Chapter 4

Three Major Classes of Chemical Reactions
The Major Classes of Chemical Reactions

4.1 The Role of Water as a Solvent

4.2 Writing Equations for Aqueous Ionic Reactions

4.3 Precipitation Reactions

4.4 Acid-Base Reactions

4.5 Oxidation-Reduction (Redox) Reactions

4.6 Elements in Redox Reactions

4.7 Reaction Reversibility and the Equilibrium State
Water as a Solvent

- Water is a polar molecule
  - since it has uneven electron distribution
  - and a bent molecular shape.
- Water readily dissolves a variety of substances.
- Water interacts strongly with its solutes and often plays an active role in aqueous reactions.
Figure 4.1 Electron distribution in molecules of $\text{H}_2$ and $\text{H}_2\text{O}$.

A. Electron charge distribution in $\text{H}_2$ is symmetrical.

B. Electron charge distribution in $\text{H}_2\text{O}$ is asymmetrical.

C. Each bond in $\text{H}_2\text{O}$ is polar.

D. The whole $\text{H}_2\text{O}$ molecule is polar.
Figure 4.2  An ionic compound dissolving in water.
Figure 4.3 The electrical conductivity of ionic solutions.

A Distilled water does not conduct a current.

B Positive and negative ions fixed in a solid do not conduct a current.

C In solution, positive and negative ions move and conduct a current.
Sample Problem 4.1 Using Molecular Scenes to Depict an Ionic Compound in Aqueous Solution

**PROBLEM:** The beakers shown below contain aqueous solutions of the strong electrolyte potassium sulfate.

(a) Which beaker best represents the compound in solution? (H₂O molecules are not shown).

(b) If each particle represents 0.10 mol, what is the total number of particles in solution?
Sample Problem 4.1

PLAN:  (a) Determine the formula and write and equation for the dissociation of 1 mol of compound. Potassium sulfate is a strong electrolyte; it therefore dissociates completely in solution. *Remember that polyatomic ions remain intact in solution.*

(b) Count the number of separate particles in the relevant beaker, then multiply by 0.1 mol and by Avogadro’s number.

SOLUTION:

(a) The formula is K₂SO₄, so the equation for dissociation is:

\[ \text{K}_2\text{SO}_4 \ (s) \rightarrow 2\text{K}^+ \ (aq) + \text{SO}_4^{2-} \ (aq) \]
Sample Problem 4.1

There should be 2 cations for every 1 anion; beaker C represents this correctly.

(b) Beaker C contains 9 particles, 6 K$^+$ ions and 3 SO$_4^{2-}$ ions.

\[
9 \times 0.1 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol}} = 5.420 \times 10^{23} \text{ particles}
\]
Sample Problem 4.2  Determining Amount (mol) of Ions in Solution

PROBLEM: What amount (mol) of each ion is in each solution?

(a) 5.0 mol of ammonium sulfate dissolved in water
(b) 78.5 g of cesium bromide dissolved in water
(c) 7.42 \times 10^{22} \text{ formula units of copper(II) nitrate dissolved in water}
(d) 35 \text{ mL of 0.84 M zinc chloride}

PLAN: Write an equation for the dissociation of 1 mol of each compound. Use this information to calculate the actual number of moles represented by the given quantity of substance in each case.
Sample Problem 4.2

SOLUTION:

(a) The formula is (NH₄)₂SO₄ so the equation for dissociation is:

(NH₄)₂SO₄ (s) → 2NH₄⁺ (aq) + SO₄²⁻ (aq)

\[
\begin{align*}
5.0 \text{ mol (NH}_4\text{)}_2\text{SO}_4 \times \frac{2 \text{ mol NH}_4^+}{1 \text{ mol (NH}_4\text{)}_2\text{SO}_4} &= 10. \text{ mol NH}_4^+ \\
5.0 \text{ mol (NH}_4\text{)}_2\text{SO}_4 \times \frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol (NH}_4\text{)}_2\text{SO}_4} &= 5.0 \text{ mol NH}_4^+
\end{align*}
\]
Sample Problem 4.2

SOLUTION:

(b) The formula is CsBr so the equation for dissociation is:

\[ \text{CsBr (s)} \rightarrow \text{Cs}^+ \, (aq) + \text{Br}^- \, (aq) \]

\[
\frac{78.5 \text{ g CsBr}}{212.8 \text{ g CsBr}} \times \frac{1 \text{ mol CsBr}}{1 \text{ mol CsBr}} \times \frac{1 \text{ mol Cs}^+}{1 \text{ mol CsBr}} = 0.369 \text{ mol Cs}^+
\]

There is one Cs\(^+\) ion for every Br\(^-\) ion, so the number of moles of Br\(^-\) is also equal to 0.369 mol.
Sample Problem 4.2

SOLUTION:

(c) The formula is Cu(NO_3)_2 so the formula for dissociation is:

Cu(NO_3)_2 (s) → Cu^{2+} (aq) + 2NO_3^{-} (aq)

7.42 \times 10^{22} \text{ formula units Cu(NO}_3)_2 \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ formula units}} = 0.123 \text{ mol Cu(NO}_3)_2

0.123 \text{ mol Cu(NO}_3)_2 \times \frac{1 \text{ mol Cu}^{2+}}{1 \text{ mol Cu(NO}_3)_2} = 0.123 \text{ mol Cu}^{2+} \text{ ions}

There are 2 NO_3^{-} ions for every 1 Cu^{2+} ion, so there are 0.246 mol NO_3^{-} ions.
SOLUTION:

(d) The formula is ZnCl$_2$ so the formula for dissociation is:

\[
\text{ZnCl}_2 \text{ (s)} \rightarrow \text{Zn}^{2+} \text{ (aq)} + 2\text{Cl}^- \text{ (aq)}
\]

\[
35 \text{ mL soln} \times \frac{1 \text{ L}}{10^3 \text{ mL}} \times \frac{0.84 \text{ mol ZnCl}_2}{1 \text{ L soln}} = 2.9 \times 10^{-2} \text{ mol ZnCl}_2
\]

\[
2.9 \times 10^{-2} \text{ mol ZnCl}_2 \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol ZnCl}_2} = 5.8 \times 10^{-2} \text{ mol Cl}^-
\]

There is 1 mol of Zn$^{2+}$ ions for every 1 mol of ZnCl$_2$, so there are \(2.9 \times 10^{-2}\) mol Zn$^{2+}$ ions.
Writing Equations for Aqueous Ionic Reactions

The **molecular equation** shows all reactants and products as if they were *intact, undisassociated compounds*.

This gives the least information about the species in solution.

\[ 2\text{AgNO}_3 (aq) + \text{Na}_2\text{CrO}_4 (aq) \rightarrow \text{Ag}_2\text{CrO}_4 (s) + 2\text{NaNO}_3 (aq) \]  

When solutions of silver nitrate and sodium chromate mix, a brick-red precipitate of silver chromate forms.
The **total ionic equation** shows all soluble ionic substances *dissociated into ions*.

This gives the most accurate information about species in solution.

\[
2\text{Ag}^+ (aq) + 2\text{NO}_3^- (aq) \rightarrow \text{Ag}_2\text{CrO}_4 (s) \\
+ 2\text{Na}^+ (aq) + \text{CrO}_4^{2-} (aq) + 2\text{Na}^+ (aq) + \text{NO}_3^- (aq)
\]

**Spectator ions** are ions that are not involved in the actual chemical change. Spectator ions appear unchanged on both sides of the total ionic equation.

\[
2\text{Ag}^+ (aq) + 2\text{NO}_3^- (aq) \rightarrow \text{Ag}_2\text{CrO}_4 (s) \\
+ 2\text{Na}^+ (aq) + \text{CrO}_4^{2-} (aq) + 2\text{Na}^+ (aq) + 2\text{NO}_3^- (aq)
\]
The net ionic equation eliminates the spectator ions and shows only the actual chemical change.

\[ 2\text{Ag}^+ (aq) + \text{CrO}_4^{2-} (aq) \rightarrow \text{Ag}_2\text{CrO}_4 (s) \]
Figure 4.4  An aqueous ionic reaction and the three types of equations.

Molecular equation

\[ 2\text{AgNO}_3(aq) + \text{Na}_2\text{CrO}_4(aq) \rightarrow \text{Ag}_2\text{CrO}_4(s) + 2\text{NaNO}_3(aq) \]

Silver nitrate  Sodium chromate  Silver chromate  Sodium nitrate

Total ionic equation

\[ 2\text{Ag}^+(aq) + 2\text{NO}_3^-(aq) + 2\text{Na}^+(aq) + \text{CrO}_4^{2-}(aq) \rightarrow \text{Ag}_2\text{CrO}_4(s) + 2\text{Na}^+(aq) + 2\text{NO}_3^-(aq) \]

Net ionic equation

\[ 2\text{Ag}^+(aq) + \text{CrO}_4^{2-}(aq) \rightarrow \text{Ag}_2\text{CrO}_4(s) \]
Precipitation Reactions

• In a precipitation reaction two soluble ionic compounds react to give an insoluble products, called a precipitate.

• The precipitate forms through the net removal of ions from solution.

• It is possible for more than one precipitate to form in such a reaction.
Figure 4.5 The precipitation of calcium fluoride.

\[
2 \text{NaF (aq)} + \text{CaCl}_2 (aq) \rightarrow \text{CaF}_2 (s) + 2 \text{NaCl (aq)}
\]

\[
2 \text{Na}^+ (aq) + 2 \text{F}^- (aq) + \text{Ca}^{2+} (aq) + 2 \text{Cl}^- (aq) \rightarrow \text{CaF}_2(s) + 2 \text{Na}^+ (aq) + 2 \text{Cl}^- (aq)
\]

\[
2 \text{NaF(aq)} + \text{CaCl}_2 (aq) \rightarrow \text{CaF}_2(s) + 2 \text{NaCl (aq)}
\]
Precipitation reactions are also called **double displacement** reactions or **metathesis**.

\[
2\text{NaI}(aq) + \text{Pb(NO}_3\text{)}_2(aq) \rightarrow \text{PbI}_2(s) + \text{NaNO}_3(aq)
\]

\[
2\text{Na}^+(aq) + 2\text{I}^-(aq) + \text{Pb}^{2+}(aq) + 2\text{NO}_3^-(aq) \rightarrow \text{PbI}_2(s) + 2\text{Na}^+(aq) + 2\text{NO}_3^-(aq)
\]

\[
\text{Pb}^{2+}(aq) + 2\text{I}^-(aq) \rightarrow \text{PbI}_2(s)
\]

Ions exchange partners and a precipitate forms, so there is an exchange of bonds between reacting species.
Predicting Whether a Precipitate Will Form

• Note the ions present in the reactants.
• Consider all possible cation-anion combinations.
• Use the solubility rules to decide whether any of the ion combinations is insoluble.
  – An insoluble combination identifies the precipitate that will form.
Soluble Ionic Compounds

1. All common compounds of Group 1A(1) ions (Li⁺, Na⁺, K⁺, etc.) and ammonium ion (NH₄⁺) are soluble.
2. All common nitrates (NO₃⁻), acetates (CH₃COO⁻ or C₂H₃O₂⁻) and most perchlorates (ClO₄⁻) are soluble.
3. All common chlorides (Cl⁻), bromides (Br⁻) and iodides (I⁻) are soluble, except those of Ag⁺, Pb²⁺, Cu⁺, and Hg₂²⁺. All common fluorides (F⁻) are soluble except those of Pb²⁺ and Group 2A(2).
4. All common sulfates (SO₄²⁻) are soluble, except those of Ca²⁺, Sr²⁺, Ba²⁺, Ag⁺, and Pb²⁺.

Insoluble Ionic Compounds

1. All common metal hydroxides are insoluble, except those of Group 1A(1) and the larger members of Group 2A(2)(beginning with Ca²⁺).
2. All common carbonates (CO₃²⁻) and phosphates (PO₄³⁻) are insoluble, except those of Group 1A(1) and NH₄⁺.
3. All common sulfides are insoluble except those of Group 1A(1), Group 2A(2) and NH₄⁺.
Sample Problem 4.3  Predicting Whether a Precipitation Reaction Occurs; Writing Ionic Equations

PROBLEM: Predict whether or not a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, total ionic, and net ionic equations, and identify the spectator ions.

(a) potassium fluoride (aq) + strontium nitrate (aq) →
(b) ammonium perchlorate (aq) + sodium bromide (aq) →

PLAN: Note reactant ions, write the possible cation-anion combinations, and use Table 4.1 to decide if the combinations are insoluble. Write the appropriate equations for the process.
Sample Problem 4.3

SOLUTION:  (a) The reactants are KF and Sr(NO₃)₂. The possible products are KNO₃ and SrF₂. KNO₃ is soluble, but SrF₂ is an insoluble combination.

Molecular equation:

\[ 2\text{KF (aq)} + \text{Sr(NO}_3\text{)}_2 (aq) \rightarrow 2 \text{KNO}_3 (aq) + \text{SrF}_2 (s) \]

Total ionic equation:

\[ 2\text{K}^+ (aq) + 2\text{F}^- (aq) + \text{Sr}^{2+} (aq) + 2\text{NO}_3^- (aq) \rightarrow 2\text{K}^+ (aq) + 2\text{NO}_3^- (aq) + \text{SrF}_2 (s) \]

K⁺ and NO₃⁻ are spectator ions

Net ionic equation:

\[ \text{Sr}^{2+} (aq) + 2\text{F}^- (aq) \rightarrow \text{SrF}_2 (s) \]
Sample Problem 4.3

**SOLUTION:** (b) The reactants are \(\text{NH}_4\text{ClO}_4\) and \(\text{NaBr}\). The possible products are \(\text{NH}_4\text{Br}\) and \(\text{NaClO}_4\). Both are soluble, so no precipitate forms.

Molecular equation:

\[
\text{NH}_4\text{ClO}_4\ (aq) + \text{NaBr}\ (aq) \rightarrow \text{NH}_4\text{Br}\ (aq) + \text{NaClO}_4\ (aq)
\]

Total ionic equation:

\[
\text{NH}_4^+\ (aq) + \text{ClO}_4^-\ (aq) + \text{Na}^+\ (aq) + \text{Br}^-\ (aq) \rightarrow \text{NH}_4^+\ (aq) + \text{Br}^-\ (aq) + \text{Na}^+\ (aq) + \text{ClO}_4^-\ (aq)
\]

All ions are spectator ions and there is no net ionic equation.
PROBLEM: The following molecular views show reactant solutions for a precipitation reaction (with H$_2$O molecules omitted for clarity).

(a) Which compound is dissolved in beaker A: KCl, Na$_2$SO$_4$, MgBr$_2$, or Ag$_2$SO$_4$?
(b) Which compound is dissolved in beaker B: NH$_4$NO$_3$, MgSO$_4$, Ba(NO$_3$)$_2$, or CaF$_2$?
Sample Problem 4.4

PLAN:  Note the number and charge of each kind of ion and use Table 4.1 to determine the ion combinations that are soluble.

SOLUTION:

(a) Beaker A contains two 1+ ion for each 2- ion. Of the choices given, only Na$_2$SO$_4$ and Ag$_2$SO$_4$ are possible. Na$_2$SO$_4$ is soluble while Ag$_2$SO$_4$ is not.

**Beaker A therefore contains Na$_2$SO$_4$.**

(b) Beaker B contains two 1- ions for each 2+ ion. Of the choices given, only CaF$_2$ and Ba(NO$_3$)$_2$ match this description. CaF$_2$ is not soluble while Ba(NO$_3$)$_2$ is soluble.

**Beaker B therefore contains Ba(NO$_3$)$_2$.**
Sample Problem 4.4

**PROBLEM:** (c) Name the precipitate and spectator ions when solutions A and B are mixed, and write balanced molecular, total ionic, and net ionic equations for this process.

(d) If each particle represents 0.010 mol of ions, what is the maximum mass (g) of precipitate that can form (assuming complete reaction)?

**PLAN:** (c) Consider the cation-anion combinations from the two solutions and use Table 4.1 to decide if either of these is insoluble.

**SOLUTION:** The reactants are Ba(NO₃)₂ and Na₂SO₄. The possible products are BaSO₄ and NaNO₃. BaSO₄ is insoluble while NaNO₃ is soluble.
Sample Problem 4.4

Molecular equation:

\[ \text{Ba(NO}_3\text{)}_2 \ (aq) + \text{Na}_2\text{SO}_4 \ (aq) \rightarrow 2\text{NaNO}_3 \ (aq) + \text{BaSO}_4 \ (s) \]

Total ionic equation:

\[ \text{Ba}^{2+} \ (aq) + 2\text{NO}_3^{-} \ (aq) + 2\text{Na}^{+} \ (aq) + \text{SO}_4^{2-} \ (aq) \rightarrow 2\text{Na}^{+} \ (aq) + 2\text{NO}_3^{-} \ (aq) + \text{BaSO}_4 \ (s) \]

\[ \text{Na}^{+} \text{ and NO}_3^{-} \text{ are spectator ions} \]

Net ionic equation:

\[ \text{Ba}^{2+} \ (aq) + \text{SO}_4^{2-} \ (aq) \rightarrow \text{BaSO}_4 \ (s) \]
Sample Problem 4.4

**PLAN:**  (d) Count the number of each kind of ion that combines to form the solid. Multiply the number of each reactant ion by 0.010 mol and calculate the mol of product formed from each. Decide which ion is the limiting reactant and use this information to calculate the mass of product formed.

**SOLUTION:**  There are 4 Ba$^{2+}$ particles and 5 SO$_4^{2-}$ particles depicted.

\[
4 \text{ Ba}^{2+} \text{ particles} \times \frac{0.010 \text{ mol Ba}^{2+}}{1 \text{ particle}} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol Ba}^{2+}} = 0.040 \text{ mol BaSO}_4
\]

\[
4 \text{ SO}_4^{2-} \text{ particles} \times \frac{0.010 \text{ mol SO}_4^{2-}}{1 \text{ particle}} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol SO}_4^{2-}} = 0.050 \text{ mol BaSO}_4
\]
Sample Problem 4.4

Ba\(^{2+}\) ion is the limiting reactant, since it yields less BaSO\(_4\).

\[
0.040 \text{ mol BaSO}_4 \times \frac{233.4 \text{ g BaSO}_4}{1 \text{ mol BaSO}_4} = 9.3 \text{ g BaSO}_4
\]
Acid-Base Reactions

An **acid** is a substance that produces $\text{H}^+$ ions when dissolved in $\text{H}_2\text{O}$.  

$$\text{HX} \xrightarrow{\text{H}_2\text{O}} \text{H}^+ (aq) + \text{X}^- (aq)$$

A **base** is a substance that produces $\text{OH}^-$ ions when dissolved in $\text{H}_2\text{O}$.  

$$\text{MOH} \xrightarrow{\text{H}_2\text{O}} \text{M}^+ (aq) + \text{OH}^- (aq)$$

An **acid-base reaction** is also called a **neutralization** reaction.
Figure 4.7  The H\(^+\) ion as a solvated hydronium ion.

H\(^+\) interacts strongly with H\(_2\)O, forming H\(_3\)O\(^+\) in aqueous solution.
<table>
<thead>
<tr>
<th>Acids</th>
<th>Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Strong</strong></td>
<td><strong>Strong</strong></td>
</tr>
<tr>
<td>hydrochloric acid, HCl</td>
<td>sodium hydroxide, NaOH</td>
</tr>
<tr>
<td>hydrobromic acid, HBr</td>
<td>potassium hydroxide, KOH</td>
</tr>
<tr>
<td>hydriodic acid, HI</td>
<td>calcium hydroxide, Ca(OH)₂</td>
</tr>
<tr>
<td>nitric acid, HNO₃</td>
<td>strontium hydroxide, Sr(OH)₂</td>
</tr>
<tr>
<td>sulfuric acid, H₂SO₄</td>
<td>barium hydroxide, Ba(OH)₂</td>
</tr>
<tr>
<td>perchloric acid, HClO₄</td>
<td></td>
</tr>
<tr>
<td><strong>Weak</strong></td>
<td><strong>Weak</strong></td>
</tr>
<tr>
<td>hydrofluoric acid, HF</td>
<td>ammonia, NH₃</td>
</tr>
<tr>
<td>phosphoric acid, H₃PO₄</td>
<td></td>
</tr>
<tr>
<td>acetic acid, CH₃COOH (or HC₂H₃O₂)</td>
<td></td>
</tr>
</tbody>
</table>
Figure 4.8  Acids and bases as electrolytes.

Strong acids and strong bases dissociate completely into ions in aqueous solution.

They are **strong electrolytes** and conduct well in solution.

A Strong acid (or base) = strong electrolyte
Figure 4.8 Acids and bases as electrolytes.

Weak acids and weak bases dissociate very little into ions in aqueous solution.

They are *weak electrolytes* and conduct poorly in solution.
PROBLEM: How many $\text{H}^+(aq)$ ions are in 25.3 mL of 1.4 $M$ nitric acid?

PLAN: Use the volume and molarity to determine the mol of acid present. Since $\text{HNO}_3$ is a strong acid, moles acid = moles $\text{H}^+$. 

1. **volume of $\text{HNO}_3$**
   - Convert mL to L and multiply by $M$

2. **mol of $\text{HNO}_3$**
   - Mole of $\text{H}^+$ = mol of $\text{HNO}_3$

3. **mol of $\text{H}^+$**
   - Multiply by Avogadro’s number

4. **number of $\text{H}^+$ ions**
Sample Problem 4.5

**SOLUTION:**

$$\frac{35.3 \text{ mL soln}}{10^3 \text{ mL}} \times 1.4 \text{ mol HNO}_3 \times \frac{1 \text{ L soln}}{1 \text{ L soln}} = 0.035 \text{ mol HNO}_3$$

One mole of $H^+ (aq)$ is released per mole of nitric acid (HNO₃).

\[
\text{H}_2\text{O} \rightarrow \text{H}^+ (aq) + \text{NO}_3^- (aq)
\]

$$0.035 \text{ mol HNO}_3 \times \frac{1 \text{ mol H}^+}{1 \text{ mol HNO}_3} \times \frac{6.022 \times 10^{23} \text{ ions}}{1 \text{ mol}} = 2.1 \times 10^{22} \text{ H}^+ \text{ ions}$$
PROBLEM: Write balanced molecular, total ionic, and net ionic equations for the following acid-base reactions and identify the spectator ions.

(a) hydrochloric acid \((aq)\) + potassium hydroxide \((aq)\) →
(b) strontium hydroxide \((aq)\) + perchloric acid \((aq)\) →
(c) barium hydroxide \((aq)\) + sulfuric acid \((aq)\) →

PLAN: All reactants are strong acids and bases (see Table 4.2). The product in each case is H₂O and an ionic salt. Write the molecular reaction in each case and use the solubility rules to determine if the product is soluble or not.
Sample Problem 4.6

SOLUTION:

(a) hydrochloric acid (aq) + potassium hydroxide (aq) →

Molecular equation:
HCl (aq) + KOH (aq) → KCl (aq) + H₂O (l)

Total ionic equation:
H⁺ (aq) + Cl⁻ (aq) + K⁺ (aq) + OH⁻ (aq) → K⁺ (aq) + Cl⁻ (aq) + H₂O (l)

Net ionic equation:
H⁺ (aq) + OH⁻ (aq) → H₂O (l)

Spectator ions are K⁺ and Cl⁻
Sample Problem 4.6

SOLUTION:

(b) strontium hydroxide \((aq)\) + perchloric acid \((aq)\) →

Molecular equation:
\[
\text{Sr(OH)}_2 (aq) + 2\text{HClO}_4 (aq) \rightarrow \text{Sr(ClO}_4)_2 (aq) + 2\text{H}_2\text{O} (l)
\]

Total ionic equation:
\[
\text{Sr}^{2+} (aq) + 2\text{OH}^- (aq) + 2\text{H}^+ (aq) + 2\text{ClO}_4^- (aq) \rightarrow \text{Sr}^{2+} (aq) + 2\text{ClO}_4^- (aq) + 2\text{H}_2\text{O} (l)
\]

Net ionic equation:
\[
2\text{H}^+ (aq) + 2\text{OH}^- (aq) \rightarrow 2\text{H}_2\text{O} (l) \text{ or } \text{H}^+ (aq) + \text{OH}^- (aq) \rightarrow \text{H}_2\text{O} (l)
\]

Spectator ions are \text{Sr}^{2+} and \text{ClO}_4^-
Sample Problem 4.6

SOLUTION:

(c) barium hydroxide \((aq)\) + sulfuric acid \((aq)\) →

Molecular equation:
\[
\text{Ba(OH)}_2 \,(aq) + \text{H}_2\text{SO}_4 \,(aq) \rightarrow \text{BaSO}_4 \,(s) + 2\text{H}_2\text{O} \,(l)
\]

Total ionic equation:
\[
\text{Ba}^{2+} \,(aq) + 2\text{OH}^- \,(aq) + 2\text{H}^+ \,(aq) + \text{SO}_4^{2-} \,(aq) \rightarrow \text{BaSO}_4 \,(s) + \text{H}_2\text{O} \,(l)
\]

The net ionic equation is the same as the total ionic equation since there are no spectator ions.

This reaction is both a neutralization reaction and a precipitation reaction.
Figure 4.9  An aqueous strong acid-strong base reaction as a proton-transfer process.

H$_3$O$^+$ (aq) + X$^-$ (aq) + M$^+$ (aq) + OH$^-$ (aq) → mix → 2H$_2$O(l) + M$^+$ (aq) + X$^-$ (aq) → Δ → 2H$_2$O(g) + MX(s)

Aqueous solutions of strong acid and strong base are mixed, resulting in the formation of water and a salt (MX). The M$^+$ and X$^-$ ions remain in solution as spectator ions.
Figure 4.11 A gas-forming reaction with a weak acid.

Molecular equation

\[ \text{NaHCO}_3 (aq) + \text{CH}_3\text{COOH}(aq) \rightarrow \text{CH}_3\text{COONa} (aq) + \text{CO}_2 (g) + \text{H}_2\text{O} (l) \]

Total ionic equation

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

Net ionic equation

\[ \text{HCO}_3^- (aq) + \text{CH}_3\text{COOH} (aq) \rightarrow \text{CH}_3\text{COO}^- (aq) + \text{CO}_2 (g) + \text{H}_2\text{O} (l) \]
Sample Problem 4.7  Writing Proton-Transfer Equations for Acid-Base Reactions

PROBLEM: Write balanced total and net ionic equations for the following reactions and use curved arrows to show how the proton transfer occurs.

(a) hydriodic acid (aq) + calcium hydroxide (aq) →

   Give the name and formula of the salt present when the water evaporates.

(b) potassium hydroxide (aq) + propionic acid (aq) →

   Note that propionic acid is a weak acid. Be sure to identify the spectator ions in this reaction.
**Sample Problem 4.7**

**PLAN:** In (a) the reactants are a strong acid and a strong base. The acidic species is therefore $\text{H}_3\text{O}^+$, which transfers a proton to the $\text{OH}^-$ from the base.

**SOLUTION:**

**Total Ionic Equation:**

$$2\text{H}_3\text{O}^+ (aq) + 2\text{I}^- (aq) + \text{Ca}^{2+} (aq) + 2\text{OH}^- (aq) \rightarrow 2\text{I}^- (aq) + \text{Ca}^{2+} (aq) + 4\text{H}_2\text{O} (l)$$

**Net Ionic Equation:**

$$\text{H}_3\text{O}^+ (aq) + \text{OH}^- (aq) \rightarrow + \text{H}_2\text{O} (l)$$

**When the water evaporates, the salt remaining is CaI$_2$, calcium iodide.**
Sample Problem 4.7

PLAN: In (b) the acid is weak; therefore it does not dissociate much and largely exists as intact molecules in solution.

SOLUTION:

Total Ionic Equation:

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

Net Ionic Equation:

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

\( \text{K}^+ \) is the only spectator ion in the reaction.
Acid-Base Titrations

• In a **titration**, the concentration of one solution is used to determine the concentration of another.
• In an acid-base titration, a standard solution of base is usually added to a sample of acid of unknown molarity.
• An **acid-base indicator** has different colors in acid and base, and is used to monitor the reaction progress.
• At the **equivalence point**, the mol of H\(^+\) from the acid equals the mol of OH\(^-\) ion produced by the base.
  – Amount of H\(^+\) ion in flask = amount of OH\(^-\) ion added
• The **end point** occurs when there is a slight excess of base and the indicator changes color permanently.
Figure 4.11  An acid-base titration.

Before titration

Temporary excess of OH⁻

H₂O(l) + M⁺(aq) + X⁻(aq)

Permanent slight excess of OH⁻
Sample Problem 4.8  Finding the Concentration of Acid from a Titration

**PROBLEM:** A 50.00 mL sample of HCl is titrated with 0.1524 \( M \) NaOH. The buret reads 0.55 mL at the start and 33.87 mL at the end point. Find the molarity of the HCl solution.

**PLAN:** Write a balanced equation for the reaction. Use the volume of base to find mol OH\(^-\), then mol H\(^+\) and finally \( M \) for the acid.

- Volume of base (difference in buret readings)
- Multiply by \( M \) of base
- Mol of OH\(^-\)
- Use mole ratio as conversion factor
- Mol of H\(^+\) and acid
- Divide by volume (L) of acid
- Molarity (\( M \)) of acid
Sample Problem 4.8

**SOLUTION:** \[ \text{NaOH (aq)} + \text{HCl (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)} \]

Volume of base = 33.87 mL – 0.55 mL = 33.32 mL

\[
\begin{aligned}
33.32 \text{ mL soln} & \times \frac{1 \text{ L}}{10^3 \text{ mL}} & \times \frac{0.1524 \text{ mol NaOH}}{1 \text{ L soln}} &= 5.078 \times 10^{-3} \text{ mol NaOH}
\end{aligned}
\]

Since 1 mol of HCl reacts with 1 mol NaOH, the amount of HCl

\[
\begin{aligned}
\frac{5.078 \times 10^{-3} \text{ mol HCl}}{50.00 \text{ mL}} & \times \frac{10^3 \text{ mL}}{1 \text{ L}} &= 0.1016 \text{ M HCl}
\end{aligned}
\]
Oxidation is the *loss* of electrons. The *reducing agent* loses electrons and is oxidized.

Reduction is the *gain* of electrons. The *oxidizing agent* gains electrons and is reduced.

*A redox reaction* involves *electron transfer* Oxidation and reduction occur together.
Figure 4.12 The redox process in compound formation.

A Formation of an ionic compound

Mg from Mg(s) + O from O_2(g) → 2e^- → Mg^{2+} + O^{2-} → Sample of the ionic compound MgO

B Formation of a covalent compound

Electrons distributed evenly: H·H + Cl·Cl → Shift of electrons: H·H + Cl·Cl → Electrons distributed unevenly: H·Cl + Cl·H → Molecules of the covalent compound HCl
### Table 4.3  Rules for Assigning an Oxidation Number (O.N.)

#### General rules

1. For an atom in its elemental form (Na, O₂, Cl₂, etc.): O.N. = 0
2. For a monoatomic ion: O.N. = ion charge
3. The sum of O.N. values for the atoms in a compound equals zero. The sum of O.N. values for the atoms in a polyatomic ion equals the ion’s charge.

#### Rules for specific atoms or periodic table groups

1. For Group 1A(1): O.N. = +1 in all compounds
2. For Group 2A(2): O.N. = +2 in all compounds
3. For hydrogen: O.N. = +1 in combination with nonmetals
4. For fluorine: O.N. = -1 in combination with metals and boron
5. For oxygen:
   - O.N. = -1 in peroxides
   - O.N. = -2 in all other compounds (except with F)
6. For Group 7A(17): O.N. = -1 in combination with metals, nonmetals (except O), and other halogens lower in the group
Sample Problem 4.9  Determining the Oxidation Number of Each Element in a Compound (or Ion)

PROBLEM: Determine the oxidation number (O.N.) of each element in these species:
(a) zinc chloride  (b) sulfur trioxide  (c) nitric acid

PLAN: The O.N.s of the ions in a polyatomic ion add up to the charge of the ion and the O.N.s of the ions in the compound add up to zero.

SOLUTION:
(a) ZnCl₂. The O.N. for zinc is +2 and that for chloride is -1.

(b) SO₃. Each oxygen is an oxide with an O.N. of -2. The O.N. of sulfur must therefore be +6.

(c) HNO₃. H has an O.N. of +1 and each oxygen is -2. The N must therefore have an O.N. of +5.
Sample Problem 4.10

Identifying Redox Reactions

PROBLEM: Use oxidation numbers to decide whether each of the following is a redox reaction or not.

(a) \[ \text{CaO (s)} + \text{CO}_2 (g) \rightarrow \text{CaCO}_3 (s) \]

(b) \[ 4 \text{KNO}_3 (s) \rightarrow 2 \text{K}_2 \text{O(s)} + 2 \text{N}_2 (g) + 5 \text{O}_2 (g) \]

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

PLAN: Use Table 4.3 to assign an O.N. to each atom. A change in O.N. for any atom indicates electron transfer.

SOLUTION: (a) \[ \text{CaO(s)} + \text{CO}_2 (g) \rightarrow \text{CaCO}_3 (s) \]

\[
\begin{array}{c}
+2 \uparrow \downarrow \downarrow \downarrow \\
\text{CaO} & \text{CO}_2 & \text{CaCO}_3 \\
+2 & -2 & +2
\end{array}
\]

This is not a redox reaction, since no species change O.N.
Sample Problem 4.10

(b) \[ 4 \text{KNO}_3(s) \rightarrow 2 \text{K}_2\text{O}(s) + 2 \text{N}_2(g) + 5 \text{O}_2(g) \]

This is a redox reaction.
N changes O.N. from +5 to 0 and is reduced.
O changes O.N. from -2 to 0 and is oxidized.
Sample Problem 4.10

(c) \( \text{NaHSO}_4(aq) + \text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \)

This is not a redox reaction since no species change O.N.
Figure 4.13  A summary of terminology for redox reactions.

- **X** loses electron(s).
- **X** is oxidized.
- **X** is the reducing agent.
- **X** increases its oxidation number.

- **Y** gains electron(s).
- **Y** is reduced.
- **Y** is the oxidizing agent.
- **Y** decreases its oxidation number.
Sample Problem 4.11 Identifying Oxidizing and Reducing Agents

PROBLEM: Identify the oxidizing agent and reducing agent in each of the following reactions:

(a) \(2\text{Al} (s) + 3\text{H}_2\text{SO}_4 (aq) \rightarrow \text{Al}_2(\text{SO}_4)_3 (aq) + 3\text{H}_2(g)\)

(b) \(\text{PbO} (s) + \text{CO} (g) \rightarrow \text{Pb} (s) + \text{CO}_2 (g)\)

(c) \(2\text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2\text{H}_2\text{O} (g)\)

PLAN: Assign an O.N. to each atom and look for those that change during the reaction.

The reducing agent contains an atom that is oxidized (increases in O.N.) while the oxidizing agent contains an atom that is reduced (decreases in O.N.).
Sample Problem 4.11

SOLUTION:

(a) \[ 2\text{Al}(s) + 3\text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 3\text{H}_2(g) \]

\[ \begin{array}{cccc}
0 & +1 & +6 & 0 \\
-2 & -2 & +3 & -2 \\
\end{array} \]

Al changes O.N. from 0 to +3 and is oxidized. Al is the reducing agent.

H changes O.N. from +1 to 0 and is reduced. \( \text{H}_2\text{SO}_4 \) is the oxidizing agent.
Sample Problem 4.11

SOLUTION:

(b) \( \text{PbO (s)} + \text{CO (g)} \rightarrow \text{Pb (s)} + \text{CO}_2 (g) \)

\[
\begin{array}{c}
\text{Pb} & \quad & \text{O.N.} & \quad & \text{CO} & \quad & \text{O.N.} & \quad & \text{PbO} & \quad & \text{CO}_2 \\
\text{+2} & \quad & \rightarrow & \quad & \text{-2} & \quad & \text{+4} & \quad & \text{-2} & \quad & \text{-2} \\
\end{array}
\]

\text{Pb changes O.N. from +2 to 0 and is reduced.}
\text{PbO is the oxidizing agent.}

\text{C changes O.N. from +2 to +4 and is oxidized.}
\text{CO is the reducing agent.}
Sample Problem 4.11

SOLUTION:

\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \]

\(0\) \hspace{1cm} \(0\) \hspace{1cm} +1 \hspace{1cm} -2

\(H_2\) changes O.N. from 0 to +1 and is oxidized. 
\(H_2\) is the reducing agent.

\(O\) changes O.N. from 0 to -2 and is reduced. 
\(O_2\) is the oxidizing agent.
Balancing Redox Equations (oxidation number method)

1. Assign O.N.s to all atoms.
2. Identify the reactants that are oxidized and reduced.
3. Compute the numbers of electrons transferred, and draw tie-lines from each reactant atom to the product atom to show the change.
4. Multiply the numbers of electrons by factor(s) that make the electrons lost equal to the electrons gained.
5. Use the factor(s) as balancing coefficients.
6. Complete the balancing by inspection and add states of matter.
Sample Problem 4.12 Balancing Redox Equations by the Oxidation Number Method

PROBLEM: Use the oxidation number method to balance the following equations:

(a) \( \text{Cu} (s) + \text{HNO}_3 (aq) \rightarrow \text{Cu(NO}_3)_2 (aq) + \text{NO}_2 (g) + \text{H}_2\text{O} (l) \)

SOLUTION:
Assign oxidation numbers and identify oxidized and reduced species:

(a) \( \text{Cu} (s) + \text{HNO}_3 (aq) \rightarrow \text{Cu(NO}_3)_2 (aq) + \text{NO}_2 (g) + \text{H}_2\text{O} (l) \)

\[
\begin{array}{cccc}
\text{Cu} & \text{HNO}_3 & \text{Cu(NO}_3)_2 & \text{NO}_2 & \text{H}_2\text{O} \\
0 & +1 & +2 & +4 & +1 \\
+5 & +2 & +5 & -2 & -2 \\
-2 & -2 & -2 & -2 & -2 \\
\end{array}
\]
Sample Problem 4.12

\[ \text{loses } 2e^-; \text{ oxidation} \]
\[
\begin{align*}
\text{Cu}(s) + \text{HNO}_3(aq) & \rightarrow \text{Cu(NO}_3)_2(aq) + \text{NO}_2(g) + \text{H}_2\text{O}(l) \\
\end{align*}
\]

\[ \text{gains } 1e^-; \text{ reduction} \]

Multiply to make e\(^-\) lost = e\(^-\) gained:

\[
\begin{align*}
\text{Cu} (s) + 2\text{HNO}_3 (aq) & \rightarrow \text{Cu(NO}_3)_2 (aq) + 2\text{NO}_2 (g) + \text{H}_2\text{O} (l) \\
\end{align*}
\]

Balance other atoms by inspection:

\[
\begin{align*}
\text{Cu} (s) + 4\text{HNO}_3 (aq) & \rightarrow \text{Cu(NO}_3)_2 (aq) + 2\text{NO}_2 (g) + 2\text{H}_2\text{O} (l) \\
\end{align*}
\]
(b) \( \text{PbS (s)} + \text{O}_2 (g) \rightarrow \text{PbO (s)} + \text{SO}_2 (g) \)

**SOLUTION:**

Assign oxidation numbers and identify oxidized and reduced species:
Sample Problem 4.12

\[ \text{PbS (s) + O}_2 (g) \rightarrow \text{PbO (s) + SO}_2 (g) \]

*loses 6e\(^{-}\); oxidation*

\[ \text{gains 2e}^{- \text{ per O}; \text{ reduction}} \]

Multiply to make e\(^{-}\) lost = e\(^{-}\) gained:

\[ \text{PbS (s) + } \frac{3}{2} \text{O}_2 (g) \rightarrow \text{PbO (s) + SO}_2 (g) \]

Balance other atoms by inspection, and multiply to give whole-number coefficients:

\[ 2\text{PbS (s) + 3O}_2 (g) \rightarrow 2\text{PbO (s) + 2SO}_2 (g) \]
Figure 4.14  The redox titration of $\text{C}_2\text{O}_4^{2-}$ with $\text{MnO}_4^{-}$

Net ionic equation:

$$2\text{MnO}_4^{-}(aq) + 5\text{C}_2\text{O}_4^{2-}(aq) + 16\text{H}^{+}(aq) \longrightarrow 2\text{Mn}^{2+}(aq) + 10\text{CO}_2(g) + 8\text{H}_2\text{O}(l)$$
Sample Problem 4.13  Finding the Amount of Reducing Agent by Titration

PROBLEM: To measure the Ca$^{2+}$ concentration in human blood, 1.00 mL of blood was treated with Na$_2$C$_2$O$_4$ solution to precipitate the Ca$^{2+}$ as CaC$_2$O$_4$. The precipitate was filtered and dissolved in dilute H$_2$SO$_4$ to release C$_2$O$_4^{2-}$, which was titrated with KMnO$_4$ solution. The solution required 2.05mL of 4.88x10$^{-4}$ M KMnO$_4$ to reach the end point. The balanced equation is

\[
2 \text{KMnO}_4(aq) + 5 \text{CaC}_2\text{O}_4(s) + 8 \text{H}_2\text{SO}_4(aq) \rightarrow 2\text{MnSO}_4(aq) + \text{K}_2\text{SO}_4(aq) + 5\text{CaSO}_4(s) + 10\text{CO}_2(g) + 8\text{H}_2\text{O}(l)
\]

Calculate the amount (mol) of Ca$^{2+}$ in 1.00 mL of blood.
Sample Problem 4.13

**PLAN:** Calculate the mol of KMnO₄ from the volume and molarity of the solution. Use this to calculate the mol of C₂O₄²⁻ and hence the mol of Ca²⁺ ion in the blood sample.

1. **volume of KMnO₄ soln**
2. Convert mL to L and multiply by $M$
3. **mol of KMnO₄**
4. **molar ratio**
5. **mol of CaC₂O₄**
6. Ratio of elements in formula
7. **mol of Ca²⁺**
Sample Problem 4.13

SOLUTION:

\[
\frac{2.05 \text{ mL soln}}{10^3 \text{ mL}} \times \frac{1 \text{ L}}{4.88 \times 10^{-4} \text{ mol KMnO}_4} = 1.00 \times 10^{-6} \text{ mol KMnO}_4
\]

\[
1.00 \times 10^{-6} \text{ mol KMnO}_4 \times \frac{5 \text{ mol CaC}_2\text{O}_4}{2 \text{ mol KMnO}_4} = 2.50 \times 10^{-6} \text{ mol CaC}_2\text{O}_4
\]

\[
2.50 \times 10^{-6} \text{ mol CaC}_2\text{O}_4 \times \frac{1 \text{ mol Ca}^{2+}}{1 \text{ mol CaC}_2\text{O}_4} = 2.50 \times 10^{-6} \text{ mol Ca}^{2+}
\]
Elements in Redox Reactions
Types of Reaction

• Combination Reactions
  – Two or more reactants combine to form a new compound:
  – \( X + Y \rightarrow Z \)

• Decomposition Reactions
  – A single compound decomposes to form two or more products:
  – \( Z \rightarrow X + Y \)

• Displacement Reactions
  – double displacement: \( AB + CD \rightarrow AC + BD \)
  – single displacement: \( X + YZ \rightarrow XZ + Y \)

• Combustion
  – the process of combining with \( O_2 \)
Figure 4.15  Combining elements to form an ionic compound.

Macroscopic level

Atomic level

Symbolic level

$$\text{2K(s)} + \text{Cl}_2(g) \rightarrow \text{2KCl(s)}$$

2K(s) Potassium + Cl$_2$(g) Chlorine → 2KCl(s) Potassium chloride
Figure 4.16 Decomposition of the compound mercury(II) oxide to its elements.
Figure 4.17 The active metal lithium displaces \( \text{H}_2 \) from water.
Figure 4.18 The displacement of $H_2$ from acid by nickel.

$$\text{Ni (s)} + 2\text{H}^+ (aq) \rightarrow \text{Ni}^{2+} (aq) + \text{H}_2 (g)$$

O.N. increasing
oxidation occurring
reducing agent

O.N. decreasing
reduction occurring
oxidizing agent
Figure 4.19 A more reactive metal (Cu) displacing the ion of a less reactive metal (Ag⁺) from solution.

\[
\begin{align*}
2\text{AgNO}_3(aq) + \text{Cu}(s) &\rightarrow \text{Cu(NO}_3)_2(aq) + 2\text{Ag}(s) \\
\end{align*}
\]
Figure 4.20  The activity series of the metals.

Copyright © The McGraw-Hill Companies, Inc. Permission required for reproduction or display.

Li  K  Ba  Ca  Na  Mg  Al  Mn  Zn  Cr  Fe  Cd  Co  Ni  Sn  Pb

Strength as reducing agent

Can displace H₂ from water

Ba(s) + 2H₂O(l) → Ba²⁺(aq) + 2OH⁻(aq) + H₂(g)
(also see Figure 4.18)

Can displace H₂ from steam

Zn(s) + 2H₂O(g) → Zn(OH)₂(s) + H₂(g)

Can displace H₂ from acid

Sn(s) + 2H⁺(aq) → Sn²⁺(aq) + H₂(g)
(also see Figure 4.19)

H₂

 Cannot displace H₂ from any source

Ag(s) + 2H⁺(aq) → no reaction
Sample Problem 4.14  Identifying the Type of Redox Reaction

PROBLEM: Classify each of the following redox reactions as a combination, decomposition, or displacement reaction. Write a balanced molecular equation for each, as well as total and net ionic equations for part (c), and identify the oxidizing and reducing agents:

(a) magnesium (s) + nitrogen (g) $\rightarrow$ magnesium nitride (aq)

(b) hydrogen peroxide (l) $\rightarrow$ water (l) + oxygen gas

(c) aluminum (s) + lead(II) nitrate (aq) $\rightarrow$ aluminum nitrate (aq) + lead (s)

PLAN: Combination reactions combine reactants, decomposition reactions involve more products than reactants and displacement reactions have the same number of reactants and products.
Sample Problem 4.14

SOLUTION:

(a) This is a combination reaction, since Mg and N\textsubscript{2} combine:

\[3\text{Mg (s)} + \text{N}_2 (g) \rightarrow \text{Mg}_3\text{N}_2 (s)\]

Mg is the reducing agent; N\textsubscript{2} is the oxidizing agent.
(b) This is a decomposition reaction, since \( \text{H}_2\text{O}_2 \) breaks down:

\[
2 \text{H}_2\text{O}_2 (l) \rightarrow + 2\text{H}_2\text{O} (l) + \text{O}_2 (g)
\]

\[
\begin{array}{ccc}
+1 & +1 & 0 \\
-2 & -2 & -2 \\
\end{array}
\]

H\(_2\)O\(_2\) is \textit{both} the reducing and the oxidizing agent.
Sample Problem 4.14

(c) This is a displacement reaction, since Al displaces Pb$^{2+}$ from solution.

$$2\text{Al} \ (s) + 3\text{Pb(NO}_3\text{)}_2 \ (aq) \rightarrow 2\text{Al(NO}_3\text{)}_3 \ (aq) + 3\text{Pb} \ (s)$$

The net ionic equation is:

$$2\text{Al} \ (s) + 3\text{Pb}^{2+} \ (aq) \rightarrow 2\text{Al}^{3+} \ (aq) + 3\text{Pb} \ (s)$$

Al is the reducing agent; Pb(NO$_3$)$_2$ is the oxidizing agent.

The total ionic equation is:

$$2\text{Al} \ (s) + 3\text{Pb}^{2+} \ (aq) + 2\text{NO}_3^- \ (aq) \rightarrow 2\text{Al}^{3+} \ (aq) + 3\text{NO}_3^- \ (aq) + 3\text{Pb} \ (s)$$
Figure 4.21 The equilibrium state.

A Nonequilibrium system

CaCO₃(s) \xrightarrow{\Delta} \text{Reaction goes to completion because CO}_2 \text{ escapes.} \rightarrow \text{CO}_2 \text{ forms and escapes.} \rightarrow \text{CaO}

CaO(s) + CO₂(g)

B Equilibrium system

CaCO₃(s) \xrightarrow{\Delta} \text{Reaction reaches equilibrium because all substances are present.} \rightarrow \text{Mixture of CaO and CaCO}_3

CaCO₃(s) \xrightarrow{\text{CaO(s) + CO}_2(g)}