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)

2. **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)

3. **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)
4.1: General Properties of Aqueous Solutions
4.2: Precipitation Reactions
4.3: Acid-Base Reactions
4.4: Oxidation-Reduction Reactions
4.5: Concentration of Solutions
4.6: Solution Stoichiometry and Chemical Analysis
4.E: Reactions in Aqueous Solution (Exercises)
4.S: Reactions in Aqueous Solution (Summary)

5.1: The Nature of Energy
5.2: The First Law of Thermodynamics
5.3: Enthalpy
5.4: Enthalpy of Reaction
5.5: Calorimetry
5.6: Hess’s Law
5.7: Enthalpies of Formation
5.8: Foods and Fuels
5.E: Thermochemistry (Exercises)
5.S: Thermochemistry (Summary)

6.1: The Wave Nature of Light
6.2: Quantized Energy and Photons
6.3: Line Spectra and the Bohr Model
6.4: The Wave Behavior of Matter
6.5: Quantum Mechanics and Atomic Orbitals
6.6: 3D Representation of Orbitals
6.7: Many-Electron Atoms
6.8: Electron Configurations
6.9: Electron Configurations and the Periodic Table
6.E: Electronic Structure of Atoms (Exercises)
6.S: Electronic Structure of Atoms (Summary)
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)

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)

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)
Chapter 10: Gases

10.1: Characteristics of Gases
10.2: Pressure
10.3: The Gas Laws
10.4: The Ideal Gas Equation
10.5: Further Applications of the Ideal-Gas Equations
10.6: Gas Mixtures and Partial Pressures
10.7: Kinetic-Molecular Theory
10.8: Molecular Effusion and Diffusion
10.9: Real Gases - Deviations from Ideal Behavior
10.E: Exercises
10.S: Gases (Summary)

Chapter 11: Liquids and Intermolecular Forces

11.1: A Molecular Comparison of Gases, Liquids, and Solids
11.2: Intermolecular Forces
11.3: Some Properties of Liquids
11.4: Phase Changes
11.5: Vapor Pressure
11.6: Phase Diagrams
11.7: Structure of Solids
11.8: Bonding in Solids
11.E: Liquids and Intermolecular Forces (Exercises)
11.S: Liquids and Intermolecular Forces (Summary)

Chapter 12: Solids and Modern Materials

12.1: Classes of Materials
12.2: Materials for Structure
12.3: Materials for Medicine
12.4: Materials for Electronics
12.5: Materials for Optics
12.6: Materials for Nanotechnology
12.E: Solids and Modern Materials (Exercises)
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. 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: Noble 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
10. 23.E: Metals and Metallurgy (Exercises)

• 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)
6. 6.E: Electronic Structure (Exercises)
7. 7.E: Periodic Trends (Exercises)
8. 8.E: Chemical Bonding Basics (Exercises)
9. 9.E: Bonding Theories (Exercises)
10. 10.E: Gases (Exercises)
11. 11.E: Liquids and Intermolecular Forces (Exercises)
13. 13.E: Properties of Solutions (Exercises)
15. 15.E: Chemical Equilibrium (Exercises)
16. 16.E: Acid–Base Equilibria (Exercises)
17. 17.E: Additional Aspects of Aqueous Equilibria (Exercises)
18. 18.E: Chemistry of the Environment (Exercises)
19. 19.E: Chemical Thermodynamics (Exercises)
20. 20.E: Electrochemistry (Exercises)
21. 21.E: Nuclear Chemistry (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)

21.1: Radioactivity

• nucleons – neutron and proton
- all atoms of a given element have the same number of protons, atomic number
- isotopes – atoms with the same atomic number but different mass numbers
- three isotopes of uranium: uranium-233, uranium-235, uranium-238
  \(^{233}_{92}U, ^{235}_{92}U, ^{238}_{92}U\)
  (superscript is mass number, subscript atomic number)
- radionuclides – nuclei that are radioactive
- radioisotopes – atoms containing radionuclides

21.1.1 Nuclear Equations

- alpha particles – helium-4 particles
- alpha radiation – stream of alpha particles
- emission of radiation is one way that an unstable nucleus is transformed into a more stable one

\[^{238}_{92}U \rightarrow ^{234}_{92}Th + ^{4}_{2}He\]
- superscript = mass number
- subscript = atomic number
- radioactive decay – when a nucleus spontaneously decomposes
- sum of the mass numbers is the same on both sides of the equation
- sum of the atomic numbers same on both sides of the equation
- radioactive properties of the nucleus are independent of the state of chemical combination of the atom
- chemical form does not matter when writing nuclear equations

21.1.2 Types of Radioactive Decay

- three most common type of radioactive decay: alpha(\(\alpha\)), beta(\(\beta\)), and gamma(\(\gamma\)) radiation

<table>
<thead>
<tr>
<th>Property</th>
<th>(\alpha)</th>
<th>(\beta)</th>
<th>(\gamma)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Charge</td>
<td>2+</td>
<td>1-</td>
<td>0</td>
</tr>
<tr>
<td>Mass</td>
<td>(6.64 \times 10^{-24}) g</td>
<td>(9.11 \times 10^{-28}) g</td>
<td>0</td>
</tr>
<tr>
<td>Relative penetrating power</td>
<td>1</td>
<td>100</td>
<td>10,000</td>
</tr>
<tr>
<td>Nature of radiation</td>
<td>(^{\alpha}\text{He}) (\neq \text{nuclei})</td>
<td>electrons</td>
<td>High-energy photons</td>
</tr>
</tbody>
</table>

- beta particles – high speed electrons emitted by an unstable nucleus

\[^{131}_{53}I \rightarrow ^{131}_{54}Xe + ^{0.1}_{-1}e\]
- beta decay results in increasing the atomic number
- $^0_1n \rightarrow ^1_1p + ^0_1e$
- **gamma radiation** – high-energy protons
- gamma radiation does not change atomic number or mass number or a nucleus
- almost always accompanies other radioactive emission
- represents the energy lost when the remaining nucleons reorganize into more stable arrangements
- **positron** – particle that has same mass as an electron but opposite charge
- represented by $^0_1e$
- emission of a positron has effect of converting a proton to a neutron and decreasing atomic number of nucleus by 1
- **electron capture** – the capture by the nucleus of an inner-shell electron from the electron cloud surrounding the nucleus
- has effect of converting a proton to neutron
- $^1_1p + ^0_1e \rightarrow ^0_1n$

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutron</td>
<td>( ^1_0n )</td>
</tr>
<tr>
<td>Proton</td>
<td>( ^1_1H ) or ( ^1_1p )</td>
</tr>
<tr>
<td>Electron</td>
<td>( ^0_{-1}e )</td>
</tr>
<tr>
<td>Alpha Particle</td>
<td>( ^4_2\text{He} ) or ( ^4_2\text{\alpha} )</td>
</tr>
<tr>
<td>Beta Particle</td>
<td>( ^0_{-1}\text{e} ) or ( ^0_{-1}\text{\beta} )</td>
</tr>
<tr>
<td>Positron</td>
<td>( ^0_1e )</td>
</tr>
</tbody>
</table>

### 21.2: Patterns of Nuclear Stability

#### 21.2.1 Neutron-to-Proton Ratio
- **strong nuclear force** – a strong force of attraction between a large number of protons in the small volume of the nucleus
- stable nuclei with low atomic numbers up to 20 have nearly equal number of neutrons and protons
- for higher atomic numbers, the number of neutrons are greater than the number of protons
- the neutron-to-proton ratio of stable nuclei increase with increasing atomic number
- **belt of stability** – area where all stable nuclei are found
  - ends at bismuth
  - all nuclei with 84 or more protons are radioactive
• an even number of protons and neutrons is more stable than an odd number

  ◦ determining type of radioactive decay
    ▪ 1) nuclei above the belt of stability
    ▪ high neutron-to-proton ratios
    ▪ move toward belt of stability by emitting a beta particle
    ▪ decreases the number of neutrons and increases the number of protons in a nucleus
    ▪ 2) nuclei below the belt of stability
    ▪ low neutron-to-proton ratios
    ▪ move toward belt of stability by positron emission or electron capture
    ▪ increase number of neutrons and decrease the number of protons
    ▪ positron emission more common with lower nuclear charges
    ▪ electron capture becomes more common with increasing nuclear charge
    ▪ 3) nuclei with atomic numbers

  Image81.gif

  ◦ alpha emission
  ◦ decreases both number of neutrons and protons by 2

21.2.2 Radioactive Series

• ◦ some nuclei cannot game stability by a single emission
  ◦ radioactive series or nuclear disintegration series – series of nuclear reactions that begin with an unstable nucleus to a stable one
  ◦ three types of radioactive series found in nature
    ▪ uranium-238 to lead-206, uranium-235 to lead-207, and thorium-232 to lead-208

21.2.3 Further Observations

• ◦ nuclei with 2, 8, 20, 28, 50, or 82 protons or 2, 8, 20, 28, 50, 82, or 126 neutrons are more stable than with nuclei without these numbers
  ◦ numbers called magic numbers
  ◦ nuclei with even number of protons and neutrons more stable than with odd number of protons and neutrons
  ◦ observations made in terms of the shell model of the nucleus
    ▪ nucleons reside in shells
  ◦ magic numbers represent closed shells in nuclei

21.3: Nuclear Transmutations

• ◦ nuclear transmutations – nuclear reactions caused by the collision of one nucleus with a neutron or by another nucleus
  ◦ first conversion of one nucleus into another performed by Ernest Rutherford in 1919
• converted nitrogen-14 to oxygen-17
  \[ ^{14}_7 \text{N} + ^{4}_2 \text{He} \rightarrow ^{17}_8 \text{O} + ^{1}_1 \text{H} \]
  \[ ^{14}_7 \text{N} (\alpha, p) ^{17}_8 \text{O} \]

21.3.1 Using Charged Particles
• particle accelerators – used to accelerate particles at very high speeds
  • cyclotron, and synchrotron

21.3.2 Using Neutrons
• neutrons do not need to be accelerated

21.3.4 Transuranium Elements
• transuranium elements – elements with atomic numbers above 92 that are produced by artificial transmutations

21.4: Rates of Radioactive Decay
• radioactive decay is a first-order process
  • has characteristic of half life, which is the time required for half of any given quantity of a substance to react
  • half-life unaffected by external conditions

21.4.1 Dating
• radiocarbon dating assumes that the ratio of carbon-14 to carbon-12 in the atmosphere has been constant for at least 50,000 years
  • age of rocks can be determined by ratio of uranium-238 to lead-206

21.4.2 Calculations Based on Half-life
• rate = kN
  \( \ln \frac{N_t}{N_0} = -k t \)
  \( k = \frac{0.693}{t_{1/2}} \)

21.5 Detection of Radioactivity
• Geiger counter – device used to measure and detect radioactivity
  • Based on ionization of matter caused by radiation
  • Phosphors – substances that give off light when exposed to radiation
  • Scintillation counter – used to detect and measure radiation based on tiny flashes of light produced when
radiation strikes a suitable phosphor

21.5.1 Radiotracers

- radioisotopes can be used to follow an element through its chemical reactions
- isotopes of same element have same properties
- radiotracer – radioisotopes used to trace an element

21.6: Energy Changes in Nuclear Reactions

\[ E = mc^2 \]

E = energy, m = mass, c = speed of light

If system loses mass, it loses energy (exothermic)

If system gains mass, it gains energy (endothermic)

21.6.1 Nuclear Binding Energies

- masses of nuclei always less than masses of individual nucleons
- **mass defect** – mass difference between a nucleus and its constituent nucleons
- energy is needed to break nucleus into separated protons and neutrons, addition of energy must also have an increase in mass
- nuclear binding energy – energy required to separate a nucleus into its individual nucleons
  - the larger to nuclear binding energy the more stable the nucleus toward decomposition
- **fission** – energy produced when heavy nuclei split
- **fusion** – energy produced when light nuclei fuse

21.7: Nuclear Fission

- fission and fusion both exothermic
- chain reaction – reaction in which the neutrons produced in one fission cause further fission reactions
- in order for a fission chain reaction to occur, the sample of fissionable material must have a certain minimum mass
- **critical mass** – amount of fissionable material large enough to maintain the chain reaction with a constant rate of fission
- **supercritical mass** – mass in excess of a critical mass

21.7.1 Nuclear Reactors

- nuclear reactors the fission is controlled to generate a constant power
- reactor core consists of fissionable fuel, control rods, a moderator, and cooling fluid
• fission products are extremely radioactive and are thus hard to store
• about 20 half-lives needed for products to react acceptable levels for biological exposure

21.8: Nuclear Fusion

• fusion is appealing because of availability of light isotopes and fusion products are not radioactive
• high energies needed to overcome attraction of nuclei
• thermonuclear reactions – fusion reactions
• lowest temperature required is about 40,000,000 K

21.9: Biological Effects of Radiation

• when matter absorbs radiation, the energy of the radiation can cause either excitation or ionization
• ionization radiation more harmful than nonionization radiation
• most of energy of radiation absorbed by water molecules
• free radical – a substance with one or more unpaired electrons
• can attack other biomolecules to produce more free radicals
• gamma rays most dangerous
• tissues that take most damage are the ones that reproduce at a rapid rate
• bone marrow, blood forming tissues, lymph nodes

21.9.1 Radiation Doses

• ◦ becquerel (Bq) – SI unit for activity of the radiation source; rate at which nuclear disintegrations are occurring
  ◦ 1 (Bq) = 1 nuclear disintegration/s
  ◦ curie (Ci) = 3.7x10^{10} disintegrations/s = rate of decay of 1g of radium
  ◦ two units used to measure amount of exposure to radiation: gray (Gy) and rad
  ◦ gray – SI unit of absorbed dose = absorption of 1 J of energy per kilogram of tissue
  ◦ rad (radiation absorbed dose) – absorption of 1x10^{-2} J of energy per kilogram of tissue
  ◦ 1 Gy = 100 rads
  ◦ relative biological effectiveness – RBE
    • 1 for gamma and beta radiation, 10 for alpha radiation
    • exact value varies with dose rate, total dose, and type of tissue affected
    • rem (roentgen equivalent for man) – product of the radiation dose in rads and the RBE of the radiation
gives the effective dosage
    • rem is unit of radiation damage that is usually used in medicine
    • number of rems = (number of rads)(RBE)
  ◦ Sievert (Sv) – SI unit for dosage
    • 1 Sv = 100 rem
- annual exposure = 360mrem

21.9.2 Radon

- radon exposure estimated to account for more than half annual exposure
- half-life of radon is 3.82 days
- decays into radioisotope polonium
- atoms of polonium can be trapped in lungs giving out alpha radiation causing lung cancer
- recommended levels of radon-222 in homes is to be less than 4 pCi per liter of air