CHEM 1210 Lecture Notes
OpenStax Chemistry 2e

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COURSE ADMINISTRATIVE DETAILS

- My office hours
- Intro to my research
- Introductory Quiz
- Grading details
    - Exams - 40, Final - 20, Online Homework - 15, Book Homework - 15, Quizzes - 10
    - Online homework
    - Frequent quizzes
- Importance of reading and learning on your own
- Learning resources
    - My Office Hours
    - Tutoring services - https://www.suu.edu/academicsuccess/tutoring/
- Show how to access Canvas
    - Calendar, Grades, Modules, etc.
    - Quizzes
    - Textbook
- Introduction to chemistry
    - Ruby fluorescence
    - Levomethamphetamine
    - Submerged salt crystals grow due dynamic equilibrium
    - Rubber band elasticity
    - Structure of the periodic table
    - Salt on ice and purifying hydrogen peroxide

CHAPTER 1

ESSENTIAL IDEAS

1.1 Chemistry in Context

- Modern chemistry is the end result of thousands of years of humans trying to explain and control the materials around them
- Early forays into chemistry (such as alchemy) had deep mystical roots and often relied on serendipity to make good progress
- Modern chemistry is a rigorous science, relying on falsifiability and the scientific methd (Figure 1.4)
- We sometimes refer to chemistry as “The Central Science” (Figure 1.3)
- To adequately describe and understand chemical phenomena, we often talk from different perspectives

  Macroscopic Domain This is what we observe with bulk substances. Two chemicals react to produce a new chemical

  Microscopic Domain We now understand that all microscopic effects are governed by the behavior of *microscopic* actors (molecules, atoms, electrons, etc.)

  Symbolic Domain Effectively communicating chemical ideas requires new language. Chemical formulas, equations, and mechanisms are all symbolic representations

  All three domains are on display in Figure 1.5

1.2 Phases and Classification of Matter

- Three primary phases of matter are shown in Figure 1.5 (and 1.6)
- Plasmas are like a gas, but with electrically charged particles
- Mass vs Weight (for very fine measurements, the difference matters even on Earth due to buoyancy)
- Figure 1.8 illustrates the *law of conservation of matter*
- Classifying matter (Figure 1.11)
  - Pure Substances
    - Elements (Anything on the *periodic table of the elements*)
    - Compounds (Combinations of elements – can have very different properties from their constituent elements)
  - Mixtures

        * Heterogeneous mixtures (variable composition)
        * Homogeneous mixtures (i.e. solutions, continuous composition)

- Table 1.1 shows the abundance of many elements on Earth
- Atoms are the smallest particle of an element that has the properties of that element
    - Thought-experiment of dividing a sample in half ad-infinitum
    - Ancient atomic theories and modern Dalton atomic theory (discussed in detail later)
    - Atoms are *very* small; smaller than we could even detect until recently
- Molecules are collections of atoms held together with chemical bonds (more nuanced definition later)
    - Many elements occur naturally as molecules, rather than atoms
    - Figure 1.14 shows many molecular elements and compounds

1.3 Physical and Chemical Properties

- Physical Properties: Properties which can be observed without changing the chemical identity of the substance
- Chemical Properties: Properties which can only be observed through chemical reactions (e.g. flammability, acidity, electrochemical potential, etc.)
- Physical Changes: Any change which perserves the *chemical identity* of the substance (including phase changes)
- Chemical Changes: Changes which alter the chemical identities of one of more substance
- Extensive Properties: Depend on the size of hte system (double the size, double the property measurement, such as mass or volume)
- Intensive Properties: Independent of system size (density, temperature, most chemical properties)
- The periodic table groups elements according to their properties (Figure 1.22)
    - Metals conduct electricity and heat, are maleable and ductile
    - Non-metals are very diverse, but generally poor conductors
    - Metalloids exist at the boundary and share properties with both metals and non-metals
    - There are many other ways to group the elements, which we will learn later

1.4 Measurements

- All measurements are composed of three parts:
  - The magnitude of the measurement (the number itself)
  - The unit of measurement used (g, kg, lbs, etc.)
  - The degree of uncertainty in the measurement (this is usually implicit, and covered in the next section)
- Units are an essential part of any measuement. Develop a habit of *always* including units in your work
  - $u_{rms} = \sqrt{\dfrac{3RT}{M}}$ – example of how units can guide problem solving and "unit purgatory"
  - SI units are a collection of fundamental units from which all other units can be derived (Table 1.2)
  - Metric prefixes make it more convenient to discuss very large or very small numbers (Table 1.3)
  - Scientific notation is an even more general and robust way of representing numbers
    * The quantity is represented by a number with the decimal after the first digit
    * The magnitude is represented by a power of 10
  - Practice converting between normal numbers, metric prefixes, and scientific notation
  - For temperature, we use both K and $^\circ C$ (But not $^\circ F$)
    $T(K) = T(^\circ C) + 273.15$
  - Derived units will combine the fundamental units in some way
    volume: m^3, L, ml
    velocity: ${}^{m}/_{s}$
    density: ${}^{kg}/_{m^3}$, ${}^{g}/_{cm^3}$ (Table 1.4)
    energy: $1J \equiv {}^{kgm^2}/_{s^2}$

1.5 Measurement Uncertainty, Accuracy, and Precision

- Countable quantities are considered to be *exact* (no uncertainty)
- Measurements (and groups of measurements) always have some degree of undertainty
  - Accuracy is how close a measurement is to the *true value* (usually unknown, but approximated by calibration with a well-known standard)
  - Precision is how finely a measurment is made (What is the margin of error)
  - Figure 1.27 and Table 1.5 illustrate the differences between precision and accuracy
  - Accuracy is usually improved through calibration, and moving forward we will usually assume that measurements are as accurate as an instrument allows
  - Precision is represented in the way we write the number, and can be improved with a better instrument or with repeat measurements

- Significant figures are the way that we represent precision in a number
  - The number of digits conveys the degree of precision
  - Example of me saying I'm $6ft\ 2in$ tall, vs me saying I'm $6ft\ 1.6241434in$ tall
  - For graduated measurements, we record one digit beyond the lowest graduation (Figure 1.26)
  - For digital measurements, we record the number as it is given by the instrument
  - For any given number, we should track both the *quantity* of significant figures, and the *position* of the least-significant digit
  - In a written number, digits are considered significant according to the following rules:
    - All non-zeros are significant
    - All *captive* zeros (between two other significant digits) are significant
    - Trailing zeros are *always* significant
    - Leading zeros are *never* significant
    - For scientific notation, only the digits of the quantity (not the magnitude) count
    - Logarithmic quantities follow different rules which we will revisit in CHEM 1220 (chapter 14)
    - Note that for some numbers scientific notation is *required* to convey the correct precision $(3.0 \times 10^3 m)$
- Errors propogate when multiple measurements are used in a mathematical operation
  - For addition and subtraction, the least significant digit of the answer will be in the same position as the least significant digit of hte least precise input
  - For multiplicationa and division, the quantity of significant digits in the answer will match the quantity of significant digits of the input with fewest significant digits
  - When rounding an exact 5 (no further digits beyond the 5), round up or down to make the last digit even
  - Compound problems involve multiple types of operations
    - Solve the problem in steps, applying the correct rule to each step
    - Track the significant figures (quantity and position) for each intermediate answer, but do *not* truncate or round any of these answers
    - Only round after the last step
      - Practice $\frac{12.3g+34g}{12.0cm^3+7.7cm^3} = 2.4 g/cm^3$ (wrong answer with premature rounding)

1.6 Mathematical Treatment of Measurement Results

CHAPTER 2

ATOMS, MOLECULES, AND IONS

CHAPTER 3

COMPOSITION OF SUBSTANCES AND SOLUTIONS

3.1 Formula Mass and the Mole Concept

3.2 Determining Empirical and Molecular Formulas

3.3 Molarity

3.4 Other Units for Solution Concentration

CHAPTER 4

STOICHIOMETRY OF CHEMICAL REACTIONS

CHAPTER 5

THERMOCHEMISTRY

5.1 Energy Basics

5.2 Calorimetry

5.3 Enthalpy

CHAPTER 6

ELECTRONIC STRUCTURE AND PERIODIC PROPERTIES OF ELEMENTS

6.1 Electromagnetic Energy

6.2 The Bohr Model

6.3 Development of Quantum Theory

6.4 Electronic Structure of Atoms (Electron Configurations)

6.5 Periodic Variations in Element Properties

CHAPTER 7

CHEMICAL BONDING AND MOLECULAR GEOMETRY

7.1 Ionic Bonding

7.2 Covalent Bonding

7.3 Lewis Symbols and Structures

7.4 Formal Charges and Resonance

7.5 Strengths of Ionic and Covalent Bonds

7.6 Molecular Structure and Polarity

CHAPTER 8

ADVANCED THEORIES OF COVALENT BONDING

8.1 Valence Bond Theory

8.2 Hybrid Atomic Orbitals

8.3 Multiple Bonds

8.4 Molecular Orbital Theory

CHAPTER 9

GASES

9.1 **Gas Pressure**

9.2 **Relating Pressure, Volume, Amount, and Temperature: The Ideal Gas Law**

9.3 **Stoichiometry of Gaseous Substances, Mixtures, and Reactions**

9.4 **Effusion and Diffusion of Gases**

9.5 **The Kinetic-Molecular Theory**

9.6 **Non-Ideal Gas Behavior**

CHAPTER 10

LIQUIDS AND SOLIDS

10.1 Intermolecular Forces

ERRATA