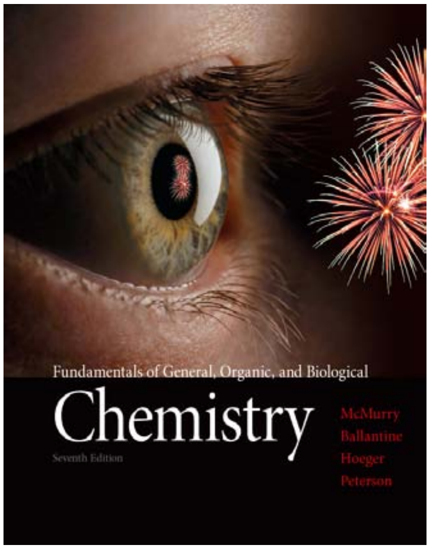
Chapter 5 Lecture



Fundamentals of General, Organic, and Biological Chemistry

7th Edition

McMurry, Ballantine, Hoeger, Peterson

Chapter Five

Classification and Balancing of Chemical Reactions

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Outline

5.1 Chemical Equations

5.2 Balancing Chemical Equations

5.3 Classes of Chemical Reactions

5.4 Precipitation Reactions and Solubility Guidelines

5.5 Acids, Bases, and Neutralization Reactions

5.6 Redox Reactions

5.7 Recognizing Redox Reactions

5.8 Net Ionic Equations

Goals

1. **How are chemical reactions written?**

   Given the identities of reactants and products, be able to write a balanced chemical equation or net ionic equation.

2. **How are chemical reactions of ionic compounds classified?**

   Be able to recognize precipitation, acid–base neutralization, and redox reactions.

3. **What are oxidation numbers, and how are they used?**

   Be able to assign oxidation numbers to atoms in compounds and identify the substances oxidized and reduced in a given reaction.

4. **What is a net ionic equation?**

   Be able to recognize spectator ions and write the net ionic equation for reactions involving ionic compounds.

5.1 Chemical Equations

- Chemical reactions can be thought of as "recipes."
- Starting materials and final products are listed, with chemical names replaced with formulas.
- A **chemical equation** is an expression in which symbols and formulas are used to represent a chemical reaction.

5.1 Chemical Equations

$$\underbrace{2\,NaHCO_3}_{\text{Reactant}} \xrightarrow{\text{Heat}} \underbrace{Na_2CO_3 + H_2O + CO_2}_{\text{Products}}$$

- A **reactant** is a substance that undergoes change in a chemical reaction and is written on the left side of the reaction arrow in a chemical equation.
- A **product** is a substance that is formed in a chemical reaction and is written on the right side of the reaction arrow in a chemical equation.
- Necessary conditions are written above the arrow.

5.1 Chemical Equations

- The **law of conservation of mass** states that matter can neither be created nor destroyed in a chemical reaction.
- Bonds between atoms in the reactants are rearranged to form new compounds, but none of the atoms disappear and no new ones are formed.
- Therefore, chemical equations must be **balanced**—the numbers and kinds of atoms must be the same on both sides of the reaction arrow.

5.1 Chemical Equations

- A **coefficient** is a number placed in front of a formula to balance a chemical equation.
- Coefficients multiply all the atoms in a formula.
- Be sure to consider coefficients.
- Substances that take part in chemical reactions may be solids, liquids, gases, or dissolved. This information is added to an equation by placing symbols after the formulas.
    - (*s*) = solid
    - (*l*) = liquid
    - (*g*) = gas
    - (*aq*) = dissolved in aqueous solution

5.2 Balancing Chemical Equations

STEP 1: Write an unbalanced equation, using the correct formulas for all reactants and products.

- For example, hydrogen and oxygen must be written as H_2 and O_2 rather than as H and O, since both elements exist as diatomic molecules.
- Subscripts in chemical formulas cannot be changed because doing so would change the identity of the substances in the reaction.

5.2 Balancing Chemical Equations

STEP 2: Add appropriate coefficients to balance the numbers of atoms of each element.

- Begin with elements that appear in only one compound or formula on each side of the equation.
- Leave elements that exist in elemental forms, such as oxygen and hydrogen, until last.

$$H_2SO_4 + NaOH \longrightarrow Na_2SO_4 + H_2O \quad \text{(Unbalanced)}$$

$$H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + H_2O \quad \text{(Balanced for Na)}$$

Add this coefficient . . . . . . to balance these 2 Na.

5.2 Balancing Chemical Equations

- If a polyatomic ion appears on both sides of an equation, it is treated as a single unit.

$$H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + H_2O \quad \text{(Balanced for Na and sulfate)}$$

One sulfate here . . . . . . and one here.

$$H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + 2\,H_2O \quad \text{(Completely balanced)}$$

4 H and 2 O here. 4 H and 2 O here.

5.2 Balancing Chemical Equations

STEP 3: Check the equation to make sure the numbers and kinds of atoms on both sides of the equation are the same.

STEP 4: Make sure the coefficients are reduced to their lowest whole-number values.

5.3 Classes of Chemical Reactions

- **Precipitation reactions** are processes in which an insoluble solid called a **precipitate** forms when reactants are combined in aqueous solution.
- Most precipitations take place when the anions and cations of two ionic compounds change partners.

5.3 Classes of Chemical Reactions

- **Acid–base neutralization reactions** are processes in which an acid reacts with a base to yield water plus an ionic compound called a **salt**.
- A neutralization reaction removes H^+ and OH^- ions from solution and yields neutral H_2O.
- *Any* ionic compound produced in an acid–base reaction is called a salt.

5.3 Classes of Chemical Reactions

- **Oxidation–reduction reactions** or **redox reactions**, are processes in which one or more electrons are transferred.
- As a result of this transfer, the number of electrons assigned to individual atoms in the various reactants change.
- Reactions involving covalent compounds are redox reactions since electrons are rearranged as bonds are broken and new bonds are formed.

5.4 Precipitation Reactions and Solubility Guidelines

- Whether a precipitation reaction will occur on mixing aqueous solutions of two ionic compounds, depends on the **solubilities** of the potential products.
- If a substance has low solubility then it is likely to precipitate from an aqueous solution. If a substance has high solubility in water, then no precipitate will form.
- **Solubility** is the amount of a compound that will dissolve in a given amount of solvent at a given temperature.

5.4 Precipitation Reactions and Solubility Guidelines

General Rules on Solubility

1. A compound is probably soluble if it contains one of the following cations.

- Group 1A cation: Li^+, Na^+, K^+, Rb^+, Cs^+
- Ammonium cation: NH_4^+

2. A compound is probably soluble if it contains one of the following anions.

- Halide: Cl^-, Br^-, I^-, except *Ag^+, Hg_2^{2+}*, and *Pb^{2+}* compounds
- Nitrate (NO_3^-), perchlorate (ClO_4^-), acetate ($CH_3CO_2^-$), sulfate (SO_4^{2-}), except *Ba^{2+}, Hg_2^{2+}*, and *Pb^{2+}* sulfates

5.4 Precipitation Reactions and Solubility Guidelines

TABLE 5.1 General Solubility Guidelines for Ionic Compounds in Water

Soluble	Exceptions
Ammonium compounds (NH_4^+)	None
Lithium compounds (Li^+)	None
Sodium compounds (Na^+)	None
Potassium compounds (K^+)	None
Nitrates (NO_3^-)	None
Perchlorates (ClO_4^-)	None
Acetates ($CH_3CO_2^-$)	None
Chlorides (Cl^-)	Ag^+, Hg_2^{2+}, and Pb^{2+} compounds
Bromides (Br^-)	Ag^+, Hg_2^{2+}, and Pb^{2+} compounds
Iodides (I^-)	Ag^+, Hg_2^{2+}, and Pb^{2+} compounds
Sulfates (SO_4^{2-})	Ba^{2+}, Hg_2^{2+}, and Pb^{2+} compounds

5.4 Precipitation Reactions and Solubility Guidelines

Gout and Kidney Stones: Problems in Solubility

- One of the major pathways in the body for the breakdown of nucleic acids is by conversion to a substance called *uric acid.*
- When too much sodium urate is produced or mechanisms for its elimination fail, its concentration in blood and urine rises.
- *Gout* is a disorder of nucleic acid metabolism. It is characterized by increased sodium urate in blood, leading to the deposit of sodium urate crystals in soft tissue around the joints. Deposits of the sharp, needlelike crystals cause an extremely painful inflammation that can lead ultimately to arthritis and even to bone destruction.
- Increased sodium urate concentration in urine can result in the formation of *kidney stones*, small crystals that precipitate in the kidney. Although often quite small, kidney stones cause excruciating pain when they pass through the ureter, the duct that carries urine from the kidney to the bladder.

5.5 Acids, Bases, and Neutralization Reactions

- The most common kind of neutralization reaction occurs between an acid (HA), and a metal hydroxide (MOH), to yield water and a salt.
- The H^+ ion from the acid combines with the OH^- ion from the base to give neutral H_2O.
- The anion from the acid combines with the cation from the base to give a salt.

$$HA\ (aq) + MOH\ (aq) \rightarrow H_2O\ (l) + MH\ (aq)$$

5.5 Acids, Bases, and Neutralization Reactions

- Another kind of neutralization reaction occurs between an acid and a carbonate or bicarbonate to yield water, a salt, and carbon dioxide.
- The carbonate ion reacts with H^+ to yield H_2CO_3.
- This is unstable and immediately decomposes to give H_2O plus CO_2.

$$2\ HA\ (aq) + M_2CO_3\ (aq) \rightarrow H_2O\ (l) + 2\ MH\ (aq) + CO_2\ (g)$$

5.5 Acids, Bases, and Neutralization Reactions

- Carbonates yield OH^- ions when dissolved in water.

M_2CO_3 (*s*) H_2O (*l*) → 2 M^+ (*aq*) + HCO_3^- (*aq*) + OH^- (*aq*)

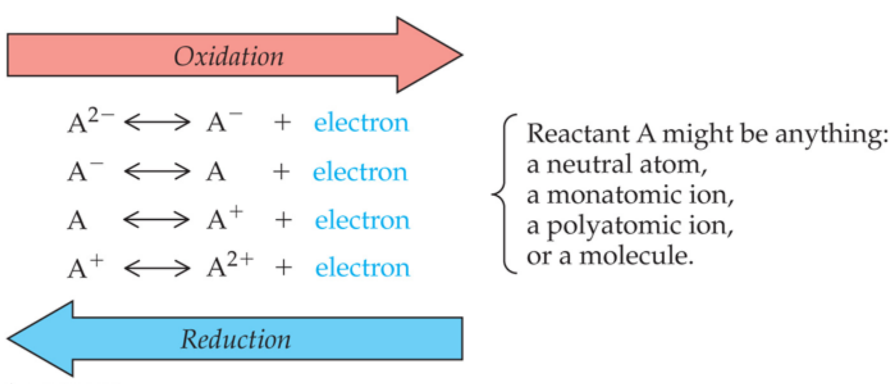
5.6 Redox Reactions

- This category of reactions is more complex.
- Historically, the word *oxidation* referred to the combination of an element with oxygen.
- *Reduction* referred to the removal of oxygen from an oxide to yield the element.

5.6 Redox Reactions

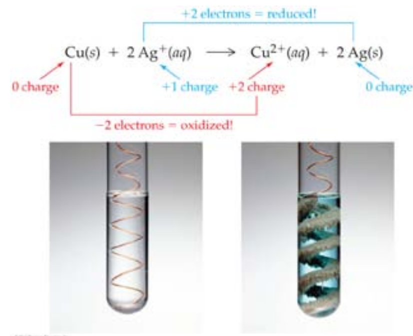
- Today the words have taken on a much broader meaning.
  - **Oxidation** is the loss of one or more electrons.
  - **Reduction** is the gain of one or more electrons.
- An oxidation–reduction reaction, or redox reaction, is one in which *electrons are transferred from one atom to another*.

5.6 Redox Reactions

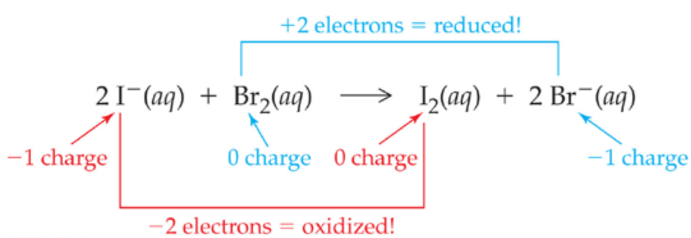


5.6 Redox Reactions

- Copper metal gives an electron to each of two silver ions, forming copper ions and silver metal.
- Copper is oxidized, and silver ions are reduced.
  - The charge on the copper increases from 0 to +2, when it loses two electrons.
  - The charge on the silver decreases from +1 to 0, when it gains an electron.



5.6 Redox Reactions



- An iodine ion gives an electron to bromine, forming iodine and bromide ions.
- An iodide ion is oxidized as its charge increases from −1 to 0.
- Bromine is reduced as its charge decreases from 0 to −1.

5.6 Redox Reactions

- Oxidation and reduction always occur together.
- When one substance loses an electron (is oxidized), another substance must gain that electron (be reduced).
- The substance that gives up an electron and causes the reduction is a **reducing agent**.
- The substance that gains an electron and causes the oxidation is an **oxidizing agent**.
  - The charge on the reducing agent increases.
  - The charge on the oxidizing agent decreases.

5.6 Redox Reactions

- A **reducing agent** loses one or more electrons.
  - Causes reduction
  - Undergoes oxidation
  - Becomes more positive (or less negative)
  - May gain oxygen atoms
- An **oxidizing agent** gains one or more electrons.
  - Causes oxidation
  - Undergoes reduction
  - Becomes more negative (or less positive)
  - May lose oxygen atoms

5.6 Redox Reactions

- The reaction of metal with water or aqueous acid is a particularly important process.
- Alkali metals and alkaline earth metals are the most powerful reducing agents.
  - They will react with pure water.
  - They have low ionization energy.
  - As ionization energy increases, reducing power decreases.
- Reactive nonmetals are powerful oxidizing agents.
  - They have the highest oxidation energies.
  - They have the most favorable electron affinity.

5.6 Redox Reactions

- Generalizations about redox behavior:
  1. In reactions involving metals and nonmetals, metals tend to lose electrons while nonmetals tend to gain electrons. The number of electrons lost or gained can often be predicted based on the position of the element in the periodic table.
  2. In reactions involving nonmetals, the "more metallic" element (farther down and/or to the left in the periodic table) tends to lose electrons, and the "less metallic" element (up and/or to the right) tends to gain electrons.

5.6 Redox Reactions

- **Corrosion** is the deterioration of a metal by oxidation, such as the rusting of iron in moist air.
  - The economic consequences of rusting are enormous. It has been estimated that up to one-fourth of the iron produced in the United States is used to replace bridges, buildings, and other structures that have been destroyed by corrosion.

5.6 Redox Reactions

- **Combustion** is the burning of a fuel by rapid oxidation with oxygen in air.
  - Gasoline, fuel oil, natural gas, wood, paper, and other organic substances of carbon and hydrogen are the most common fuels that burn in air.
  - Some metals will burn in air, including magnesium and calcium.

5.6 Redox Reactions

- **Respiration** is the process of breathing and using oxygen for the many biological redox reactions that provide the energy that living organisms need.
  - Energy is released from food molecules slowly and in complex, multistep pathways, but the overall result of respiration is similar to that of combustion reactions.

5.6 Redox Reactions

- **Bleaching** makes use of redox reactions to decolorize or lighten colored materials.
  - Dark hair is bleached to turn it blond, clothes are bleached to remove stains, wood pulp is bleached to make white paper.
  - The oxidizing agent used depends on the situation. Hydrogen peroxide is used for hair, sodium hypochlorite ($NaOCl$) for clothes, and elemental chlorine for wood pulp.
  - In all cases, colored organic materials are destroyed by reaction with strong oxidizing agents.

5.6 Redox Reactions

- **Metallurgy,** the science of extracting and purifying metals from their ores, makes use of numerous redox processes.
- Worldwide, approximately 800 million tons of iron is produced each year by reduction of the mineral hematite with carbon monoxide.

$$Fe_2O_3\ (s) + 3\ CO\ (g) \rightarrow 2\ Fe\ (s) + 3\ CO_2\ (g)$$

5.6 Redox Reactions

Batteries

- Batteries are based on redox reactions. In a battery, the two reactants are kept in separate compartments and the electrons are transferred through a wire running between them.
- The common household battery used for flashlights and radios is the *dry-cell*, developed in 1866. One reactant is a can of zinc metal, and the other is a paste of solid manganese dioxide. A graphite rod provides electrical contact, and a moist paste of ammonium chloride separates the reactants. If the zinc and the graphite are connected, zinc sends electrons flowing through the wire in a redox reaction.
- In *alkaline* batteries, the ammonium chloride paste is replaced by an alkaline, or basic, paste of NaOH or KOH.
- The batteries used in implanted medical devices, such as pacemakers, must be small, corrosion-resistant, reliable, and able to last up to 10 years. Nearly all pacemakers being implanted today—about 750,000 each year—use titanium-encased, lithium-iodine batteries.

5.7 Recognizing Redox Reactions

- When ions are involved, determine whether there is a change in charges.
- For reactions involving metals and nonmetals, predict gain or loss of electrons.
- Molecular substances can be analyzed in terms of loss and gain of oxygen.

OR

- By extending the ideas of oxidation and reduction to an increase or decrease in electron sharing, electronegativity differences can be used.
  - An atom is oxidized when it loses a share in electrons.
  - An atom is reduced when it gains a share in electrons.

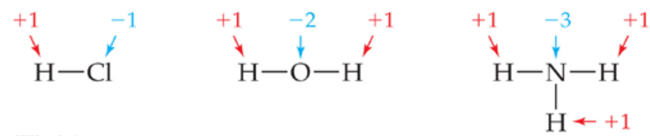
5.7 Recognizing Redox Reactions

- A formal system has been devised for keeping track of changes in electron sharing, and determining whether atoms are oxidized or reduced in reactions.
- A value called an **oxidation number** (or *oxidation state*), indicates whether the atom is neutral, electron-rich, or electron-poor.
- By comparing the oxidation number of an atom before and after a reaction, we can tell whether the atom has gained or lost shares in electrons.
- Oxidation numbers do not necessarily imply ionic charges. They are simply a convenient device for keeping track of electrons in redox reactions.

5.7 Recognizing Redox Reactions

The rules for assigning oxidation numbers are straightforward.

- An atom in its elemental state has an oxidation number of 0.
- A monatomic ion has an oxidation number equal to its charge.
- In a molecular compound, an atom usually has the same oxidation number it would have if it were a monatomic ion.
- The sum of the oxidation numbers in a neutral compound is 0. In a polyatomic ion, the sum of the charges is the charge on the ion.



5.8 Net Ionic Equations

- In reactions involving ions, it is more accurate to write the reaction as an **ionic equation**.
- An **ionic equation** is one in which ions are explicitly shown.
- Some ions undergo no change during the reaction. They appear on both sides of the reaction but play no role.
- A **spectator ion** is an ion that appears unchanged on both sides of a reaction arrow.

5.8 Net Ionic Equations

- The actual reaction can be described more simply by writing a **net ionic equation**, which includes only the ions that undergo change.
- A **net ionic equation** is an equation that does not include spectator ions.
- A net ionic equation must be balanced both for atoms and for charge, with all coefficients reduced to their lowest whole numbers.
- All insoluble and molecular compounds are represented by their full formulas.

5.8 Net Ionic Equations

- Applied to neutralization reactions,

$$KOH\ (aq) + HNO_3\ (aq) \rightarrow H_2O\ (l) + KNO_3\ (aq)$$

can be rewritten as an ionic equation,

$$K^+\ (aq) + OH^-\ (aq) + H^+\ (aq) + NO_3^-\ (aq) \rightarrow H_2O\ (l) + K^+\ (aq) + NO_3^-\ (aq).$$

Eliminating the spectator ions (K^+ and NO_3^-) yields

$$OH^-\ (aq) + H^+\ (aq) \rightarrow H_2O\ (l).$$

- This net ionic equation confirms a neutralization reaction.

5.8 Net Ionic Equations

- Applied to redox reactions,

$$Cu\ (s) + 2\ AgNO_3\ (aq) \rightarrow 2\ Ag\ (s) + Cu(NO_3)_2\ (aq)$$

can be rewritten as an ionic equation,

$$Cu\ (s) + 2\ Ag^+\ (aq) + 2\ NO_3^-\ (aq) \rightarrow$$
$$2\ Ag\ (s) + Cu^{2+}\ (aq) + 2\ NO_3^-\ (aq).$$

Eliminating the spectator ion (NO_3^-) yields

$$Cu\ (s) + 2\ Ag^+\ (aq) \rightarrow 2\ Ag\ (s) + Cu^{2+}\ (aq).$$

- This net ionic equation confirms that copper is oxidized and silver is reduced.

Chapter Summary

1. How are chemical reactions written?

- Chemical equations must be *balanced*; that is, the numbers and kinds of atoms must be the same in both the reactants and the products.
- To balance an equation, *coefficients* are placed before formulas, but the formulas themselves cannot be changed.

Chapter Summary, *Continued*

2. How are chemical reactions of ionic compounds classified?

- *Precipitation* reactions are processes in which an insoluble solid called a *precipitate* is formed. Most precipitations take place when the anions and cations of two ionic compounds change partners.
- *Acid–base neutralization* reactions are processes in which an acid reacts with a base to yield water plus an ionic compound called a *salt*. Since acids produce H^+ ions and bases produce OH^- ions when dissolved in water, a neutralization reaction removes H^+ and OH^- ions from solution and yields water.

Chapter Summary

2. How are chemical reactions of ionic compounds classified? (*Continued)*

- *Oxidation–reduction* (redox) reactions are processes in which one or more electrons are transferred between reaction partners.
  - An *oxidation* is defined as the loss of one or more electrons by an atom.
  - A *reduction* is the gain of one or more electrons.
  - An *oxidizing agent* causes the oxidation of another reactant by accepting electrons.
  - A *reducing agent* causes the reduction of another reactant by donating electrons.

Chapter Summary, *Continued*

3. What are oxidation numbers, and how are they used?

- Oxidation numbers are assigned to atoms in reactants and products to provide a measure of whether an atom is neutral, electron-rich, or electron-poor.
- By comparing the oxidation number of an atom before and after a reaction, we can tell whether the atom has gained or lost shares in electrons and thus, whether a redox reaction has occurred.

Chapter Summary, *Continued*

4. What is a net ionic equation?

- The *net ionic equation* only includes those ions that are directly involved in the ionic reaction.
- These ions are found in different phases or compounds on the reactant and product sides of the chemical equation.
- The net ionic equation does not include *spectator ions*, which appear in the same state on both sides of the equation.