Chapter 4
Types of Chemical Reactions

- 4-1 Dissolution Reactions
- 4-2 Precipitation Reactions
- 4-3 Acids and Bases and Their Reactions
- 4-4 Oxidation-Reduction Reactions
1. **Dissolution**
   - Ionic Compounds in Water
   - Molecular Compounds in Water
   - Electrolytes and Non-Electrolytes
   - Solubilities (like dissolves like)
   - Predicting Dissolution Reactions

2. **Precipitation Reactions**
   - Ionic Equations and Net Ionic Equations
   - Predicting Precipitation Reactions

3. **Acids and Bases**
   - Arrhenius Acids and Bases Theory
   - Strong and Weak acid
   - Naming Acids
   - Weak Bases
   - Modifying the Arrhenius model
   - Acid-Base Titrations

4. **Oxidation-Reduction Reactions (Redox) Reactions**
   - Oxidizing and Reducing Agents
   - Oxidation Number
   - Types of Redox Reactions
     » Combination and Decomposition
     » Oxygenation
     » Hydrogenation
     » Displacement Reactions
     » Disproportionation
Dissolution: Two (or more) substances spread out, or disperse, into each other at the level of individual atoms, molecules, or ions.

Solution:

Solvent:

Solute:

In principle, the solute and solvent can be any combination of solid (s), liquid (l), and gaseous (g) phases.
Ionic Compounds in Water:

All ionic compounds are solids in the range of temperature in which water is a liquid. They have rigid lattices in which strong forces (ionic bonds) pin the constituent ions in place.

The high melting points of ionic compounds indicate that a good deal of energy must be supplied to destroy the lattice and produce a liquid (molten) form in which the ions move more freely.

Aquation:
Ionic Compounds in Water:

Solute is said to **dissociate** into ions or to **ionize** upon dissociation.
Molecular Compounds in Water:

Molecular substances (covalent) contain no ions to which water molecules can adhere, but their molecules are often polar.

Example: members of the class of carbohydrates known as sugars have the general formula, $C_m(H_2O)_n$.

Typical sugars include sucrose, $C_{12}H_{22}O_{11}$ (table sugar); fructose, $C_6H_{12}O_6$ (fruit sugar); and ribose, $C_5H_{10}O_5$ (a subunit in the biomolecules ribonucleic acids).

$$C_6H_{12}O_6(s) \rightarrow C_6H_{12}O_6(aq)$$
Solubilities:

Miscible:

Solubility: The largest amount that can dissolve in a given amount of a solvent at a particular temperature.
• Pairs of liquids that mix in any proportion are termed miscible. Liquids that do not mix are termed immiscible
• "Like dissolves like"
  – *in other words, substances with similar intermolecular attractive forces tend to be soluble in one another*
“like dissolves like” concept

- Fructose
- Glucose
- Water
- Ethanol
- Alanine
- Sucrose
“like dissolves like” concept

- Octane
- Isobutane
- Benzene
- Cyclohexane
- Octanol
- Carbon tetrachloride

Non-Polar
Electrolytes and Non-Electrolytes:

Electrolytes:

Non-Electrolytes:

Strong electrolytes:
Predicting Dissolution Reactions

**TABLE 4-1**

<table>
<thead>
<tr>
<th>Anion</th>
<th>Soluble$^a$</th>
<th>Slightly Soluble</th>
<th>Insoluble</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{NO}_3^-$ (nitrate)</td>
<td>All</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{CH}_3\text{COO}^-$ (acetate)</td>
<td>Most</td>
<td></td>
<td>Be($\text{CH}_3\text{COO}$)$_2$</td>
</tr>
<tr>
<td>$\text{ClO}_3^-$ (chlorate)</td>
<td>All</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{ClO}_4^-$ (perchlorate)</td>
<td>Most</td>
<td>K$\text{ClO}_4$</td>
<td></td>
</tr>
<tr>
<td>$\text{F}^-$ (fluoride)</td>
<td>Group I, AgF, BeF$_2$</td>
<td>SrF$_2$, BaF$_2$, PbF$_2$</td>
<td>MgF$_2$, CaF$_2$</td>
</tr>
<tr>
<td>$\text{Cl}^-$ (chloride)</td>
<td>Most</td>
<td>PbCl$_2$</td>
<td>AgCl, Hg$_2$Cl$_2$</td>
</tr>
<tr>
<td>$\text{Br}^-$ (bromide)</td>
<td>Most</td>
<td>PbBr$_2$, HgBr$_2$</td>
<td>AgBr, Hg$_2$Br$_2$</td>
</tr>
<tr>
<td>$\text{I}^-$ (iodide)</td>
<td>Most</td>
<td></td>
<td>AgI, Hg$_2$I$_2$, Pbl$_2$, Hgl$_2$</td>
</tr>
<tr>
<td>$\text{SO}_4^{2-}$ (sulfate)</td>
<td>Most</td>
<td>CaSO$_4$, Ag$_2$SO$_4$, Hg$_2$SO$_4$</td>
<td>SrSO$_4$, BaSO$_4$, PbsO$_4$</td>
</tr>
<tr>
<td>$\text{S}^{2-}$ (sulfide)</td>
<td>Groups I and II$^c$</td>
<td>(NH$_4$)$_2$S</td>
<td>Most</td>
</tr>
<tr>
<td>$\text{CO}_3^{2-}$ (carbonate)</td>
<td>Group I, (NH$_4$)$_2$CO$_3$</td>
<td></td>
<td>Most</td>
</tr>
<tr>
<td>$\text{SO}_3^{2-}$ (sulfite)</td>
<td>Group I, (NH$_4$)$_2$SO$_3$</td>
<td></td>
<td>Most</td>
</tr>
<tr>
<td>$\text{PO}_4^{3-}$ (phosphate)</td>
<td>Group I, (NH$_4$)$_3$PO$_4$</td>
<td>Li$_3$PO$_4$</td>
<td>Most</td>
</tr>
<tr>
<td>$\text{OH}^-$ (hydroxide)</td>
<td>Group I, Ba(OH)$_2$</td>
<td>Sr(OH)$_2$, Ca(OH)$_2$</td>
<td>Most</td>
</tr>
</tbody>
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$^a$Soluble compounds have solubilities exceeding 1 g/100 g water. Slightly soluble compounds have solubilities between 0.01 and 1 g/100 g; insoluble compounds have solubilities less than 0.01 g/100 g at room temperature.

$^b$Compounds of elements from the first column in the periodic table: Li, Na, K, Rb, Cs.

$^c$Compounds of elements from the second column in the periodic table: Be, Mg, Ca, Sr, Ba.
Solubilities:

Exercise 4-1:

Predict whether the following substances are soluble in water:

(a) calcium carbonate (CaCO₃)

(b) mercury(II) sulfide (HgS)

(c) isopropanol (C₃H₇OH)
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<td>Li(_3)PO(_4)</td>
<td></td>
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<td>Group I, Ba(OH)(_2)</td>
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\(^a\)Soluble compounds are defined as those that dissolve to the extent of 1 g or more per 100 g water, slightly soluble as 0.01 to 1 g per 100 g water, and insoluble as less than 0.01 g per 100 g water at room temperature.

\(^b\)Compounds of elements from the first column in the periodic table: Li, Na, K, Rb, Cs.

\(^c\)Compounds of elements from the second column in the periodic table: Be, Mg, Ca, Sr, Ba.
Solubility Rules

1. All common sodium, potassium and ammonium salts are soluble in water.
2. The chloride and iodide salts of all common metals, except silver, lead and mercury (I), are soluble in water. Lead chloride is soluble in hot water.
3. The sulfates of all metals, excepts lead, mercury (I), barium, strontium, and calcium are soluble in water.
4. The carbonates of all metals, excepts those of Group 1A and ammonium carbonate, are insoluble in water.
5. Most metal hydroxides are insoluble in water. However, the hydroxides of group 1A are soluble and those of Group II A are moderately soluble.
Precipitation Reactions:

Whenever the concentration of a substance in solution exceeds its solubility, a new phase starts to separate. A new solid phase is usually dense enough to sink to the bottom of a liquid solution.

Solids usually precipitate ("fall down") from their solution.
Ionic Equations and Net Ionic Equations:

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

Ionic equation

Spectator ions

Net ionic equation
Exercise 4-2:

Write a net ionic equation to represent the formation of the precipitate observed when aqueous solutions of CaCl₂ and NaF are mixed. Identify the spectator ions in this process.
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Acids and Bases and Their Reactions

Acid Base Theory

1. Arrhenius Acids and Bases (Chapter 4)
   - Acids are H\(^+\) donors
   - Bases are OH\(^-\) donors

2. Arrhenius Broadened Definition (Chapter 4)
   - Acids increase H\(^+\) concentration or \([H^+]\) increases
   - Bases increase OH\(^-\) concentration or \([OH^-]\) increases

3. Brønsted Acids and Bases (Chapter 8)
   - Acids donate H\(^+\)
   - Bases accept H\(^+\)

4. Lewis Acids and Bases (Chapter 8)
   - Acids are electron pair acceptors
   - Bases are electron pair donors
Acids and Bases and Their Reactions:
Arrhenius Acids and Bases:

An **Arrhenius acid** is a substance that, when dissolved in water, delivers hydrogen ions \( \text{H}^+(aq) \) ions) to the solution; it is a hydrogen-ion donor.

An **Arrhenius base** is a substance that, when dissolved in water, delivers hydroxide ions \( \text{OH}^-(aq) \) ions) to the solution; it is a hydroxide-ion donor.

Water is simultaneously and equally both an acid and a base:
Arrhenius Acids and Bases:

Strong acids:

Strong bases:
Arrhenius Acids and Bases:

Neutralization reaction:

\[ \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]
Weak Acids

\[
\text{CH}_3\text{COOH}(aq) \rightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}^+(aq) \quad \text{(dissociation of acetic acid)}
\]

\[
\text{CH}_3\text{COOH}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCH}_3\text{COO}(aq) + \text{H}_2\text{O}(l) \quad \text{(overall equation)}
\]
Naming Acids

Binary acids:

• HCl as covalent compound is *hydrogen* chloride
• HCl as a binary acid is *hydrochloric* Acid

Oxoacids:

• \( \text{SO}_4^{2-} \) or Sulfate ion
• \( \text{H}_2\text{SO}_4 \) becomes Sulfuric Acid

Oxoacids:

• \( \text{SO}_3^{-2} \) or Sulfite ion
• \( \text{H}_2\text{SO}_3 \) becomes Sulfurous Acid

Organic acids:

• \( \text{HCOO}^- \) or Formate ion becomes Formic Acid

\( \text{R} = \text{H} \)
Modifying the Arrhenius Model

An **Arrhenius base** is a substance that, when dissolved in water, increases the concentration of **hydroxide ions** over what is present in the pure solvent.

An **Arrhenius acid** is a substance that, when dissolved in water, increases the concentration of **hydrogen ions** over what is present in the pure solvent.
Acid-Base Titrations

**Titration:** A controlled addition of measured volumes of a solution of known concentration to a second solution of unknown concentration under conditions in which the solutes react cleanly (without side reactions), completely, and rapidly. A titration is complete when the second solute is used up.

Completion is signaled by a change in some physical property, such as the color of the reacting mixture or the color of an **indicator** that has been added to it.
[\text{NaOH}]

"X"

Indicator
phenolphthalein
Exercise 4-6

Compute the molarity of a solution of sodium hydroxide if 25.64 mL of solution must be added to a solution containing 0.5333 g of KHC$_8$H$_4$O$_4$ (potassium hydrogen phthalate abbreviated KHP) to reach the phenolphthalein end-point.
Exercise 4-4

Give the name for the acid $\text{H}_2\text{SeO}_3$. Write the overall, ionic, and net ionic equations for the complete neutralization of $\text{H}_2\text{SeO}_3$ by sodium hydroxide.

Start with something we know:

Selenium (Se) is in Group VI with Sulfur (S)

$\text{H}_2\text{SO}_4$ is Sulfuric Acid and $\text{H}_2\text{SO}_3$ is Sulfurous Acid
Acids and Bases and Their Reactions

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   • Acids are H⁺ donors
   • Bases are OH⁻ donors

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   • Bases increase OH⁻ concentration or \([OH^-]\) increases

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Oxidation-Reduction or Redox Reactions

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  – Oxidation and Reduction
  – Oxidation Numbers

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Oxidation-Reduction Reactions

**Redox**: an extensive and important class of reactions that is characterized by the transfer of electrons.

\[ 2 \text{Mg}(s) + \text{O}_2(g) \rightarrow 2 \text{MgO}(s) \]

Magnesium is **oxidized**: it *gives up* electrons as the charge on its atoms increases from zero to +2.

Oxygen is **reduced**: it *gains* electrons as the charge on its atoms decreases from zero to -2 (i.e., becomes more negative).
Oxidation Numbers (also called oxidation states) are determined for the atoms in covalently bonded compounds by applying the following set of simple rules:

1. The oxidation number of the atoms in a neutral molecule must all up to zero; those in an ion must add up to the charge on the ion.

2. Alkali metal (Group I) atoms have oxidation number +1, and alkaline earth (Group II) atoms have oxidation number +2 in their compounds; atoms of Group III elements usually have oxidation number +3 in their compounds.

3. Fluorine always has an oxidation number of -1 in its compounds. The other halogens have oxidation number -1 in their compounds, except in compounds with oxygen and with other halogens, in which they can have positive oxidation numbers.
Oxidation Numbers (also called oxidation states) are determined for the atoms in covalently bonded compounds by applying the following set of simple rules:

4. Hydrogen is assigned an oxidation number of +1 in its compounds, except in metal hydrides such as LiH, in which rule 2 take precedence and hydrogen has an oxidation number of -1.

5. Oxygen is assigned an oxidation number of -2 in compounds. There are two exceptions: in compounds with fluorine, rule 3 takes precedence, and in compounds that contain O—O bonds, rules 2 and 4 take precedence. Thus, the oxidation number of oxygen in OF₂ is +2; in peroxides (e.g., H₂O₂ and Na₂O₂), its oxidation number is -1, and in superoxides (e.g., KO₂), its oxidation number is -½.
What is the Nitrogen Atom Oxidation State?

\[
\begin{align*}
\text{NO}_2^- & \quad \text{HONO}_2 \\
\text{NO}_2 & \\
\text{NO}_2^+ & \\
\text{NO}_3^- & 
\end{align*}
\]

Hint: consider HO as HO⁻
Assign Oxidation Numbers to each of the following

Ethane $\text{C}_2\text{H}_6$

Acetic Acid $\text{C}_2\text{H}_4\text{O}_2$

Ammonium Nitrate $\text{NH}_4\text{NO}_3$  Hint: $\text{NH}_4^{+1} \quad \text{NO}_3^{-1}$

Potassium superoxide $\text{KO}_2$
We can now state the following definition:

An atom is oxidized (loses electrons) if its oxidation number increases in a chemical reaction;

an atom is reduced (gains electrons) if its oxidation number decreases.
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<tr>
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<th>Electron Change</th>
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<tr>
<td>Oxidation</td>
<td>Increase</td>
<td>Loss of Electrons</td>
</tr>
<tr>
<td>Reduction</td>
<td>Decrease</td>
<td>Gain of Electrons</td>
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<td>Oxidizing Agent, does the oxidizing</td>
<td>Decrease</td>
<td>Picks Up electrons</td>
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<tr>
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<td>Increase</td>
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<tr>
<td>Substance Oxidized</td>
<td>Increase</td>
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<td>H₂O</td>
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<td>H⁺</td>
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<tr>
<td>Mg</td>
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Types of Redox Reactions

*Redox Combination and Decomposition Reactions*

\[ \text{P}_4(s) + 6 \text{Cl}_2(g) \rightarrow 4 \text{PCl}_3(l) \]

\[ \text{P}_4(s) + 10 \text{Cl}_2(g) \rightarrow 4 \text{PCl}_5(l) \]

\[ 2 \text{HgO (s)} \rightarrow 2 \text{Hg(l)} + \text{O}_2(g) \]

\[ 2 \text{Hg}_2\text{O(s)} \rightarrow 4 \text{Hg(l)} + \text{O}_2(g) \]

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

*Oxygenation and Hydrogenation*

\[ 4 \text{Li (s)} + \text{O}_2(\text{g}) \rightarrow 2 \text{Li}_2\text{O(s)} \]

\[ 2 \text{ZnS(s)} + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{ZnO(s)} + 2 \text{SO}_2(\text{g}) \]

\[ 2 \text{Na(l)} + \text{H}_2(\text{g}) \rightarrow 2 \text{NaH(s)} \]

\[ \text{Fe}_2\text{O}_3(s) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{Fe(s)} + 3 \text{H}_2\text{O(g)} \]
Types of Redox Reactions

Displacement Reactions

\[ 2 \text{AgNO}_3(aq) + \text{Cu}(s) \rightarrow \text{Cu(NO}_3)_2(aq) + 2 \text{Ag}(s) \]  
(overall equation)

\[ 2 \text{Ag}^+(aq) + \text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2 \text{Ag}(s) \]  
(net ionic equation)

Referring to the information (page 178) of Ability to Replace

\[ \text{Cl}_2(g) + 2 \text{KI}(aq) \rightarrow \text{I}_2(s) + 2 \text{KCl}(aq) \]  
(overall equation)

\[ \text{Cl}_2(g) + 2 \text{I}^-(aq) \rightarrow \text{I}_2(s) + 2 \text{Cl}^-(aq) \]  
(net ionic equation)
Types of Redox Reactions

*Disproportionation*

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

\[ 2 \text{Cu}^+(aq) \rightarrow \text{Cu}^{2+}(aq) + \text{Cu}(s) \]  \hspace{1cm} \text{(net ionic equation)}
1. **Dissolution**
   - Ionic Compounds in Water
   - Molecular Compounds in Water
   - Electrolytes and Non-Electrolytes
   - Solubilities (like dissolves like)
   - Predicting Dissolution Reactions

2. **Precipitation Reactions**
   - Ionic Equations and Net Ionic Equations
   - Predicting Precipitation Reactions

3. **Acids and Bases**
   - Arrhenius Acids and Bases Theory
   - Strong and Weak acid
   - Naming Acids
   - Weak Bases
   - Modifying the Arrhenius model
   - Acid-Base Titrations

4. **Oxidation-Reduction Reactions (Redox) Reactions**
   - Oxidizing and Reducing Agents
   - Oxidation Number
   - Types of Redox Reactions
     » Combination and Decomposition
     » Oxygenation
     » Hydrogenation
     » Displacement Reactions
     » Disproportionation
Chapter 4: Types of Chemical Reactions

Examples and Exercises:

All (4-1 to 4-12)

Homework:

1, 6, 12, 50, 54, 55, 59, 71

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