

CHEM 1210 Lecture Notes

OpenStax Chemistry 2e

Matthew Rowley

October 10, 2025

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 method (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

Quiz 1.1 - Scientific Method**Homework 1.1**

- 1: Thinking in terms of Chemistry
- 3: The scientific method
- 5: Domains of inquiry

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 preserves the *chemical identity* of the substance (including phase changes)
- Chemical Changes: Changes which alter the chemical identities of one or more substance
- Extensive Properties: Depend on the size of the 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 malleable 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

Quiz 1.2 - Matter, Properties, and Change

Homework 1.2

- 17: Classifying matter
- 27: Classifying changes

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 measurement. Develop a habit of *always* including units in your work

- $u_{rms} = \sqrt{\frac{3RT}{M}}$ – 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 uncertainty

- 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 measurement 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 *6 ft 2 in* tall, vs me saying I'm *6 ft 1.6241434 in* 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 propagate 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 the least precise input
 - For multiplication 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 $\frac{12.3g + 34g}{12.0cm^3 + 7.7cm^3} = 2.4g/cm^3$ (wrong answer with premature rounding)

Quiz 1.3 - Significant Figures

Homework 1.3

- 45: Scientific Notation
- 49: Counting Significant Figures
- 53: Significant Figures and Calculations

1.6 Mathematical Treatment of Measurement Results

- Some quantities are calculated based on two or more measurements (such as velocity and density)
- These formulas can be used to relate all three quantities together (i.e. $velocity = \frac{distance}{time}$)
- The derived quantity can be interpreted as a *conversion factor* between the other two quantities
- Conversion factors and unit conversions
 - Elementary school perspective of *ft* to *in* conversions
 - Conversion factors are a ratio between two identical quantities
 - Converting units involves multiplying by 1 in the form of a conversion factor
 - Units guide the problem solving
- Dimensional Analysis is a problem-solving framework based on a series of unit conversions
 - Don't dive straight into calculations and equations
 - Identify the units you expect for the answer
 - Identify the starting point
 - Create a plan to convert units from the starting point to the answer
 - Carry out the calculations
 - Practice converting 65.0 miles/hour into m/s
 - The "railroad ties" or "picket fence" method can help organize your work
- Dimensional analysis is not the only way to solve problems, but it is versatile and robust; usually my preferred choice
- Practice a more abstract problem:
Find the $miles/gal$ if a car consumes 8036 g of gasoline while driving for 40.0 min at 75 $miles/hour$

Quiz 1.4 - Dimensional Analysis

Homework 1.4

- 65: Simple unit conversion
- 87: Density from volume and mass
- 89: Mass from volume
- 91: Volume from mass

CHAPTER 2

ATOMS, MOLECULES, AND IONS

2.1 Early Ideas in Atomic Theory

- 1807 Dalton's Atomic Theory: (1, 2 and 5 are not strictly true) (Figures 2.2-2.4)
 1. Matter is composed of atoms
 2. Atoms of a given element all have identical properties to each other
 3. Atoms of one element differ in properties from elements of a different element
 4. Chemical compounds consist of atoms of different elements combined in a specific ratio
 5. Chemical reactions *rearrange* the atoms which are already there, but cannot create or destroy atoms
- Development of Dalton's theory:
 - Dalton relied on prior work by Proust who demonstrated the law of definite proportions (Table 2.1)
 - This was not at all expected - my analogy with bread, or metal alloys
 - Dalton further noted that ratios of these proportions followed the law of multiple proportions (Copper (I or II) Chloride example in the book)

2.2 Evolution of Atomic Theory

- About a century later, scientists discovered that atoms are made of even smaller components
- J.J. Thomson discovered the electron, and its charge/mass ratio (Figure 2.6)
- Millikan's oil drop experiment found the fundamental charge (and thus mass) of an electron (Figure 2.7)
- Figure 2.8 shows some early ideas of how the positive and negative charges were distributed in an atom
- Ernest Rutherford discovered the atomic nucleus, consisting of very concentrated positive charge (Figures 2.9 and 2.10)

“It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.”
- Different *isotopes* of atoms were discovered with techniques that produced isotopically pure samples
- Finally, the neutron itself was discovered in 1932, explaining what particle led to different isotopes

2.3 Atomic Structure and Symbolism

- Atoms are made up of protons, neutrons, and electrons
- Figure 2.11 shows the small scale of the atom and nucleus
- Atomic units make discussions about atoms convenient
 - The Atomic Mass Unit amu , Da , or $u = 1.6605 \times 10^{-24}g$
 - The fundamental charge $e = 1.602 \times 10^{-19}C$
 - The Angstrom $\text{\AA} = 10^{-10}m$
- Table 2.2 summarizes the properties of elementary particles in atoms
- We track the composition of an atom with three numbers:
 - The atomic number is the number of protons $Z = p$
 - The mass number is the number of protons and neutrons $A = p + n$
 - The number of neutrons is therefore $n = A - Z$
 - The charge is the protons minus the electrons $q = p - e$
 - The number of electrons is $e = p - q$
- A positively charged atom is called a cation, and a negatively charged atom is called an anion
- Chemical symbols are a shorthand way of representing everything we need about an atom
 - There is a 1 or 2 letter symbol for each element (Table 2.3 shows some make sense, some don't)
 - A is written as a left superscript
 - Z is written as a left subscript, but can be left off
 - q is written as a right superscript with the magnitude first, then the sign. q is left off if $q = 0$
 - Example: ${}^{13}_6C^{2+}$ (6 protons, 7 neutrons, 4 electrons)
- Isotopes are different versions of elements with different mass numbers
 - For the most part, different isotopes of an element behave exactly the same in chemistry
 - Isotope abundance can be found by mass spectrometry, among other methods (Figure 2.15)
 - Table 2.4 shows the natural abundances of the isotopes of several light elements
 - Atomic weight (atomic mass) is the weighted average of all the isotopes found on the Earth
$$M = \sum_i mass_i \times \frac{\%abundance_i}{100\%}$$

2.4 Chemical Formulas

- We can represent the actual structure and makeup of molecules at several levels of abstraction (Figures 2.16 and 2.17)
- Molecular formulas
 - Each element is listed, with the number of atoms for each element written as a subscript (H_2O)
 - The order of elements follows certain patterns, with the least electronegative element often written first
- Structural Formulas show how atoms are connected with covalent bonds represented as lines
- Ball and Stick models show the three-dimensional geometry of a molecule
- Space-filling models show the actual volume of space taken up by each atom in a molecule
- Figure 2.18 illustrates the difference between subscripts within a formula, and stoichiometric coefficients in front of formulas
- Empirical formulas show the mathematically simplified ratios of elements
 - Some experimental techniques (especially early ones) could *only* give the empirical formula
 - Ionic compounds are always reported with the empirical formula
 - To find the empirical formula, divide all subscripts by their greatest common factor
 - Some very different compounds share an empirical formula (carbohydrates CH_2O include formaldehyde, acetic acid, and sugar)
 - The molecular formula can be calculated from the empirical formula's weight and the molecular weight (Chapter 3)
- Arranging the same group of atoms in different ways produces different isomers
 - Isomers share the same chemical formula, but can have very different properties
 - Structural isomers differ in how the atoms are connected to each other (Figure 2.23)
 - Optical isomers (or spatial isomers) are non-superimposable mirror images (Figure 2.24, glove analogy)

Quiz 2.1 - Atomic Theories

Homework 2.1

- 7: Properties of protons and neutrons
- 11: Atomic symbols from composition
- 19: Composition from atomic symbols
- 23: Atomic weight
- 29: Molecular and empirical formulas

2.5 The Periodic Table

- As scientists discovered and studied more and more elements, they started to notice certain natural groupings according to physical and chemical properties
- Mendeleev arranged the atoms according to these groups and atomic weight, producing the first periodic table (Figure 2.25)
- Mendeleev even predicted the existence of and properties of yet-undiscovered elements based on gaps in his table
- Vocabulary around the periodic table:
 - Rows are periods or series
 - Columns are groups or families
- Figure 2.26 is a typical periodic table, showing the metals, non-metals and metalloids (contrast with my preferred table)
- Figure 2.27 shows many of the names we use for important groups of elements
- The structure of the periodic table encodes rich information about the electrons in the elements, as we will learn in chapter 6

Quiz 2.2 - Periodic Table

Homework 2.2

- 37: Classifying elements
- 41: Using group names

2.6 Ionic and Molecular Compounds

- An atom which gains or loses electrons (carries a charge) is called an *ion*
 - Positively charged ions are called cations, and are smaller than their neutral atom counterparts (Figure 2.28)
 - Negatively charged ions are called anions, and are larger than their neutral atom counterparts
 - We can predict which charge different elements will naturally tend to take based on their position in the periodic table (Figure 2.29)
 - Many other elements can take two or more charges, especially the transition metals
- Some ions are composed of more than one atom and are called polyatomic ions
 - Table 2.5 gives some common polyatomic ions. Memorize the formula, name, and charge of these and their acids

- Notice some trends in the names of oxyanions (per-ate, -ate, -ite, and hypo-ite)
- We will learn about the acid names listed here (and more) in section 2.7
- Ionic compounds are held together by ionic bonds (coulombic attractions)
 - Show dot diagram of how NaCl and CaCl form from elements
 - Metal + non-metal is not an adequate definition of ionic compounds (NH_4NO_3)
 - Ionic compounds form an extended lattice of ions (covered more in CHEM 1220)
 - Ions will combine to form neutral compounds
 - Practice producing compound formulas from ions (include paranthesis for polyatomic ions where appropriate)
 - Practice finding ions from formulas of compounds
- Molecular compounds are held together by covalent bonds (shared electrons)
 - Show a dot diagram of H_2 , H_2O , O_2 , and N_2
 - Bonds between non-metals are covalent bonds
 - Molecular compounds combine non-metals into discrete units called molecules
 - Single, double, and triple bonds involve sharing two, four, and six electrons

2.7 Chemical Nomenclature

- Naming ionic compounds
 - Cation names are the name of the element, with the charge in roman numerals in parenthesis *if* the element could take multiple charges
 - Anion names are the element name with an “-ide” ending (some element like P remove more than others)
 - Polyatomic ion names are the same as you learned earlier
 - For ionic compounds the name is simply cation name + anion name
 - There is no indication of the quantity of atoms, that is inferred from charge neutrality
 - Practice getting formulas from names and names from formulas
- Naming hydrates
 - Some ionic compounds will incorporate water into their ionic lattice
 - The formulas will have a \cdot then list the number of waters
 - The names add the degree of hydration using prefixes from table 2.10 and adding “hydrate”
 - The waters can be driven off with high temperature, producing the *anhydrous* form

Quiz 2.3 - Naming Ionic Compounds

Homework 2.3

- 47: Predicting bond type in compounds
- 49: Formulas from ions
- 51: Names from formulas
- 57: Names from formulas with transition metals
- 59: Formulas from names

Resuming section 2.7 Chemical Nomenclature

- Naming molecular compounds
 - There are many ways to name molecular compounds, we will focus on just two here
 - Naming binary molecular compounds
 - * # + name + # + name with “-ide” ending
 - * Least electronegative element (leftmost on the periodic table) goes first
 - * #s come from table 2.10
 - * Omit “mono-” for the first element
 - * Practice going from formula to name and vice-versa (Table 2.11)
 - Naming molecular acids
 - * Molecular acid names are based on the name of the anion formed when all H^+ are removed
 - * -ide ions form hydro-ic acids (Table 2.12)
 - * -ate ions form -ic acids
 - * -ite ions form -ous acids
 - * preserve the hypo- and per- prefixes
 - * Table 2.13 shows some oxyacid names

Quiz 2.4 - Naming Molecular Compounds

No textbook homework to accompany this quiz due to a lack of appropriate questions!

CHAPTER 3

COMPOSITION OF SUBSTANCES AND SOLUTIONS

3.1 Formula Mass and the Mole Concept

- For chemists, the amount of a substance we care about is not grams, but the quantity of atoms
- Analogy of hammers and feathers
- The mole is defined as the relationship between $AMUs$ and g

$$6.022 \times 10^{23} AMU = 1g \longrightarrow 1mol AMU = 1g \longrightarrow 1AMU = 1 \frac{g}{mol}$$
- Any time we would use the AMU we could replace it with g/mol , and this is called the *molar mass*
- Practice converting between grams and moles using atomic weight
- For compounds, we can find the molar mass, or formula mass by adding up the atomic weights of all the atoms in the formula
- We can count the literal number of atoms, but it is usually best to remain in units of moles
- Practice finding formula masses for compounds, and converting between grams and moles for compounds

3.2 Determining Empirical and Molecular Formulas

- Empirical analysis is a method to determine empirical formulas
 - In our text, the masses of C, H, and O are given directly
 - To get the formula, turn each mass into moles, and divide by the smallest mole number (Figure 3.11)
 - Actually, this technique involves a careful combustion reaction
 - The mass of C comes from the amount of CO_2 produced
 - The mass of H comes from the amount of H_2O produced (remember the 2!)
 - The mass of O comes from subtracting those from the original amount
 - Turn the masses to moles
 - Practice: 2.443g of unknown produce 5.368g of CO_2 and 2.930g of H_2O ($M = 60.096$, propanol)
 - Practice: 3.450g of unknown produce 5.057g of CO_2 and 2.070g of H_2O ($M = 180.16$, hexose)
- % Composition gives the % by mass of each element in a compound

- From masses, we can calculate the % mass composition easily
- If % mass is given, we can just assume 100.0g and turn the %s into g to do empirical analysis
- From the empirical or molecular formula, we can get % mass composition: $\%_X = \frac{nM_X}{M_{Formula}}$
- We can determine the molecular formula from the empirical formula if we know the molar mass
 - The molecular formula will be an integer multiple of the empirical formula (could be 1)
 - That formula is the ratio $n = \frac{M_{molar}}{M_{empirical}}$
 - Find the molecular formulas of the two examples above using their molar masses

Quiz 3.1 - Molar Mass

Homework 3.1

- 3: Calculating molar mass
- 17: Mass from moles
- 18: Moles from mass (sorry, no solution!)

Quiz 3.2 - Empirical Analysis

Homework 3.2

- 33: % composition from formulas
- 37: Empirical formulas from % composition
- 39: Molecular formula from % composition and molar mass

3.3 Molarity

- Molarity is the most common way of describing the *concentration* of a solute in a solution
 - The unit for molarity is M , which is unfortunate because that is also used for molar masses
 - Molarity is the moles of solute divided by the liters of solution $M = \frac{mol_{solute}}{L_{solution}}$
 - Technically, the solution volume could be more or less than the solvent volume before dissolving. When we make a solution, we always describe the process as “add enough water to meet the mark”
 - The molarity can now be used as a conversion between moles and volume
 - Practice: Find molar concentration of solution made by dissolving 13.5g NaCl in 100.0ml of solution

- Practice: How many moles of NaCl are contained in 35.62 *ml* of the solution?
- Practice: How many *ml* would be needed to provide $1.74 \times 10^{-3} \text{ mol}$ of NaCl?
- Solutions can be diluted or concentrated
 - The number of moles of solute is the same, just the volume of solution changes (by adding or removing solvent)
 - $C_1 V_1 = C_2 V_2$
 - Practice: Find the molar concentration when 36.43 *ml* of 1.5 *M* HCl are diluted to 100.0 *ml*?
 - Practice: How many *ml* of 0.125 *M* NaNO₃ should be used to make 50.0 *ml* of $2.4 \times 10^{-3} \text{ M}$?

3.4 Other Units for Solution Concentration

- There are other, less common units of concentration (including some not covered until chapter 11)
- Mass %, or % by mass
 - $\%m/m = \frac{m_A}{m_{total}} 100\%$
- Volume %, or % by volume
 - $\%v/v = \frac{V_A}{V_{solution}} 100\%$
- m/V %
 - $\%m/v = \frac{g_A}{V_{solution}} 100\%$
- ppm and ppb
 - $ppm = \frac{m_A}{m_{total}} 10^6 ppm$
 - $ppb = \frac{m_A}{m_{total}} 10^9 ppb$

Quiz 3.3 - Concentration

Homework 3.3

- 49: Mass from molarity and volume
- 53: Calculating molarity
- 61: Dilution
- 73: ppm
- 77: %m/m

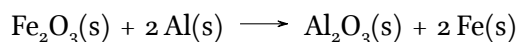
CHAPTER 4

STOICHIOMETRY OF CHEMICAL REACTIONS

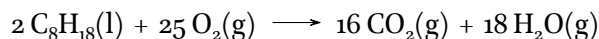
4.1 Writing and Balancing Chemical Equations

- A chemical equation relates the amounts of reactants and products in a chemical reaction
- Consider the reaction: $\text{CH}_4 + 2 \text{O}_2 \longrightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$ (Figure 4.2)
- Reactants on the left, and products are on the right
- The large numbers are called *stoichiometric coefficients*
- The subscript numbers are part of the chemical identity of the reactants and products
- The stoichiometric coefficients tell the right ratio to combine reactants, and the ratio of products produced (Figure 4.3)
- Balancing chemical reactions:
 - To balance a chemical equation, we will add the proper coefficients to give the same numbers and types of atoms on both sides of the reaction arrow
 - Start with any repeated atom moieties, such as polyatomic ions, and balance them as a unit rather than counting individual atoms
 - Next balance elements that appear in only one compound on each side
 - Finally, balance any remaining elements (O is often best left to balance last)
 - If necessary, multiply all coefficients by an integer

Practice: Balance the following chemical equation: $\text{Fe}_2\text{O}_3(\text{s}) + \text{Al}(\text{s}) \longrightarrow \text{Al}_2\text{O}_3(\text{s}) + \text{Fe}(\text{s})$



Practice: Balance the following chemical equation: $\text{C}_8\text{H}_{18}(\text{l}) + \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$



Practice: Balance the following chemical equation: $\text{Ag}_2\text{SO}_4(\text{aq}) + \text{NaCl}(\text{aq}) \longrightarrow \text{AgCl}(\text{s}) + \text{Na}_2\text{SO}_4(\text{aq})$



- Chemical reactions will sometimes include additional information
 - Phases (s), (l), (g), (aq)
 - Reaction conditions, such as temperature, or the presence of a catalyst (often written over the arrow)
 - Enthalpy of reaction or equilibrium constant (covered in later chapters)

- Reactions with ionic compounds can often be written differently and simplified
 - The normal equation is called a *molecular equation* (despite involving ionic compounds)

$$\text{CaCl}_2(\text{aq}) + 2 \text{AgNO}_3(\text{aq}) \longrightarrow \text{Ca}(\text{NO}_3)_2(\text{aq}) + 2 \text{AgCl}(\text{s})$$
 - The *complete ionic equation* will break *soluble* ionic compounds into the separate cations and anions
 - This is a more accurate representation of soluble ionic compounds, which are not actually *together* in any chemically meaningful way

$$\text{Ca}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq}) + 2 \text{Ag}^{+}(\text{aq}) + 2 \text{NO}_3^{-}(\text{aq}) \longrightarrow \text{Ca}^{2+}(\text{aq}) + 2 \text{NO}_3^{-}(\text{aq}) + 2 \text{AgCl}(\text{s})$$
 - The *net ionic equation* removes *spectator ions*, which didn't actually participate in any chemical change
 - My silly analogy about party-goers and the social interactions in a party environment

$$\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \longrightarrow \text{AgCl}(\text{s})$$

Quiz 4.1 - Balance Chemical Equations

Homework 4.1

- 5: Balancing equations
- 11: Net ionic equations

4.2 Classifying Chemical Reactions

- We can categorize chemical reactions in several ways, but for this class we will focus on the types of chemical changes caused by a reaction
- Precipitation reactions form one or more solid ionic product from all aqueous ionic reactants
 - First, predict the products by switching cations and anions
 - Next, predict the phase by using solubility rules (Table 4.1)
 - Identify and eliminate the spectator ions to get a net ionic equation
 - Some combinations of reactants will have no reaction, and some may have two solid products
 - More about precipitation reaction in CHEM 1220

Quiz 4.2 - Precipitation Reactions

Homework 4.2

- 28: Predicting products of reactions with ionic compounds

Resuming Section 4.2: Classifying Chemical Reactions

- Acid-Base reactions involve the transfer of a H^+ between reactants
 - Acids will react with water to produce hydronium ion
 - Bases will react with water to produce hydroxide ions
 - *Strong* acids/bases react with water to completion, while *weak* acids/bases only react partially (equilibrium)
 - Table 4.2 lists six common strong acids
 - Alternative definition: Acids are proton donors and bases are proton acceptors
 - A neutralization reaction produces water and a salt (ionic compound)
 - More about acid/base reactions in CHEM 1220
- Redox reactions involve the transfer of one or more electrons
 - Consider the reaction: $2 \text{Na(s)} + \text{Cl}_2\text{(g)} \longrightarrow 2 \text{NaCl(s)}$
 - OIL RIG (or LEO says GER) to remember which is oxidation, and which is reduction
 - Oxidizing agents and reductant agents. “Agent” describes the effect on the reaction partner
 - Sometimes the electron transfer is not as obvious: $\text{CH}_4 + 2 \text{O}_2 \longrightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$
 - We assign oxidation numbers to keep track of electrons
 - * Elements have ox. # of 0
 - * Monoatomic ions have ox. # equal to their charge
 - * Within compounds, O has ox. # = -2 , and H has ox # = $+1$
 - * Oxidation #s add up to the total overall charge (this rule subsumes the first 2)
 - Balancing redox reactions needs additional steps to account for the transferred electrons
 1. Split the reaction into half-reactions (need to find oxidation #s to do this)
 2. Balance all elements except H and O
 3. Add the electrons, based on the changes to oxidation #s
 4. Balance charge by adding H^+ in acid, or OH^- in base
 5. Balance H and O together by adding water
 6. Multiply half-reactions to balance their electrons, then add them together
 7. Make any cancellations of water, H^+ , or OH^-
 - You can balance redox reactions as a whole instead of as half-reactions if you prefer
 - Electrochemical series and predicting if redox reactions will be spontaneous or not (Not in the book!)
 - More about Redox reaction in ...CHEM 1220!

Quiz 4.3 - Redox Reactions

Homework 4.3

- 17: Assigning oxidation states
- 19: Classify acid/base and redox reactions
- 39: Balancing half-reactions

4.3 Reaction Stoichiometry

- Measured amounts should always be converted to moles in order to make comparisons to other chemical species in an equation
 - We often want to calculate the corresponding amount of different chemicals in a chemical reaction
 - Making direct comparisons will not work due to different molar masses, etc.
 - Figure 4.11 shows how to convert to moles from different measurements
 - My version of this figure ("Chemistry Land")
- Practice: $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \longrightarrow 2 \text{NH}_3(\text{g})$ Find other amounts for 0.75g of H_2

Quiz 4.4 - Stoichiometry**Homework 4.4**

- 49: Stoichiometry with solid reactants
- 57: Stoichiometry with aqueous reactions

4.4 Reaction Yields

- Often one or more reactants are provided in excess, with only one *limiting reactant*
 - Pick a product, and calculate how much product would be produced for each reactant
 - The lowest amount is the theoretical yield. The other amounts can be discarded
 - The reactant which lead to the lowest amount is the limiting reactant
 - Calculate the rest of the corresponding amounts starting from the limiting reactant
 - Double-check your work using the conservation of mass
- Practice: $2 \text{C}_4\text{H}_{10} + 13 \text{O}_2 \longrightarrow 8 \text{CO}_2 + 10 \text{H}_2\text{O}$ with 5.00g C_4H_{10} and 9.00g O_2
- Reactions rarely go perfectly, and the actual amount of product will be different from the theoretical yield

- $\%yield = \frac{actual\ yield}{theoretical\ yield} \times 100\%$
- Yields below 100% can mean that the reaction didn't go to completion, or some product was lost during a purification stage
- Yields above 100% usually indicate that some contaminants (including perhaps excess reactant) are mixed in with the product

Quiz 4.5 - Limiting Reactants

Homework 4.5

- 61: Limiting Reactant Problem
- 63: Percent Yield

4.5 Quantitative Chemical Analysis

- Quantitative analysis is the family of techniques which determine the amount of substance in a sample
- Titrations find the concentration of an aqueous *analyte*
 - React the unknown with a co-reactant of known concentration by adding titrant with a buret
 - The chemical environment will change (*pH*, redox potential, etc.) once the reaction reaches the equivalence point
 - Probes or color indicators can show when to stop the titration. This is called the end-point, and is ideally very close to the equivalence point
 - $\frac{C_T V_T}{\nu_T} = \frac{C_A V_A}{\nu_A}$
- Gravimetric analysis uses a chemical reaction to change the phase of the analyte so it can be separated and weighed
 - Your % copper lab found the amount of copper by reducing it to a solid
 - Your hydrates lab found the amount of water by weighing it before and after the baking step
- Combustion analysis - We actually covered this earlier with % composition of compounds

Quiz 4.6 - Titrations

Homework 4.6

- 79: Acid/base Titration
- 81: Precipitation Titration

CHAPTER 5

THERMOCHEMISTRY

5.1 Energy Basics

- Thermochemistry is the study of heat and energy changes in chemical reactions
- It also includes topics like entropy and spontaneity
- Energy can come in two forms, kinetic and potential (Figure 5.3)
 - Kinetic energy is the energy of motion: $KE = \frac{1}{2}mv^2$
 - For chemists, the kinetic energy that matters is the movement of individual atoms, molecules, etc., manifested as temperature (Figure 5.4)
 - Heat will always flow from colder objects to warmer objects (Figure 5.6)
 - Increases in temperature almost always cause objects to increase in volume (Figure 5.5)
 - Potential energy is stored energy: Gravitational, electrostatic, chemical, etc.
 - For chemists, the potential energy that matters is the energy of chemical bonds and intermolecular forces (Figure 5.2)
- Energy has several common units
 - The SI unit is the Joule: $1 J = 1 kg \frac{m^2}{s^2}$
 - The calorie: $1 cal = 4.184 J$
 - The Calorie (kcalorie): $1 Cal = 1000 cal = 4184 J$
- When we talk about changes and transfers of energy, we need to carefully define our system
 - The *system* is the part of the universe where the reaction occurs, such as a beaker or chamber with reactants
 - The *surroundings* is the rest of the universe
 - Open systems can exchange both heat and matter with the surroundings (an open beaker)
 - Closed systems can exchange heat, but not matter with the surroundings (a closed chamber)
 - Isolated systems cannot exchange either heat or matter with the surroundings (a closed, insulated flask)
- Heat from chemical and physical changes is usually associated with temperature changes or phase changes (Figure 5.7)
 - Positive heat is observed by an “upward” phase change or a cold temperature

- Holding ice in your hand will melt the ice, and make your hand cold
 - Negative heat is observed by a “downward” phase change or a hot temperature
 - Burning wood in a fire feels warm because of the negative system heat
 - Heat and temperature change are related by several equations (Note that we are always measuring the temperature of the *surroundings*)
 - For a complete system, $q = C\Delta T$ where C is the *heat capacity*
 - For a pure substance, $q = mc_S\Delta T$ where c_S is the *specific heat* of the substance and m is the mass
 - Table 5.1 shows c_S for many common substances
- Practice: Find how much heat is required to raise the temperature of 15.0g of iron by 32°C
- Practice: Find the temperature change when 42.5J of heat are added to 0.374g of aluminum

Quiz 5.1 - Energy

Homework 5.1

○ :

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

- Many physical properties of solids, liquids, and gases can be explained by the strength of attractive forces between particles (Figure 10.5)
- Phase changes happen due to the interplay between kinetic energy and intermolecular forces (Figure 10.2)
- Pressure can also play a role in phase changes, as discussed later
- These *intermolecular forces* come in different varieties
 - Dispersion Forces Non-polar molecules, impacted by polarizability, molecular weight, and surface area
 - * Dominant in non-polar molecules
 - * Created by induced dipoles (Figure 10.6)
 - * Impacted by polarizability (Table 10.1)
 - * Impacted by molecular weight (hydrocarbons from methane to wax)
 - * Impacted by molecule shape (Figure 10.7 compares the boiling points of pentane isomers)
 - Dipole-Dipole Forces
 - * Dominant in polar molecules
 - * Results from attraction between permanent dipoles (Figure 10.9)
 - Hydrogen Bonding
 - * Dominant only in molecules capable of hydrogen bonding
 - * Must contain a hydrogen-donor atom (H attached to N, O, or F)
 - * Must contain a hydrogen-acceptor atom (lone pair of electrons attached to N, O, or F)
 - * Hydrogen bonds are more than just particularly strong dipole-dipole forces. They have strong directionality according to VSEPR
 - * Figures 10.10, 10.14, and other figures on the Internet show water, DNA, and proteins all organized by hydrogen bonds
 - * Figures 10.11 and 10.12 illustrate how much hydrogen bonds exceed dipole-dipole forces in strength

ERRATA