} \text{ mol H}_2\text{SO}_4$$

The original number of moles of acid was:

$$0.100 \text{ L} \times \frac{0.500 \text{ mol H}_2\text{SO}_4}{1 \text{ L soln}} = 0.0500 \text{ mol H}_2\text{SO}_4$$

The amount of sulfuric acid that reacted with the metal, X, is

$$(0.0500 \text{ mol H}_2\text{SO}_4) - (8.35 \times 10^{-3} \text{ mol H}_2\text{SO}_4) = 0.04165 \text{ mol H}_2\text{SO}_4.$$

Since the mole ratio from the balanced equation is 1 mole X : 1 mole H₂SO₄, then the amount of X that reacted is 0.04165 mol X.

$$\text{molar mass X} = \frac{1.00 \text{ g X}}{0.04165 \text{ mol X}} = \mathbf{24.0 \text{ g/mol}}$$

The element is **magnesium**.

4.113 Add a known quantity of compound in a given quantity of water. Filter and recover the undissolved compound, then dry and weigh it. The difference in mass between the original quantity and the recovered quantity is the amount that dissolved in the water.

4.114 First, calculate the number of moles of glucose present.

$$\frac{0.513 \text{ mol glucose}}{1000 \text{ mL soln}} \times 60.0 \text{ mL} = 0.03078 \text{ mol glucose}$$

$$\frac{2.33 \text{ mol glucose}}{1000 \text{ mL soln}} \times 120.0 \text{ mL} = 0.2796 \text{ mol glucose}$$

Add the moles of glucose, then divide by the total volume of the combined solutions to calculate the molarity.

$$60.0 \text{ mL} + 120.0 \text{ mL} = 180.0 \text{ mL} = 0.180 \text{ L}$$

$$\text{Molarity of final solution} = \frac{(0.03078 + 0.2796) \text{ mol glucose}}{0.180 \text{ L}} = 1.72 \text{ mol/L} = \mathbf{1.72 \text{ M}}$$

4.115 First, you would accurately measure the electrical conductance of pure water. The conductance of a solution of the slightly soluble ionic compound X should be greater than that of pure water. The increased conductance would indicate that some of the compound X had dissolved.

4.116 Iron(II) compounds can be oxidized to iron(III) compounds. The sample could be tested with a small amount of a strongly colored oxidizing agent like a KMnO₄ solution, which is a deep purple color. A loss of color would imply the presence of an oxidizable substance like an iron(II) salt.

4.117 The three chemical tests might include:

- (1) Electrolysis to ascertain if hydrogen and oxygen were produced,
- (2) The reaction with an alkali metal to see if a base and hydrogen gas were produced, and
- (3) The dissolution of a metal oxide to see if a base was produced (or a nonmetal oxide to see if an acid was produced).

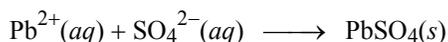
- 4.118** Since both of the original solutions were strong electrolytes, you would expect a mixture of the two solutions to also be a strong electrolyte. However, since the light dims, the mixture must contain fewer ions than the original solution. Indeed, H^+ from the sulfuric acid reacts with the OH^- from the barium hydroxide to form water. The barium cations react with the sulfate anions to form insoluble barium sulfate.



Thus, the reaction depletes the solution of ions and the conductivity decreases.

- 4.119** (a) Check with litmus paper, react with carbonate or bicarbonate to see if CO_2 gas is produced, react with a base and check with an indicator.
- (b) Titrate a known quantity of acid with a standard NaOH solution. Since it is a monoprotic acid, the moles of NaOH reacted equals the moles of the acid. Dividing the mass of acid by the number of moles gives the molar mass of the acid.
- (c) Visually compare the conductivity of the acid with a standard NaCl solution of the same molar concentration. A strong acid will have a similar conductivity to the NaCl solution. The conductivity of a weak acid will be considerably less than the NaCl solution.
- 4.120** You could test the conductivity of the solutions. Sugar is a nonelectrolyte and an aqueous sugar solution will not conduct electricity; whereas, NaCl is a strong electrolyte when dissolved in water. Silver nitrate could be added to the solutions to see if silver chloride precipitated. In this particular case, the solutions could also be tasted.

- 4.121** (a) $\text{Pb}(\text{NO}_3)_2(aq) + \text{Na}_2\text{SO}_4(aq) \longrightarrow \text{PbSO}_4(s) + 2\text{NaNO}_3(aq)$



- (b) First, calculate the moles of Pb^{2+} in the polluted water.

$$0.00450 \text{ g Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.05 \text{ g Na}_2\text{SO}_4} \times \frac{1 \text{ mol Pb}(\text{NO}_3)_2}{1 \text{ mol Na}_2\text{SO}_4} \times \frac{1 \text{ mol Pb}^{2+}}{1 \text{ mol Pb}(\text{NO}_3)_2} = 3.168 \times 10^{-5} \text{ mol Pb}^{2+}$$

The volume of the polluted water sample is 500 mL (0.500 L). The molar concentration of Pb^{2+} is:

$$[\text{Pb}^{2+}] = \frac{\text{mol Pb}^{2+}}{\text{L of soln}} = \frac{3.168 \times 10^{-5} \text{ mol Pb}^{2+}}{0.500 \text{ L soln}} = 6.34 \times 10^{-5} \text{ M}$$

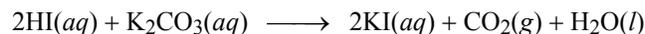
- 4.122** In a redox reaction, the oxidizing agent gains one or more electrons. In doing so, the oxidation number of the element gaining the electrons must become more negative. In the case of chlorine, the -1 oxidation number is already the most negative state possible. The chloride ion *cannot* accept any more electrons; therefore, hydrochloric acid is *not* an oxidizing agent.

- 4.123** (a) An acid and a base react to form water and a salt. Potassium iodide is a salt; therefore, the acid and base are chosen to produce this salt.



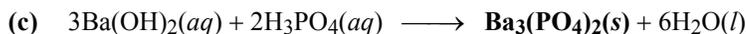
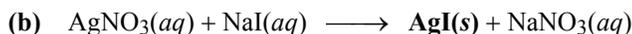
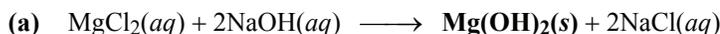
The water could be evaporated to isolate the KI .

- (b) Acids react with carbonates to form carbon dioxide gas. Again, chose the acid and carbonate salt so that KI is produced.



4.124 The reaction is too violent. This could cause the hydrogen gas produced to ignite, and an explosion could result.

4.125 All three products are water insoluble. Use this information in formulating your answer.



4.126 The solid sodium bicarbonate would be the better choice. The hydrogen carbonate ion, HCO_3^- , behaves as a Brønsted base to accept a proton from the acid.



The heat generated during the reaction of hydrogen carbonate with the acid causes the carbonic acid, H_2CO_3 , that was formed to decompose to water and carbon dioxide.

The reaction of the spilled sulfuric acid with sodium hydroxide would produce sodium sulfate, Na_2SO_4 , and water. There is a possibility that the Na_2SO_4 could precipitate. Also, the sulfate ion, SO_4^{2-} is a weak base; therefore, the “neutralized” solution would actually be *basic*.



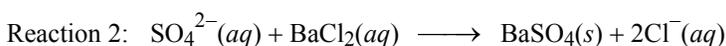
Also, NaOH is a caustic substance and therefore is not safe to use in this manner.

- 4.127**
- (a) A soluble sulfate salt such as sodium sulfate or sulfuric acid could be added. Barium sulfate would precipitate leaving sodium ions in solution.
 - (b) Potassium carbonate, phosphate, or sulfide could be added which would precipitate the magnesium cations, leaving potassium cations in solution.
 - (c) Add a soluble silver salt such as silver nitrate. AgBr would precipitate, leaving nitrate ions in solution.
 - (d) Add a solution containing a cation other than ammonium or a Group 1A cation to precipitate the phosphate ions; the nitrate ions will remain in solution.
 - (e) Add a solution containing a cation other than ammonium or a Group 1A cation to precipitate the carbonate ions; the nitrate ions will remain in solution.

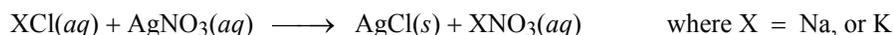
- 4.128**
- (a) Table salt, NaCl, is very soluble in water and is a strong electrolyte. Addition of AgNO_3 will precipitate AgCl.
 - (b) Table sugar or sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, is soluble in water and is a nonelectrolyte.
 - (c) Aqueous acetic acid, CH_3COOH , the primary ingredient of vinegar, is a weak electrolyte. It exhibits all of the properties of acids (Section 4.3).
 - (d) Baking soda, NaHCO_3 , is a water-soluble strong electrolyte. It reacts with acid to release CO_2 gas. Addition of $\text{Ca}(\text{OH})_2$ results in the precipitation of CaCO_3 .
 - (e) Washing soda, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, is a water-soluble strong electrolyte. It reacts with acids to release CO_2 gas. Addition of a soluble alkaline-earth salt will precipitate the alkaline-earth carbonate. Aqueous washing soda is also slightly basic (Section 4.3).
 - (f) Boric acid, H_3BO_3 , is weak electrolyte and a weak acid.
 - (g) Epsom salt, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$, is a water-soluble strong electrolyte. Addition of $\text{Ba}(\text{NO}_3)_2$ results in the precipitation of BaSO_4 . Addition of hydroxide precipitates $\text{Mg}(\text{OH})_2$.

- (h) Sodium hydroxide, NaOH, is a strong electrolyte and a strong base. Addition of $\text{Ca}(\text{NO}_3)_2$ results in the precipitation of $\text{Ca}(\text{OH})_2$.
- (i) Ammonia, NH_3 , is a sharp-odored gas that when dissolved in water is a weak electrolyte and a weak base. NH_3 in the gas phase reacts with HCl gas to produce solid NH_4Cl .
- (j) Milk of magnesia, $\text{Mg}(\text{OH})_2$, is an insoluble, strong base that reacts with acids. The resulting magnesium salt may be soluble or insoluble.
- (k) CaCO_3 is an insoluble salt that reacts with acid to release CO_2 gas. CaCO_3 is discussed in the Chemistry in Action essays entitled, “An Undesirable Precipitation Reaction” and “Metal from the Sea” in Chapter 4.

With the exception of NH_3 and vinegar, all the compounds in this problem are white solids.



4.130 The balanced equation for the reaction is:



From the amount of AgCl produced, we can calculate the moles of XCl reacted (X = Na, or K).

$$1.913 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.35 \text{ g AgCl}} \times \frac{1 \text{ mol XCl}}{1 \text{ mol AgCl}} = 0.013345 \text{ mol XCl}$$

Let x = number of moles NaCl. Then, the number of moles of KCl = $0.013345 \text{ mol} - x$. The sum of the NaCl and KCl masses must equal the mass of the mixture, 0.8870 g. We can write:

$$\text{mass NaCl} + \text{mass KCl} = 0.8870 \text{ g}$$

$$\left[x \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} \right] + \left[(0.013345 - x) \text{ mol KCl} \times \frac{74.55 \text{ g KCl}}{1 \text{ mol KCl}} \right] = 0.8870 \text{ g}$$

$$x = 6.6958 \times 10^{-3} = \text{moles NaCl}$$

$$\text{mol KCl} = 0.013345 - x = 0.013345 \text{ mol} - (6.6958 \times 10^{-3} \text{ mol}) = 6.6492 \times 10^{-3} \text{ mol KCl}$$

Converting moles to grams:

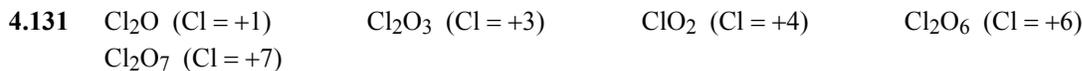
$$\text{mass NaCl} = (6.6958 \times 10^{-3} \text{ mol NaCl}) \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 0.3913 \text{ g NaCl}$$

$$\text{mass KCl} = (6.6492 \times 10^{-3} \text{ mol KCl}) \times \frac{74.55 \text{ g KCl}}{1 \text{ mol KCl}} = 0.4957 \text{ g KCl}$$

The percentages by mass for each compound are:

$$\% \text{ NaCl} = \frac{0.3913 \text{ g}}{0.8870 \text{ g}} \times 100\% = 44.11\% \text{ NaCl}$$

$$\% \text{ KCl} = \frac{0.4957 \text{ g}}{0.8870 \text{ g}} \times 100\% = 55.89\% \text{ KCl}$$



4.132 The number of moles of oxalic acid in 5.00×10^2 mL is:

$$\frac{0.100 \text{ mol H}_2\text{C}_2\text{O}_4}{1000 \text{ mL soln}} \times (5.00 \times 10^2 \text{ mL}) = 0.0500 \text{ mol H}_2\text{C}_2\text{O}_4$$

The balanced equation shows a mole ratio of 1 mol Fe_2O_3 : 6 mol $\text{H}_2\text{C}_2\text{O}_4$. The mass of rust that can be removed is:

$$0.0500 \text{ mol H}_2\text{C}_2\text{O}_4 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{6 \text{ mol H}_2\text{C}_2\text{O}_4} \times \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = \mathbf{1.33 \text{ g Fe}_2\text{O}_3}$$

4.133 Since aspirin is a monoprotic acid, it will react with NaOH in a 1:1 mole ratio.

First, calculate the moles of aspirin in the tablet.

$$12.25 \text{ mL soln} \times \frac{0.1466 \text{ mol NaOH}}{1000 \text{ mL soln}} \times \frac{1 \text{ mol aspirin}}{1 \text{ mol NaOH}} = 1.7959 \times 10^{-3} \text{ mol aspirin}$$

Next, convert from moles of aspirin to grains of aspirin.

$$1.7959 \times 10^{-3} \text{ mol aspirin} \times \frac{180.15 \text{ g aspirin}}{1 \text{ mol aspirin}} \times \frac{1 \text{ grain}}{0.0648 \text{ g}} = \mathbf{4.99 \text{ grains aspirin in one tablet}}$$

4.134 The precipitation reaction is: $\text{Ag}^+(aq) + \text{Br}^-(aq) \longrightarrow \text{AgBr}(s)$

In this problem, the relative amounts of NaBr and CaBr_2 are not known. However, the total amount of Br^- in the mixture can be determined from the amount of AgBr produced. Let's find the number of moles of Br^- .

$$1.6930 \text{ g AgBr} \times \frac{1 \text{ mol AgBr}}{187.8 \text{ g AgBr}} \times \frac{1 \text{ mol Br}^-}{1 \text{ mol AgBr}} = 9.0149 \times 10^{-3} \text{ mol Br}^-$$

The amount of Br^- comes from both NaBr and CaBr_2 . Let x = number of moles NaBr. Then, the number of moles of $\text{CaBr}_2 = \frac{9.0149 \times 10^{-3} \text{ mol} - x}{2}$. The moles of CaBr_2 are divided by 2, because 1 mol of CaBr_2 produces 2 moles of Br^- . The sum of the NaBr and CaBr_2 masses must equal the mass of the mixture, 0.9157 g. We can write:

$$\text{mass NaBr} + \text{mass CaBr}_2 = 0.9157 \text{ g}$$

$$\left[x \text{ mol NaBr} \times \frac{102.89 \text{ g NaBr}}{1 \text{ mol NaBr}} \right] + \left[\left(\frac{9.0149 \times 10^{-3} - x}{2} \right) \text{ mol CaBr}_2 \times \frac{199.88 \text{ g CaBr}_2}{1 \text{ mol CaBr}_2} \right] = 0.9157 \text{ g}$$

$$2.95x = 0.014751$$

$$x = 5.0003 \times 10^{-3} = \text{moles NaBr}$$

Converting moles to grams:

$$\text{mass NaBr} = (5.0003 \times 10^{-3} \text{ mol NaBr}) \times \frac{102.89 \text{ g NaBr}}{1 \text{ mol NaBr}} = 0.51448 \text{ g NaBr}$$

The percentage by mass of NaBr in the mixture is:

$$\% \text{ NaBr} = \frac{0.51448 \text{ g}}{0.9157 \text{ g}} \times 100\% = \mathbf{56.18\% \text{ NaBr}}$$

- 4.135** (a) $\text{CaF}_2(s) + \text{H}_2\text{SO}_4(aq) \longrightarrow 2\text{HF}(g) + \text{CaSO}_4(s)$
 $2\text{NaCl}(s) + \text{H}_2\text{SO}_4(aq) \longrightarrow 2\text{HCl}(aq) + \text{Na}_2\text{SO}_4(aq)$
- (b) HBr and HI *cannot* be prepared similarly, because Br^- and I^- would be oxidized to the element, Br_2 and I_2 , respectively.
- $2\text{NaBr}(s) + 2\text{H}_2\text{SO}_4(aq) \longrightarrow \text{Br}_2(l) + \text{SO}_2(g) + \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l)$
- (c) $\text{PBr}_3(l) + 3\text{H}_2\text{O}(l) \longrightarrow 3\text{HBr}(g) + \text{H}_3\text{PO}_3(aq)$

4.136 There are two moles of Cl^- per one mole of CaCl_2 .

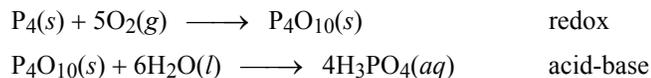
(a) $25.3 \text{ g CaCl}_2 \times \frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2} \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2} = 0.4559 \text{ mol Cl}^-$

$$\text{Molarity Cl}^- = \frac{0.4559 \text{ mol Cl}^-}{0.325 \text{ L soln}} = 1.40 \text{ mol/L} = \mathbf{1.40 \text{ M}}$$

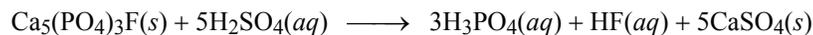
(b) We need to convert from mol/L to grams in 0.100 L.

$$\frac{1.40 \text{ mol Cl}^-}{1 \text{ L soln}} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}^-} \times 0.100 \text{ L soln} = \mathbf{4.96 \text{ g Cl}^-}$$

4.137 Electric furnace method:



Wet process:



This is a precipitation and an acid-base reaction.

- 4.138** (a) $\text{NH}_4^+(aq) + \text{OH}^-(aq) \longrightarrow \text{NH}_3(aq) + \text{H}_2\text{O}(l)$
- (b) From the amount of NaOH needed to neutralize the 0.2041 g sample, we can find the amount of the 0.2041 g sample that is NH_4NO_3 .
 First, calculate the moles of NaOH.

$$\frac{0.1023 \text{ mol NaOH}}{1000 \text{ mL of soln}} \times 24.42 \text{ mL soln} = 2.4982 \times 10^{-3} \text{ mol NaOH}$$

Using the mole ratio from the balanced equation, we can calculate the amount of NH_4NO_3 that reacted.

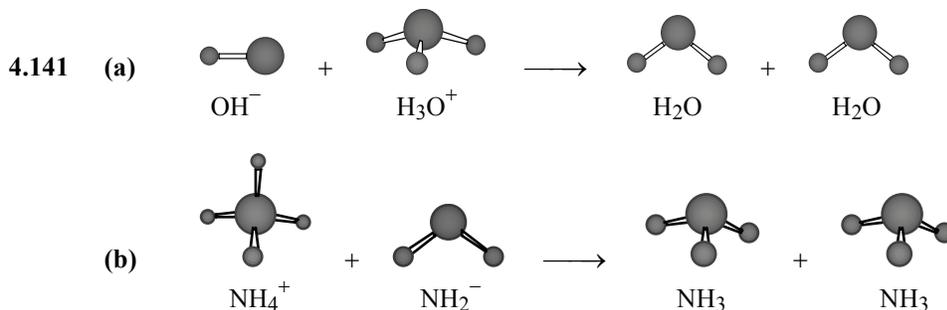
$$(2.4982 \times 10^{-3} \text{ mol NaOH}) \times \frac{1 \text{ mol NH}_4\text{NO}_3}{1 \text{ mol NaOH}} \times \frac{80.052 \text{ g NH}_4\text{NO}_3}{1 \text{ mol NH}_4\text{NO}_3} = 0.19999 \text{ g NH}_4\text{NO}_3$$

The purity of the NH_4NO_3 sample is:

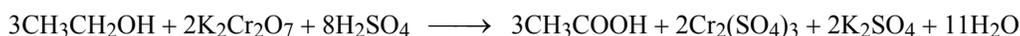
$$\% \text{ purity} = \frac{0.19999 \text{ g}}{0.2041 \text{ g}} \times 100\% = \mathbf{97.99\%}$$

4.139 In a redox reaction, electrons must be transferred between reacting species. In other words, oxidation numbers must change in a redox reaction. In both O_2 (molecular oxygen) and O_3 (ozone), the oxidation number of oxygen is zero. This is *not* a redox reaction.

4.140 Using the rules for assigning oxidation numbers given in Section 4.4, H is +1, F is -1, so the oxidation number of O must be **zero**.



4.142 The balanced equation is:



From the amount of $\text{K}_2\text{Cr}_2\text{O}_7$ required to react with the blood sample, we can calculate the mass of ethanol ($\text{CH}_3\text{CH}_2\text{OH}$) in the 10.0 g sample of blood.

First, calculate the moles of $\text{K}_2\text{Cr}_2\text{O}_7$ reacted.

$$\frac{0.07654 \text{ mol K}_2\text{Cr}_2\text{O}_7}{1000 \text{ mL soln}} \times 4.23 \text{ mL} = 3.238 \times 10^{-4} \text{ mol K}_2\text{Cr}_2\text{O}_7$$

Next, using the mole ratio from the balanced equation, we can calculate the mass of ethanol that reacted.

$$3.238 \times 10^{-4} \text{ mol K}_2\text{Cr}_2\text{O}_7 \times \frac{3 \text{ mol ethanol}}{2 \text{ mol K}_2\text{Cr}_2\text{O}_7} \times \frac{46.068 \text{ g ethanol}}{1 \text{ mol ethanol}} = 0.02238 \text{ g ethanol}$$

The percent ethanol by mass is:

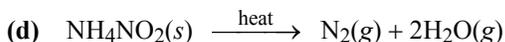
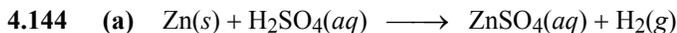
$$\% \text{ by mass ethanol} = \frac{0.02238 \text{ g}}{10.0 \text{ g}} \times 100\% = \mathbf{0.224\%}$$

This is well above the legal limit of 0.1 percent by mass ethanol in the blood. The individual should be prosecuted for drunk driving.

- 4.143** Notice that nitrogen is in its highest possible oxidation state (+5) in nitric acid. It is reduced as it decomposes to NO_2 .



The yellow color of “old” nitric acid is caused by the production of small amounts of NO_2 which is a brown gas. This process is accelerated by light.



- 4.145** Because the volume of the solution changes (increases or decreases) when the solid dissolves.

- 4.146** NH_4Cl exists as NH_4^+ and Cl^- . To form NH_3 and HCl , a proton (H^+) is transferred from NH_4^+ to Cl^- . Therefore, this is a Brønsted acid-base reaction.

- 4.147 (a)** The precipitate CaSO_4 formed over Ca preventing the Ca from reacting with the sulfuric acid.

(b) Aluminum is protected by a tenacious oxide layer with the composition Al_2O_3 .

(c) These metals react more readily with water.



(d) The metal should be placed below Fe and above H .

(e) Any metal above Al in the activity series will react with Al^{3+} . Metals from Mg to Li will work.

- 4.148 (a)**

First Solution:

$$0.8214 \text{ g KMnO}_4 \times \frac{1 \text{ mol KMnO}_4}{158.04 \text{ g KMnO}_4} = 5.1974 \times 10^{-3} \text{ mol KMnO}_4$$

$$M = \frac{\text{mol solute}}{\text{L of soln}} = \frac{5.1974 \times 10^{-3} \text{ mol KMnO}_4}{0.5000 \text{ L}} = 1.0395 \times 10^{-2} M$$

Second Solution:

$$\begin{aligned} M_1V_1 &= M_2V_2 \\ (1.0395 \times 10^{-2} M)(2.000 \text{ mL}) &= M_2(1000 \text{ mL}) \\ M_2 &= 2.079 \times 10^{-5} M \end{aligned}$$

Third Solution:

$$M_1V_1 = M_2V_2$$

$$(2.079 \times 10^{-5} M)(10.00 \text{ mL}) = M_2(250.0 \text{ mL})$$

$$M_2 = 8.316 \times 10^{-7} M$$

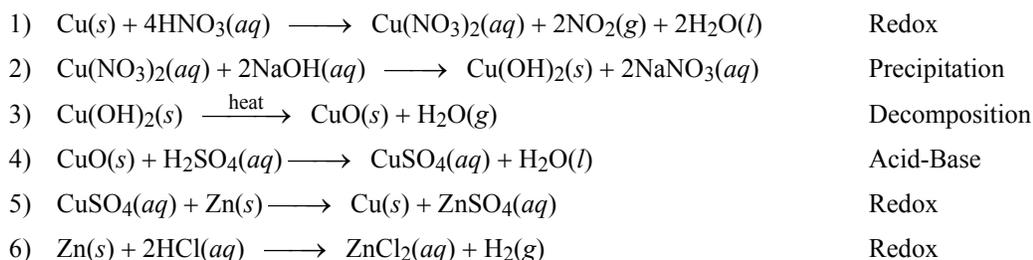
- (b) From the molarity and volume of the final solution, we can calculate the moles of KMnO_4 . Then, the mass can be calculated from the moles of KMnO_4 .

$$\frac{8.316 \times 10^{-7} \text{ mol KMnO}_4}{1000 \text{ mL of soln}} \times 250 \text{ mL} = 2.079 \times 10^{-7} \text{ mol KMnO}_4$$

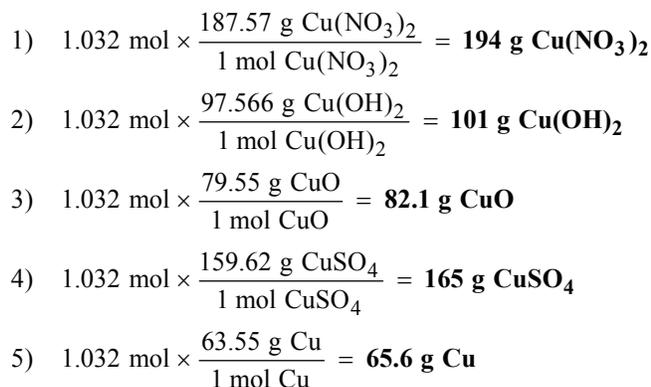
$$2.079 \times 10^{-7} \text{ mol KMnO}_4 \times \frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} = 3.286 \times 10^{-5} \text{ g KMnO}_4$$

This mass is too small to directly weigh accurately.

- 4.149 (a)** The balanced equations are:



- (b) We start with 65.6 g Cu, which is $65.6 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} = 1.032 \text{ mol Cu}$. The mole ratio between product and reactant in each reaction is 1:1. Therefore, the theoretical yield in each reaction is 1.032 moles.



- (c) All of the reaction steps are clean and almost quantitative; therefore, the recovery yield should be high.

- 4.150** The first titration oxidizes Fe^{2+} to Fe^{3+} . This titration gives the amount of Fe^{2+} in solution. Zn metal is added to reduce all Fe^{3+} back to Fe^{2+} . The second titration oxidizes all the Fe^{2+} back to Fe^{3+} . We can find the amount of Fe^{3+} in the original solution by difference.

Titration #1: The mole ratio between Fe^{2+} and MnO_4^- is 5:1.

$$23.0 \text{ mL soln} \times \frac{0.0200 \text{ mol MnO}_4^-}{1000 \text{ mL soln}} \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} = 2.30 \times 10^{-3} \text{ mol Fe}^{2+}$$

$$[\text{Fe}^{2+}] = \frac{\text{mol solute}}{\text{L of soln}} = \frac{2.30 \times 10^{-3} \text{ mol Fe}^{2+}}{25.0 \times 10^{-3} \text{ L soln}} = \mathbf{0.0920 \text{ M}}$$

Titration #2: The mole ratio between Fe^{2+} and MnO_4^- is 5:1.

$$40.0 \text{ mL soln} \times \frac{0.0200 \text{ mol MnO}_4^-}{1000 \text{ mL soln}} \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} = 4.00 \times 10^{-3} \text{ mol Fe}^{2+}$$

In this second titration, there are more moles of Fe^{2+} in solution. This is due to Fe^{3+} in the original solution being reduced by Zn to Fe^{2+} . The number of moles of Fe^{3+} in solution is:

$$(4.00 \times 10^{-3} \text{ mol}) - (2.30 \times 10^{-3} \text{ mol}) = 1.70 \times 10^{-3} \text{ mol Fe}^{3+}$$

$$[\text{Fe}^{3+}] = \frac{\text{mol solute}}{\text{L of soln}} = \frac{1.70 \times 10^{-3} \text{ mol Fe}^{3+}}{25.0 \times 10^{-3} \text{ L soln}} = \mathbf{0.0680 \text{ M}}$$

4.151 Place the following metals in the correct positions on the periodic table framework provided in the problem.

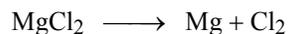
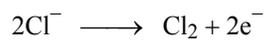
(a) Li, Na (b) Mg, Fe (c) Zn, Cd

Two metals that do not react with water or acid are Ag and Au.

4.152 (a) The precipitation reaction is: $\text{Mg}^{2+}(aq) + 2\text{OH}^-(aq) \longrightarrow \text{Mg}(\text{OH})_2(s)$

The acid-base reaction is: $\text{Mg}(\text{OH})_2(s) + 2\text{HCl}(aq) \longrightarrow \text{MgCl}_2(aq) + 2\text{H}_2\text{O}(l)$

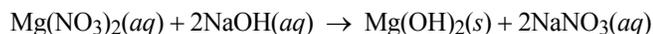
The redox reactions are: $\text{Mg}^{2+} + 2e^- \longrightarrow \text{Mg}$



(b) NaOH is much more expensive than CaO.

(c) Dolomite has the advantage of being an additional source of magnesium that can also be recovered.

4.153 The reaction between $\text{Mg}(\text{NO}_3)_2$ and NaOH is:



Magnesium hydroxide, $\text{Mg}(\text{OH})_2$, precipitates from solution. Na^+ and NO_3^- are spectator ions. This is most likely a limiting reagent problem as the amounts of both reactants are given. Let's first determine which reactant is the limiting reagent before we try to determine the concentration of ions remaining in the solution.

$$1.615 \text{ g Mg}(\text{NO}_3)_2 \times \frac{1 \text{ mol Mg}(\text{NO}_3)_2}{148.33 \text{ g Mg}(\text{NO}_3)_2} = 0.010888 \text{ mol Mg}(\text{NO}_3)_2$$

$$1.073 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{39.998 \text{ g NaOH}} = 0.026826 \text{ mol NaOH}$$

From the balanced equation, we need twice as many moles of NaOH compared to $\text{Mg}(\text{NO}_3)_2$. We have more than twice as much NaOH ($2 \times 0.010888 \text{ mol} = 0.021776 \text{ mol}$) and therefore $\text{Mg}(\text{NO}_3)_2$ is the limiting reagent. NaOH is in excess and ions of Na^+ , OH^- , and NO_3^- will remain in solution. Because Na^+ and NO_3^- are spectator ions, the number of moles after reaction will equal the initial number of moles. The excess moles of OH^- need to be calculated based on the amount that reacts with Mg^{2+} . The combined volume of the two solutions is: $22.02 \text{ mL} + 28.64 \text{ mL} = 50.66 \text{ mL} = 0.05066 \text{ L}$.

$$[\text{Na}^+] = 0.026826 \text{ mol NaOH} \times \frac{1 \text{ mol Na}^+}{1 \text{ mol NaOH}} \times \frac{1}{0.05066 \text{ L}} = \mathbf{0.5295 \text{ M}}$$

$$[\text{NO}_3^-] = 0.010888 \text{ mol Mg}(\text{NO}_3)_2 \times \frac{2 \text{ mol NO}_3^-}{1 \text{ mol Mg}(\text{NO}_3)_2} \times \frac{1}{0.05066 \text{ L}} = \mathbf{0.4298 \text{ M}}$$

The moles of OH^- reacted are:

$$0.010888 \text{ mol Mg}^{2+} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Mg}^{2+}} = 0.021776 \text{ mol OH}^- \text{ reacted}$$

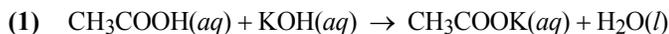
The moles of excess OH^- are:

$$0.026826 \text{ mol} - 0.021776 \text{ mol} = 0.005050 \text{ mol OH}^-$$

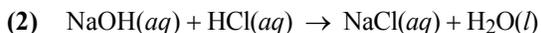
$$[\text{OH}^-] = \frac{0.005050 \text{ mol}}{0.05066 \text{ L}} = \mathbf{0.09968 \text{ M}}$$

The concentration of Mg^{2+} is approximately zero as almost all of it will precipitate as $\text{Mg}(\text{OH})_2$.

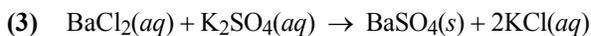
4.154 Let's set up a table showing each reaction, the volume of solution added, and the species responsible for any electrical conductance of the solution. Note that if a substance completely dissociates into +1 ions and -1 ions in solution, its conductance unit will be twice its molarity. Similarly, if a substance completely dissociates into +2 ions and -2 ions in solution, its conductance unit will be four times its molarity.



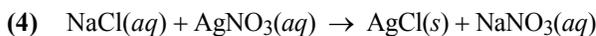
<u>Volume (added)</u>	<u>Conductance unit</u>	
0 L, KOH	$[\text{CH}_3\text{COOH}] = 1.0 \text{ M}$, (negligible ions, weak acid)	0 unit
1 L, KOH	$[\text{CH}_3\text{COOK}] = \frac{1.0 \text{ mol}}{2.0 \text{ L}} = 0.50 \text{ M}$, (CH_3COO^- , K^+)	1 unit
2 L, KOH	$[\text{CH}_3\text{COOK}] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} \text{ M}$, $[\text{KOH}] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} \text{ M}$, (K^+ , OH^-)	1.3 units



<u>Volume (added)</u>	<u>Conductance unit</u>	
0 L, NaOH	$[\text{HCl}] = 1.0 M, (\text{H}^+, \text{Cl}^-)$	2 units
1 L, NaOH	$[\text{NaCl}] = \frac{1.0 \text{ mol}}{2.0 \text{ L}} = 0.50 M, (\text{Na}^+, \text{Cl}^-)$	1 unit
2 L, NaOH	$[\text{NaCl}] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} M, [\text{NaOH}] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} M, (\text{Na}^+, \text{OH}^-)$	1.3 units



<u>Volume (added)</u>	<u>Conductance unit</u>	
0 L, BaCl ₂	$[\text{K}_2\text{SO}_4] = 1.0 M, (2\text{K}^+, \text{SO}_4^{2-})$	4 units
1 L, BaCl ₂	$[\text{KCl}] = \frac{2.0 \text{ mol}}{2.0 \text{ L}} = 1.0 M, (\text{K}^+, \text{Cl}^-)$	2 units
2 L, BaCl ₂	$[\text{KCl}] = \frac{2.0 \text{ mol}}{3.0 \text{ L}} = \frac{2}{3} M, [\text{BaCl}_2] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} M, (\text{Ba}^{2+}, 2\text{Cl}^-)$	2.7 units



<u>Volume (added)</u>	<u>Conductance unit</u>	
0 L, NaCl	$[\text{AgNO}_3] = 1.0 M, (\text{Ag}^+, \text{NO}_3^-)$	2 units
1 L, NaCl	$[\text{NaNO}_3] = \frac{1.0 \text{ mol}}{2.0 \text{ L}} = 0.50 M, (\text{Na}^+, \text{NO}_3^-)$	1 unit
2 L, NaCl	$[\text{NaNO}_3] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} M, [\text{NaCl}] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} M, (\text{Na}^+, \text{Cl}^-)$	1.3 units



<u>Volume (added)</u>	<u>Conductance unit</u>	
0 L, CH ₃ COOH	$[\text{NH}_3] = 1.0 M, (\text{negligible ions, weak base})$	0 unit
1 L, CH ₃ COOH	$[\text{CH}_3\text{COONH}_4] = \frac{1.0 \text{ mol}}{2.0 \text{ L}} = 0.50 M, (\text{CH}_3\text{COO}^-, \text{NH}_4^+)$	1 unit
2 L, CH ₃ COOH	$[\text{CH}_3\text{COONH}_4] = \frac{1.0 \text{ mol}}{3.0 \text{ L}} = \frac{1}{3} M$	0.67 unit

Matching this data to the diagrams shown, we find:

Diagram (a): Reactions (2) and (4)

Diagram (c): Reaction (3)

Diagram (b): Reaction (5)

Diagram (d): Reaction (1)