A general chemistry Libretexts Textmap organized around the textbook

**Chemistry: The Central Science**
by Brown, LeMay, Busten, Murphy, and Woodward

1. Chapter 1: Introduction: Matter and Measurement
   2. 1.1: The Study of Chemistry
   3. 1.2: Classification of Matter
   4. 1.3: Properties of Matter
   5. 1.4: Units of Measurement
   6. 1.5: Uncertainty in Measurement
   7. 1.6: Dimensional Analysis
   8. 1.E: Matter and Measurement (Exercises)
   9. 1.S: Matter and Measurement (Summary)

1. Chapter 2: Atoms, Molecules, and Ions
   2. 2.1: The Atomic Theory of Matter
   3. 2.2: The Discovery of Atomic Structure
   4. 2.3: The Modern View of Atomic Structure
   5. 2.4: Atomic Mass
   6. 2.5: The Periodic Table
   7. 2.6: Molecules and Molecular Compounds
   8. 2.7: Ions and Ionic Compounds
   9. 2.8: Naming Inorganic Compounds
   10. 2.9: Some Simple Organic Compounds
   11. 2.E: Atoms, Molecules, and Ions (Exercises)
   12. 2.S: Atoms, Molecules, and Ions (Summary)

1. Chapter 3: Stoichiometry: Chemical Formulas and Equations
   2. 3.1: Chemical Equations
   3. 3.2: Some Simple Patterns of Chemical Reactivity
   4. 3.3: Formula Masses
   5. 3.4: Avogadro's Number and the Mole
   6. 3.5: Empirical Formulas from Analysis
   7. 3.6: Quantitative Information from Balanced Equations
   8. 3.7: Limiting Reactants
   9. 3.E: Stoichiometry (Exercises)
1. Chapter 4: Reactions in Aqueous Solution
2. 4.1: General Properties of Aqueous Solutions
3. 4.2: Precipitation Reactions
4. 4.3: Acid-Base Reactions
5. 4.4: Oxidation-Reduction Reactions
6. 4.5: Concentration of Solutions
7. 4.6: Solution Stoichiometry and Chemical Analysis
8. 4.E: Reactions in Aqueous Solution (Exercises)
9. 4.S: Reactions in Aqueous Solution (Summary)

10. 3.S: Stoichiometry (Summary)

1. Chapter 5: Thermochemistry
2. 5.1: The Nature of Energy
3. 5.2: The First Law of Thermodynamics
4. 5.3: Enthalpy
5. 5.4: Enthalpy of Reaction
6. 5.5: Calorimetry
7. 5.6: Hess’s Law
8. 5.7: Enthalpies of Formation
9. 5.8: Foods and Fuels
10. 5.E: Thermochemistry (Exercises)
11. 5.S: Thermochemistry (Summary)

1. Chapter 6: Electronic Structure of Atoms
2. 6.1: The Wave Nature of Light
3. 6.2: Quantized Energy and Photons
4. 6.3: Line Spectra and the Bohr Model
5. 6.4: The Wave Behavior of Matter
6. 6.5: Quantum Mechanics and Atomic Orbitals
7. 6.6: 3D Representation of Orbitals
8. 6.7: Many-Electron Atoms
9. 6.8: Electron Configurations
10. 6.9: Electron Configurations and the Periodic Table
11. 6.E: Electronic Structure of Atoms (Exercises)
12. 6.S: Electronic Structure of Atoms (Summary)
• 7
1. Chapter 7: Periodic Properties of the Elements
   2. 7.1: Development of the Periodic Table
   3. 7.2: Effective Nuclear Charge
   4. 7.3: Sizes of Atoms and Ions
   5. 7.4: Ionization Energy
   6. 7.5: Electron Affinities
   7. 7.6: Metals, Nonmetals, and Metalloids
   8. 7.7: Group Trends for the Active Metals
   9. 7.8: Group Trends for Selected Nonmetals
10. 7.E: Periodic Properties of the Elements (Exercises)
11. 7.S: Periodic Properties of the Elements (Summary)

• 8
1. Chapter 8: Basic Concepts of Chemical Bonding
   2. 8.1: Chemical Bonds, Lewis Symbols, and the Octet Rule
      3. 8.2: Ionic Bonding
      4. 8.3: Covalent Bonding
      5. 8.4: Bond Polarity and Electronegativity
      6. 8.5: Drawing Lewis Structures
      7. 8.6: Resonance Structures
      8. 8.7: Exceptions to the Octet Rule
      9. 8.8: Strength of Covalent Bonds
10. 8.E: Basic Concepts of Chemical Bonding (Exercises)
11. 8.S: Basic Concepts of Chemical Bonding (Summary)

• 9
1. Chapter 9: Molecular Geometry and Bonding Theories
   2. 9.1: Molecular Shapes
   3. 9.2: The VSEPR Model
   4. 9.3: Molecular Shape and Molecular Polarity
   5. 9.4: Covalent Bonding and Orbital Overlap
      6. 9.5: Hybrid Orbitals
      7. 9.6: Multiple Bonds
      8. 9.7: Molecular Orbitals
   9. 9.8: Second-Row Diatomic Molecules
10. 9.E: Exercises
11. 9.S: Molecular Geometry and Bonding Theories (Summary)
1. Chapter 13: Properties of Solutions
   2. 13.1: The Solution Process
   3. 13.2: Saturated Solutions and Solubility
   4. 13.3: Factors Affecting Solubility
   5. 13.4: Ways of Expressing Concentration
   6. 13.5: Colligative Properties
   7. 13.6: Colloids
   8. 13.E: Properties of Solutions (Exercises)
   9. 13.S: Properties of Solutions (Summary)

14

1. Chapter 14: Chemical Kinetics
   2. 14.1: Factors that Affect Reaction Rates
   3. 14.2: Reaction Rates
   4. 14.3: Concentration and Rates (Differential Rate Laws)
   5. 14.4: The Change of Concentration with Time (Integrated Rate Laws)
   6. 14.5: Temperature and Rate
   7. 14.6: Reaction Mechanisms
   8. 14.7: Catalysis
   9. 14.E: Exercises
   10. 14.S: Chemical Kinetics (Summary)

15

1. Chapter 15: Chemical Equilibrium
   2. 15.1: The Concept of Equilibrium
   3. 15.2: The Equilibrium Constant
   4. 15.3: Interpreting & Working with Equilibrium Constants
   5. 15.4: Heterogeneous Equilibria
   6. 15.5: Calculating Equilibrium Constants
   7. 15.6: Applications of Equilibrium Constants
   8. 15.7: Le Châtelier's Principle
   9. 15.E: Exercises
   10. 15.S: Chemical Equilibrium (Summary)

16

1. Chapter 16: Acid–Base Equilibria
   2. 16.1: Acids and Bases: A Brief Review
   3. 16.2: Brønsted–Lowry Acids and Bases
   4. 16.3: The Autoionization of Water
5. **16.4: The pH Scale**
6. **16.5: Strong Acids and Bases**
7. **16.6: Weak Acids**
8. **16.7: Weak Bases**
9. **16.8: Relationship Between $K_a$ and $K_b$**
10. **16.9: Acid-Base Properties of Salt Solutions**
11. **16.10: Acid-Base Behavior and Chemical Structure**
12. **16.11: Lewis Acids and Bases**

• 17

1. **Chapter 17: Additional Aspects of Aqueous Equilibria**
   2. **17.1: The Common-Ion Effect**
   3. **17.2: Buffered Solutions**
   4. **17.3: Acid-Base Titrations**
   5. **17.4: Solubility Equilibria**
   6. **17.5: Factors that Affect Solubility**
   7. **17.6: Precipitation and Separation of Ions**
   8. **17.7: Qualitative Analysis for Metallic Elements**
10. **17.S: Additional Aspects of Aqueous Equilibria (Summary)**

• 18

1. **Chapter 18: Chemistry of the Environment**
   2. **18.1: Earth's Atmosphere**
   3. **18.2: Outer Regions of the Atmosphere**
   4. **18.3: Ozone in the Upper Atmosphere**
   5. **18.4: Chemistry of the Troposphere**
   6. **18.5: The World Ocean**
   7. **18.6: Fresh Water**
   8. **18.7: Green Chemistry**

• 19

1. **Chapter 19: Chemical Thermodynamics**
   2. **19.1: Spontaneous Processes**
   4. **19.3: The Molecular Interpretation of Entropy**
5. 19.4: Entropy Changes in Chemical Reactions
6. 19.5: Gibbs Free Energy
7. 19.6: Free Energy and Temperature
8. 19.7: Free Energy and the Equilibrium Constant
9. 19.E: Chemical Thermodynamics (Exercises)
   • 20
   1. Chapter 20: Electrochemistry
   2. 20.1: Oxidation States & Redox Reactions
   3. 20.2: Balanced Oxidation-Reduction Equations
   4. 20.3: Voltaic Cells
   5. 20.4: Cell Potential Under Standard Conditions
   6. 20.5: Gibbs Energy and Redox Reactions
   7. 20.6: Cell Potential Under Nonstandard Conditions
   8. 20.7: Batteries and Fuel Cells
   9. 20.8: Corrosion
   10. 20.9: Electrolysis
   11. 20.E: Electrochemistry (Exercises)
   • 21
   1. Chapter 21: Nuclear Chemistry
   2. 21.1: Radioactivity
   3. 21.2: Patterns of Nuclear Stability
   4. 21.3: Nuclear Transmutations
   5. 21.4: Rates of Radioactive Decay
   6. 21.6: Energy Changes in Nuclear Reactions
   7. 21.7: Nuclear Fission
   8. 21.8: Nuclear Fusion
   9. 21.9: Biological Effects of Radiation
   10. 21.E: Exercises
   11. 21.S: Nuclear Chemistry (Summary)
   • 22
   1. Chapter 22: Chemistry of the Nonmetals
   2. 22.1: General Concepts: Periodic Trends and Reactions
   3. 22.2: Hydrogen
   4. 22.3: Group 18: Nobel Gases
   5. 22.4: Group 17: The Halogens
   6. 22.5: Oxygen
7. **22.6: The Other Group 16 Elements: S, Se, Te, and Po**
   8. **22.7: Nitrogen**
9. **22.8: The Other Group 15 Elements: P, As, Sb, and Bi**
   10. **22.9: Carbon**
11. **22.10: The Other Group 14 Elements: Si, Ge, Sn, and Pb**
   12. **22.11: Boron**
13. **22.E: Chemistry of the Nonmetals (Exercises)**
   14. **22.S: Chemistry of the Nonmetals (Summary)**

• 23

1. **Chapter 23: Metals and Metallurgy**
2. **23.1: Occurance and Distribution of Metals**
   3. **23.2: Pyrometallurgy**
   4. **23.3: Hydrometallurgy**
   5. **23.4: Electrometallurgy**
   6. **23.5: Metallic Bonding**
   7. **23.6: Alloys**
   8. **23.7: Transition Metals**
9. **23.8: Chemistry of Selected Transition Metals**

• 24

1. **Chapter 24: Chemistry of Coordination Chemistry**
2. **24.1: Metal Complexes**
3. **24.2: Ligands with more than one Donor Atom**
4. **24.3: Nomenclature of Coordination Chemistry**
   5. **24.4: Isomerization**
   6. **24.5: Color and Magnetism**
   7. **24.6: Crystal Field Theory**
8. **24.E: Chemistry of Coordination Chemistry (Exercises)**

• 25

1. **Chapter 25: Chemistry of Life: Organic and Biological Chemistry**
2. **25.1: General Characteristics of Organic Molecules**
   3. **25.2: Introduction to Hydrocarbons**
   4. **25.3: Alkanes**
   5. **25.4: Unsaturated Hydrocarbons**
   6. **25.5: Functional Groups**
   7. **25.6: Compounds with a Carbonyl Group**
8. **25.7: Chirality in Organic Chemistry**
9. **25.8: Introduction to Biochemistry**
10. **25.9: Proteins**
11. **25.10: Carbohydrates**
12. **25.11: Nucleic Acids**
13. **25.E: Organic and Biological Chemistry (Exercises)**
14. **25.S: Organic and Biological Chemistry (Summary)**

**Homework**

1. **1.E: Matter and Measurement (Exercises)**
2. **2.E: Atoms, Molecules, and Ions (Exercises)**
3. **3.E: Stoichiometry (Exercises)**
4. **4.E: Aqueous Reactions (Exercises)**
5. **5.E: Thermochemistry (Exercises)**
8. **8.E: Chemical Bonding Basics (Exercises)**
10. **10.E: Gases (Exercises)**
11. **11.E: Liquids and Intermolecular Forces (Exercises)**
15. **15.E: Chemical Equilibrium (Exercises)**
17. **17.E: Additional Aspects of Aqueous Equilibria (Exercises)**
18. **18.E: Chemistry of the Environment (Exercises)**
19. **19.E: Chemical Thermodynamics (Exercises)**
22. **22.E: Chemistry of the Nonmetals (Exercises)**
23. **23.E: Metals and Metallurgy (Exercises)**
24. **24.E: Chemistry of Coordination Chemistry (Exercises)**
25. **25.E: Organic and Biological Chemistry (Exercises)**

These is a summary of key concepts of the chapter in the Textmap created for "Chemistry: The Central Science" by Brown et al.
14.1: Factors that Affect Reaction Rates

**chemical kinetics** – area of chemistry dealing with speeds/rates of reactions

- rates of reactions affected by four factors
  1. concentrations of reactants
  2. temperature at which reaction occurs
  3. presence of a catalyst
  4. surface area of solid or liquid reactants and/or catalysts

14.2: Reaction Rates

- **reaction rate** – speed of a chemical reaction

\[
\text{average rate} = \frac{\text{change in moles B}}{\text{change in time}} = \frac{\Delta \text{moles B}}{\Delta t} \text{ if } A \to B
\]

\[
\Delta \text{moles B} = \text{moles B at final time} - \text{moles B at initial time}
\]

\[
\text{average rate} = -\frac{\Delta \text{moles A}}{\Delta t} \text{ if } A \to B
\]

14.2.1 Rates in Terms of Concentrations

- rate calculated in units of M/s
- brackets around a substance indicate the concentration
- **instantaneous rate** – rate at a particular time
- instantaneous rate obtained from the straight line tangent that touches the curve at a specific point
- slopes give instantaneous rates
- instantaneous rate also referred to as the rate

14.2.2 Reaction Rates and Stoichiometry

- for the irreversible reaction \((aA+bB\to cC+dD)\)

\[
\text{rate} = -\frac{1}{a}\frac{\Delta [A]}{\Delta t} = -\frac{1}{b}\frac{\Delta [B]}{\Delta t} = \frac{1}{c}\frac{\Delta [C]}{\Delta t} = \frac{1}{d}\frac{\Delta [D]}{\Delta t}
\]

14.3: Concentration and Rate

- equation used only if C and D only substances formed
- Rate = \(k[A][B]\)
- **Rate law** – expression that shows that rate depends on concentrations of reactants
- \(k = \text{rate constant}\)
14.3.1 Reaction Order

- Rate = \( k[\text{reactant 1}]^m[\text{reactant 2}]^n \)
- \( m, n \) are called reaction orders
- \( m+n \), overall reaction order
- reaction orders do not have to correspond with coefficients in balanced equation
- values of reaction order determined experimentally
- reaction order can be fractional or negative

14.3.2 Units of Rates Constants

- units of rate constant depend on overall reaction order of rate law
- for reaction of second order overall
- units of rate = (units of rate constant)(units of concentration)\(^2\)
- units of rate constant = \( \text{M}^{-1}\text{s}^{-1} \)

14.3.3 Using Initial Rates to Determine Rate Laws

- zero order – no change in rate when concentration changed
- first order – change in concentration gives proportional changes in rate
- second order – change in concentration changes rate by the square of the concentration change, such as \( 2^2 \) or \( 3^2 \), etc...
- rate constant does not depend on concentration

14.4: The Change of Concentration with Time

- rate laws can be converted into equations that give concentrations of reactants or products

14.4.1 First-Order Reactions

\[
\text{rate} = -\frac{\Delta [A]}{\Delta t} = k[A]
\]

and in integral form:

\[
\ln[A]_t - \ln[A]_0 = -kt
\]

or

\[
\ln\left[\frac{[A]_t}{[A]_0}\right] = -kt
\]

\[
\ln[A]_t = -kt + \ln[A]_0
\]

- corresponds to a straight line with \( y = mx + b \)
- equations used to determine:
1. concentration of reactant remaining at any time
2. time required for given fraction of sample to react
3. time required for reactant concentration to reach a certain level

14.3.2 Half-Life

- half-life of first order reaction
  \[
  t_{\frac{1}{2}} = -\frac{\ln\frac{1}{2}}{k} = \frac{0.693}{k}
  \]

- **half-life** – time required for concentration of reactant to drop to one-half of initial value
- \(t_{\frac{1}{2}}\) of first order independent of initial concentrations
- half-life same at any given time of reaction
- in first order reaction – concentrations of reactant decreases by ½ in each series of regularly spaced time intervals

14.3.3 Second-Order Reactions

- rate depends on reactant concentration raised to second power or concentrations of two different reactants each raised to first power

  \[
  \text{Rate} = k[A]^2
  \]

  \[
  \frac{1}{[A]_t} = kt + \frac{1}{[A]_0}
  \]

  \[
  \text{half life} = t_{\frac{1}{2}} = \frac{1}{k[A]_0}
  \]

- half life dependent on initial concentration of reactant

14.5: Temperature and Rate

- rate constant must increase with increasing temperature, thus increasing the rate of reaction

14.5.1 The Collision Model

- **collision model** – molecules must collide to react
- greater frequency of collisions the greater the reaction rate
- for most reactions only a small fraction of collisions leads to a reaction

14.5.2 Activation Energy

- Svante August Arrhenius
- Molecules must have a minimum amount of energy to react
- Energy comes from kinetic energy of collisions
- Kinetic energy used to break bonds
- Activation energy, \(E_a\) – minimum energy required to initiate a chemical reaction
• Activated complex or transition state – atoms at the top of the energy barrier
• Rate depends on temperature and $E_a$
• Lower $E_a$ means faster reaction
• Reactions occur when collisions between molecules occur with enough energy and proper orientation

14.5.3 The Arrhenius Equation

• reaction rate data:
• the Arrhenius Equation:

\[ k = A e^{\frac{-E_a}{RT}} \]

• \(k\) = rate constant, \(E_a\) = activation energy, \(R\) = gas constant (8.314 J/(mol K)), \(T\) = absolute temperature, \(A\) = frequency factor
• \(A\) relates to frequency of collisions, favorable orientations

\[ \ln k = -\frac{E_a}{RT} + \ln A \]

• the \(\ln k\) vs. \(1/\text{T}\) graph (also known as an Arrhenius plot) has a slope \(\frac{-E_a}{R}\) and the y-intercept \(\ln A\)
• for two temperatures:

\[ \ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right) \]

• used to calculate rate constant, \(k_1\) and \(T_1\)

14.6: Reaction Mechanisms

• reaction mechanism – process by which a reaction occurs

14.6.1 Elementary Steps

• elementary steps – each step in a reaction
• molecularity – if only one molecule involved in step
• unimolecular – if only one molecule involved in step
• bimolecular – elementary step involving collision of two reactant molecules
• termolecular – elementary step involving simultaneous collision of three molecules
• elementary steps in multi-step mechanism must always add to give chemical equation of overall process
• intermediate – product formed in one step and consumed in a later step

14.6.2 Rate Laws of Elementary Steps

• if reaction is known to be an elementary step then the rate law is known
• rate of unimolecular step is first order (Rate = k[A])
• rate of bimolecular steps is second order (Rate = k[A][B])
• first order in \([A]\) and \([B]\)
• if double \([A]\) than number of collisions of \(A\) and \(B\) will double

14.6.3 Rate Laws of Multi-step Mechanisms

• rate-determining step – slowest elementary step
• determines rate law of overall reaction

14.6.4 Mechanisms with an Initial Slow Step vs. Mechanisms with an Initial Fast Step

• intermediates are usually unstable, in low concentration, and difficult to isolate
• when a fast step precedes a slow one, solve for concentration of intermediate by assuming that equilibrium is established in fast step

---

**14.7: Catalysis**

• catalyst – substance that changes speed of chemical reaction without undergoing a permanent chemical change

14.7.1 Homogeneous Catalysis

• homogeneous catalyst – catalyst that is present in same phase as reacting molecule
• catalysts alter \(E_a\) or \(A\)
• generally catalysts lowers overall \(E_a\) for chemical reaction
• catalysts provides a different mechanism for reaction

14.7.2 Heterogeneous Catalysis

• exists in different phase from reactants
• initial step in heterogeneous catalyst is adsorption
• adsorption – binding of molecules to surface
• adsorption occurs because ions/atoms at surface of solid extremely reactive

14.7.3 Enzymes

• biological catalysts
• large protein molecules with molecular weights 10,000 – 1 million amu
• catalase – enzyme in blood and liver that decomposes hydrogen peroxide into water and oxygen
• substrates – substances that undergo reaction at the active site
• lock-and-key model – substrate molecules bind specifically to the active site
• enzyme-substrate complex – combination of enzyme and substrate
• binding between enzyme and substrate involves intermolecular forces (dipole-dipole, hydrogen bonding, and London dispersion forces)
• product from reaction leaves enzyme allowing for another substrate to enter enzyme
• enzyme inhibitors – molecules that bind strongly to enzymes
• **turnover number** – number of catalyzed reactions occurring at a particular active site
• large turnover numbers = low activation energies