CHAPTER 4
REACTIONS IN AQUEOUS SOLUTION

Chapter Learning Goals

Section 4.1 Classify reactions as precipitation, acid-base neutralization, or redox.

Section 4.2 Classify substances as strong electrolytes, weak electrolytes, or nonelectrolytes.

Section 4.3 Write molecular, ionic, and net ionic equations for precipitation, acid-base, and redox reactions.

Section 4.4 Use solubility guidelines to determine whether a compound is likely to be soluble in water.

Use solubility guidelines to predict whether a precipitate might form when aqueous salt solutions are mixed.

Use solubility guidelines to predict the product of a precipitation reaction.

Show how a precipitation reaction can be used to prepare a substance.

Section 4.5 Identify the common strong acids and strong bases.

Section 4.6 Assign oxidation numbers to each atom in a chemical species.

Section 4.7 In a redox reaction, identify the species oxidized, the species reduced, the oxidizing agent, and the reducing agent.

Section 4.8 Use the activity series to predict whether a redox reaction will occur when a metal is placed in contact with a solution containing an ion of a different metal.

Use the activity series to predict the products of a redox reaction.

Section 4.9 Balance redox reactions by the oxidation-number method in acid or base.

Section 4.10 Balance redox reactions by the half-reaction method in acid or base.

Section 4.11 Determine the concentration of a soution using data from a redox titration.

Section 4.12 Identify the oxidation and reduction half-reactions in 1) combustion, 2) bleaching, 3) battery, 4) metallurgical, 5) corrosion, and 6) respiration reactions.

Lecture Outline

4.1. Some Ways that Chemical Reactions Occur

A. Reactions – reactants transformed to products by some energetic driving force

B. Driving forces of reactions

1. Precipitation Reactions – solid precipitate forms and drops out of solution
   a. Driving force – removal of material from solution
2. Acid-Base Neutralization – acid reacts with base to produce salt and water
   b. Driving force – formation of water
3. Oxidation–Reduction (Redox) Reaction – transfer of electrons between reactants
   a. Driving force – decrease in electrical potential

4.2. Electrolytes in Aqueous Solution

A. Electrolytes – dissolve in water to produce ionic solutions

1. Strong Electrolytes – completely dissociate in water
2. Weak Electrolytes – incompletely dissociate in water; establish dynamic equilibrium between the forward and backward reactions

B. Dissociation – fragmentation of a neutral species into ions

C. Nonelectrolytes – do not produce ions in aqueous solution

D. Chemical properties determine solubility in water

4.3. Aqueous Reactions and Net Ionic Equations

A. Molecular Equations – reactant and product formulas written as if they were molecules. Shows complete formulas of all reactants and products.

B. Ionic Equations – any reactants or products that completely dissociate in water (strong electrolytes) shown as free ions. Shows only reactants and products which dissociate into ions.

C. Spectator Ions – ions not undergoing change during reaction

1. Spectator ion identity not important
2. Role is to balance charge
3. Must be accounted for in determining molar mass of compound dissolved

D. Net Ionic Equation – shows only ions and molecules that participate in reaction

1. Equation for net change taking place during reaction
2. Spectator ions not included

4.4. Precipitation Reactions and Solubility Guidelines

A. Solubility – amount of compound dissolved in a given amount of solvent at a given temperature

1. Low solubility – precipitate forms
2. High solubility – no precipitate will form
3. Solubilities predicted using guidelines found in textbook
4. Compounds considered insoluble if less than $1x10^{-3}$ moles of compound dissolves in 1 liter of water

B. Solubility Guidelines

1. Use to predict if a precipitate will form in a reaction
2. Use to prepare and isolate a specific compound by carrying out a precipitation reaction

4.5. Acids, Bases, and Neutralization Reactions

A. Arrhenius Acid – dissociates in water to produce H^+. Not all compounds containing H are acids.

1. Hydronium Ion
   a. H^+ attached to a water molecule
   b. Represented as H_3O^+
2. Strong acids – strong electrolytes
3. Weak acids – weak electrolytes
4. Polyprotic acids
   a. Acids with more than one acidic hydrogen
   b. Dissociate in steps

B. Arrhenius Base – dissociates in water to produce OH^-. Not all compounds containing OH are bases.
1. Strong bases – strong electrolytes; most metal hydroxides
2. Weak bases – weak electrolytes
$NH_3(g) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$

C. Neutralization Reaction – reaction between acid and base producing salt and water
1. Salt produced from cation of base and anion of acid
2. Net reaction for a strong acid + strong base: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
3. Net reaction for a weak acid + strong base: $HA(aq) + OH^-(aq) \rightarrow H_2O(l) + A^-(aq)$

D. Common acid and bases – textbook Table 4.2

4.6. Oxidation-Reduction (Redox) Reactions

A. Oxidation and Reduction
1. Oxidation
a. Loss of one or more electrons
b. Increase in oxidation number
2. Reduction
a. Gain of one or more electrons
b. Decrease in oxidation number
3. Two processes occur simultaneously

B. Acronyms
1. **LEO** the lion goes **GER** (**L**ose **E**lectrons **O**xidation, **G**ain **E**lectrons **R**eduction)
2. **OIL RIG** (**O**xidation **I**nvolves **L**oss of electrons, **R**eduction **I**nvolves **G**ain of electrons)

C. Redox Reaction
1. Electrons transferred from one substance to another.
2. Number of electrons lost by substance being oxidized = number of electrons gained by substance being reduced

D. Oxidation Number
1. Means of determining whether an atom is neutral, electron-rich, or electron-poor
2. Oxidation numbers of atoms in a compound are the same as the charges the atoms would have if the compound were ionic.
3. Does not necessarily imply ionic charges
4. For any main group element, maximum oxidation number = group number; minimum oxidation number - group number – 8.
5. Rules for assigning oxidation number (textbook pages 127 and 128)

4.7. Identifying Redox Reactions

A. Oxidation and reduction – occur together
1. One atom loses one or more electrons (is oxidized)
2. Another atom must gain those electrons (be reduced)

B. Reducing Agent – substance that causes reduction to occur
1. Loses one or more electrons
a. Undergoes oxidation
b. Oxidation number of atom increases
2. Metals act as reducing agents

C. Oxidizing Agent – the substance that causes oxidation to occur
1. Gains one or more electrons
a. Undergoes reduction
b. Oxidation number of atom decreases
2. Reactive nonmetals act as oxidizing agents

D. Species to the left and bottom of the periodic table tend to lose electrons and are good reducing agents; species at the top and right of the periodic table tend to gain electrons and are good oxidizing agents.

4.8. The Activity Series of the Elements

A. Simple redox process
1. Reaction of an aqueous cation with a free element to produce a different ion and a different element

2. Process dependent on ease of oxidation of the element and ease of reduction of the cation

B. Activity Series – ranks elements in order of reducing ability in aqueous solution (textbook Table 4.3)

1. Any element higher in the activity series will react with the ion of any element lower in the activity series
2. Position of hydrogen – indicates which metals will react with $H^+(aq)$ to produce $H_2(g)$
3. Most reactive metals – top of the activity series
4. Least reactive metals – bottom of the activity series

4.9. Balancing Redox Reactions: The Oxidation-Number Method

A. Oxidation-number method for balancing redox reactions – focuses on chemical changes

B. Key – net change in the total of all oxidation numbers = zero

C. Increase in oxidation number for oxidized atoms matched by a corresponding decrease in oxidation number for reduced atoms

D. Steps – textbook page 136

4.10. Balancing Redox Reactions by the Half-Reaction Method

A. Half-reaction method of balancing redox reactions – focuses on transfer of electrons

B. Key – overall reaction broken into two parts (half-reactions).

1. Oxidation part of the reaction – half-reaction is an oxidation when electrons appear on the right side as products.
2. Reduction part of the reaction – half-reaction is a reduction when electrons appear on the left side as reactants.

C. The number of electrons used to balance the half-reaction should agree with the change in oxidation number of the element undergoing oxidation or reduction.

D. Steps – textbook page 140

4.11. Redox Titrations

A. Redox titration – method for determining concentration of an oxidizing or reducing agent in solution

1. Unknown must react in 100% yield
2. Color change should signal the end of reaction

B. Strategy for redox titrations – textbook Figure 4.5: titration of $KMnO_4$ with $H_2C_2O_4$

1. Measure a known amount of one substance
2. Using the mole ratio of the balanced equation, determine number of moles of a second substrate
3. Volume of solution containing the molar amount of a second substance measured during titration
4. Determine molar concentration of the second substance

4.12. Some Applications of Redox Reactions

A. Redox reactions

1. Involve almost every element in the periodic table
2. Occur in a vast number of processes

B. Examples of redox reactions

1. Combustion
2. Bleaching
3. Batteries
4. Metallurgy
5. Corrosion
6. Respiration