# 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
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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

**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 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

**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 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: $1\,J \equiv kg\,m^2/s^2$

## 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)

**Quiz 1.3 - Significant Figures**

**Homework 1.3**

- 45: Scientific Notation

- 49: Counting Significant Figures

- 53: Significnat 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 quantites

- 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

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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 concetrated 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 techniqes 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 Å$= 10^{-10} m$

- ○ 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 charge 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: $^{13}_{6}C^{2+}$ (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 chemisry
    - · 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 · 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 of 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!**

# 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 $g$

  $$6.022 \times 10^{23} AMU = 1g \longrightarrow 1 mol\, AMU = 1g \longrightarrow 1 AMU = 1\frac{g}{mol}$$

- 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 emipirical analysis
- From the empirical or molecular formula, we can get % mass composition: $\%_X = \dfrac{nM_X}{M_{Formula}}$

○ 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 $n = \dfrac{M_{molar}}{M_{empirical}}$
- 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 $M = \dfrac{mol_{solute}}{L_{solution}}$
- 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 $35.62ml$ of the solution?
  - · Practice: How many $ml$ would be needed to provide $1.74 \times 10^{-3}mol$ 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)
  - · $C_1V_1 = C_2V_2$
  - · Practice: Find the molar concentration when $36.43ml$ of $1.5M$ HCl are diluted to $100.0ml$?
  - · Practice: How many $ml$ of $0.125M$ NaNO$_3$ should be used to make $50.0ml$ of $2.4 \times 10^{-3}M$?

## 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

  - · $\%m/m = \dfrac{m_A}{m_{total}}100\%$

- ○ Volume %, or % by volume

  - · $\%v/v = \dfrac{V_A}{V_{solution}}100\%$

- ○ m/V %

  - · $\%m/v = \dfrac{g_A}{V_{solution}}100\%$

- ○ ppm and ppb

  - · $ppm = \dfrac{m_A}{m_{total}}10^6ppm$
  - · $ppb = \dfrac{m_A}{m_{total}}10^9ppb$

**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: $CH_4$ + 2 $O_2$ $\longrightarrow$ $CO_2$ + 2 $H_2O$ (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: $Fe_2O_3(s)$ + $Al(s)$ $\longrightarrow$ $Al_2O_3(s)$ + $Fe(s)$

   $Fe_2O_3(s)$ + 2 $Al(s)$ $\longrightarrow$ $Al_2O_3(s)$ + 2 $Fe(s)$
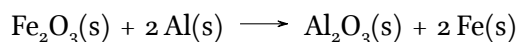
   Practice: Balance the following chemical equation: $C_8H_{18}(l)$ + $O_2(g)$ $\longrightarrow$ $CO_2(g)$ + $H_2O(g)$

   2 $C_8H_{18}(l)$ + 25 $O_2(g)$ $\longrightarrow$ 16 $CO_2(g)$ + 18 $H_2O(g)$
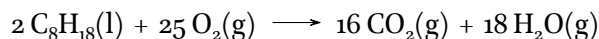
   Practice: Balance the following chemical equation: $Ag_2SO_4(aq)$ + $NaCl(aq)$ $\longrightarrow$ $AgCl(s)$ + $Na_2SO_4(aq)$

   $Ag_2SO_4(aq)$ + 2 $NaCl(aq)$ $\longrightarrow$ 2 $AgCl(s)$ + $Na_2SO_4(aq)$

○ 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)
      $$CaCl_2(aq) + 2\,AgNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + 2\,AgCl(s)$$

    - · 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
      $$Ca^{2+}(aq) + 2\,Cl^-(aq) + 2\,Ag^+(aq) + 2\,NO_3^-(aq) \longrightarrow Ca^{2+}(aq) + 2\,NO_3^-(aq) + 2\,AgCl(s)$$

    - · 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
      $$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

**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

    - · Consdier the reaction: $2\,Na(s) + Cl_2(g) \longrightarrow 2\,NaCl(s)$
    - · OIL RIG (or LEO says GER) to remember which is oxidation, and which is reduction
    - · Oxidizing agents and reducint agents. "Agent" describes the effect on the reaction partner
    - · Sometimes the electron transfer is not as obvious: $CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$
    - · 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: $N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$ 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: $2\,C_4H_{10} + 13\,O_2 \longrightarrow 8\,CO_2 + 10\,H_2O$ 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

- $\%yield = \dfrac{actual\ yield}{theoretical\ yield} \times 100\%$

- 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
    - $\dfrac{C_T V_T}{\nu_T} = \dfrac{C_A V_A}{\nu_A}$

- 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

# 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: $KE = \dfrac{1}{2}mv^2$

    - 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: Gravitaional, 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: $1\ J = 1\ kg\dfrac{m^2}{s^2}$

    - 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)

    - These examples all refer to heat from the *system's* perspective

       · 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*)

       · When heat is added to a system, it will either undergo a phase change, or heat up

       · The specific heat gives how much heat is required to warm a given substance

       · 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: Heat capacity of samples

○ 9: $q = mc_S\Delta T$

○ 11: $q = mc_S\Delta T$, but solve for $c_S$

## 5.2   Calorimetry

○ Whenever heat is transferred, the total energy of the universe remains constant

○ For a heat transfer in an isolated system, $q_1 = -q_2$

○ Consider a block of hot metal placed in a beaker of room temperature water. Heat will flow from the block into the water until the two temperature are equal to each other

○ $q_1 = -q_2$ becomes $m_1 c_1 (T_f - T_{i,1}) = -m_2 c_2 (T_f - T_{i,2})$
    Practice: A $10.0\ g$ block of iron is heated to $93.5\ °C$ and placed in $25\ ml$ of $23.0\ °C$ water.
        What is the final temperature? $(25.9\ °C)$

○ We can also measure the heat transfer associated with a chemical reaction, called *calorimetry*

○ Exothermic processes release heat into the surroundings (warm to the touch)

○ Endothermic processes absorb heat from the surroundings (cool to the touch)

- ○ Figure 5.11 shows how a calorimeter would work for endothermic and exothermic processes

- ○ Figures 5.12 and 5.13 show a constant pressure (coffee-cup) calorimeter

    - · The reaction is carried out in aqueous solution
    - · $q_{rxn}$ is the heat released or absorbed by the reaction
    - · The heat of reaction is exchanged with the solution: $q_{rxn} = -q_{soln} = -mc\Delta T$
    - · $m$ is the solution mass, which will include the water and any solutes
    - · $c$ is the solution specific heat, but this is simplified by assuming $c_{soln} = c_{water} = 4.184\frac{J}{g\,^\circ C}$
    - · $\Delta H = \dfrac{q_{rxn}}{n_{rxn}} = \dfrac{-mc\Delta T}{n_{rxn}}$ where $n_{rxn}$ is the moles of reaction: $n_{rxn} = \left(\dfrac{n_A}{\nu_A}\right)$
    - · Demo – NaOH enthalpy of solvation $\left(10\ g,\ 100\ ml,\ 44.2\ \frac{kJ}{mol}\right)$

- ○ Figure 5.17 shows a constant volume (bomb) calorimeter

    - · The reaction is carried out in a chamber charged with high pressure $O_2$
    - · $q_{rxn}$ is exchanged with the whole bomb-calorimeter apparatus
    - · The calorimeter is calibrated to give a *heat capacity* $(C_{cal})$ with units $\frac{J}{^\circ C}$
    - · $q_{rxn} = -q_{cal} = -C_{cal}\Delta T$
    - · For constant volume, we measure $U$ instead of $H$ because $w = 0$
    - · $\Delta U = \dfrac{-C_{cal}\Delta T}{n_{rxn}}$

**Quiz 5.2 - Calorimetry**

**Homework 5.2**

- ○ 19: Reaching thermal equilibrium
- ○ 25: Coffee cup calorimetry
- ○ 31: Bomb calorimetry

## 5.3   Enthalpy

- ○ The sum of all types of energy in a system is the *internal energy*, $U$
- ○ Any change in the internal energy must come from heat or work: $\Delta U = q + w$
- ○ Work is defined as $w = f \cdot d$ or $w = -P \cdot \Delta V$
    - · Positive work is when the system volume decreases
    - · Negative work is when the system volume increases, or a force moves part of the surroundings
    - · This $PV$ work is actually a bit problematic when trying to keep track of energy

* $T$ is easy to measure with a thermometer, but both $P$ and (especially) $V$ are more difficult to measure
    * Most of our work as chemists is done at constant pressure (open flask or in a balloon)
    * Under constant pressure, we can use *Enthalpy* ($H$) instead of internal energy ($U$)
    * Technically, $H = U + PV$
    * While $U = q + w$, $H = q$ under constant pressure conditions (derivation in the book)
    * So, we only need to worry about heat when we deal with $H$

- Heat is usually associated with temperature changes or phase changes, as discussed previously

- First Law of Thermodynamics: The energy of the universe is constant

- Enthalpy in chemical reactions:

    · A balanced chemical reaction may also include an enthalpy of reaction $\Delta H$

    · This tells how much heat is produced or consumed with one mole of reaction

    · $\Delta H$ can be a conversion factor between heat and amounts of reactants or products
      Practice: Consider the reaction $N_2(g)$ + $3 H_2(g) \longrightarrow 2 NH_3(g)$ $\Delta H = -92\frac{kJ}{mol}$
                If $2.25$ $g$ of $H_2$ are consumed in the above reaction, how much heat is released? $(34.2\ kJ)$
                    If $54.6\ kJ$ of heat are released, how many $g$ of $NH_3$ will be produced? $(20.2\ g)$

**Quiz 5.3 - Enthalpy**

**Homework 5.3**

- 41: Enthalpy of reaction from calorimetry

- 45: Heath from enthalpy of reaction

- 47: How much reactant from target heat output

**Resume Section 5.3 – Enthalpy**

- The "Standard State" is notated by $^\circ$ and is at $1M$ concentration and $1atm$ (or $1bar$), and technically doesn't include a temperature but values are often tabulated at $25°C$

- Standard Enthalpies of Combusion ($\Delta H_C^\circ$):

    · Standard enthalpy of combustion is the energy released when something reacts with oxygen

    · Table 5.2 Lists the enthalpies of combustion for many combustible substances

    · These values are not synonymous with energy density, due to molar mass and density complicating the values

- Because enthalpy is a state function (Figure 5.20), we can calculate values of $\Delta H$ without measuring them

- ○ Hess's Law: Any alternate path with the same starting and ending states will have the same overall $\Delta H$

  - · Drawing an energy level diagram can help to illustrate Hess's law (Figure 5.24 as an example)
  - · Find $\Delta H$ for this reaction: $C_{\text{diamond}} + O_2(g) \longrightarrow CO_2(g)$  $(\Delta H = -395.4 \, ^{kJ}/_{mol})$

    $C_{\text{diamond}} \longrightarrow C_{\text{graphite}}$  $\Delta H = -1.9 \, ^{kJ}/_{mol}$

    $C_{\text{graphite}} + O_2(g) \longrightarrow CO_2(g)$  $\Delta H = -393.5 \, ^{kJ}/_{mol}$

  - · The reverse of a reaction gives $-\Delta H$
  - · Find $\Delta H$ for this reaction: $C(s) + \frac{1}{2} O_2(g) \longrightarrow CO(g)$  $(\Delta H = -110 \, ^{kJ}/_{mol})$

    $C(s) + O_2(g) \longrightarrow CO_2(g)$  $\Delta H = -393 \, ^{kJ}/_{mol}$

    $CO(g) + \frac{1}{2} O_2(g) \longrightarrow CO_2(g)$  $\Delta H = -283 \, ^{kJ}/_{mol}$

  - · Double the reaction gives double the $\Delta H$
  - · Consider trying to find $\Delta H$ for the reaction below:

    $(\star)\, C_2H_5OH(l) + 2\, O_2(g) \longrightarrow 2\, CO(g) + 3\, H_2O(l)$  $\Delta H = ?$

  - · Find an alternate path using these reactions with known $\Delta H$:

    $(A)\, C_2H_5OH(l) + 3\, O_2(g) \longrightarrow 2\, CO_2(g) + 3\, H_2O(l)$  $\Delta H = -1367 \, ^{kJ}/_{mol}$

    $(B)\, CO(g) + \frac{1}{2} O_2(g) \longrightarrow CO_2(g)$  $\Delta H = -283 \, ^{kJ}/_{mol}$

  - · The enthalpy of the first reaction can be found from the enthalpies of the other two
  - · $\Delta H_\star = \Delta H_A - 2\Delta H_B = -801 \dfrac{kJ}{mol}$  (Draw the energy level diagram)

    Practice: Find the enthalpy of reaction $(\star)$ using reactions $(A)$, $(B)$, and $(C)$

    $(\star)\, CS_2(l) + 3\, O_2(g) \longrightarrow CO_2(g) + 2\, SO_2(g)$  $(\Delta H = -1075.0)$

    $(A)\, C(s) + O_2(g) \longrightarrow CO_2(g)$  $\Delta H = -393.5 \, ^{kJ}/_{mol}$

    $(B)\, S(s) + O_2(g) \longrightarrow SO_2(g)$  $\Delta H = -296.8 \, ^{kJ}/_{mol}$

    $(C)\, C(s) + 2\, S(s) \longrightarrow CS_2(l)$  $\Delta H = 87.9 \, ^{kJ}/_{mol}$

**Quiz 5.4 - Hess's Law**

**Homework 5.4**

- ○ 59: Hess's Law and enthalpy of reaction
- ○ 63: Hess's Law and enthalpy of reaction

**Resume Section 5.3 – Enthalpy**

- ○ Standard Enthalpies of Formation $\Delta H_f^\circ$

  - · To apply Hess's law to arbitrary reactions, you would need to devise an alternate path from an encyclopedia of known reactions – this would be *very* inconvenient

- Instead of using random reactions from one state to another, it is useful to devise a *standard state* for each element
- The standard state is the most stable form of that element e.g. for O, it is $O_2(g)$, not $O_2(l)$ or $O_3(g)$
- Each compound will have a *standard formation reaction* which forms it from its elements in their standard state

    For water, that's $H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(l)$     (This is one time a $\frac{1}{2}$ coefficient is acceptable)
- The enthalpy for this reaction is called the compound's *Standard Enthalpy of Formation* $\left( \Delta H_f^\circ \right)$
- Elements in their standard state have $\Delta H_f^\circ = 0$
- Any reaction can be framed as a combination of standard formation reactions

    * First, the reactants are broken down into their elements (the *reverse* of formation reactions)
    * Then, the elements are reassembled into the products (formation reactions)
    * The energy level diagram for any reaction is the same: reactants $\rightarrow$ elements $\rightarrow$ products
    * This pathway doesn't need to be *practical*, it is enough to be theoretically *possible*
    * $\Delta H_{rxn} = \sum\limits_{products} \nu \cdot \Delta H_f^\circ - \sum\limits_{reactants} \nu \cdot \Delta H_f^\circ$
    * This formula is general, for any reaction at all
    * Instead of an encyclopedia of thousands of reactions, we only need a table of $\Delta H_f^\circ$ values
    * Appendix G includes a large number of thermodynamic values like $\Delta H_f^\circ$

    Practice: Find $\Delta H_{rxn}$ for $P_4O_{10}(s) + 6 H_2O(l) \longrightarrow 4 H_3PO_4(aq)$    $\left( \Delta H_{rxn} = 454.6 \dfrac{kJ}{mol} \right)$

    Find $\Delta H_{rxn}$ for $C_3H_8(g) + 5 O_2(g) \longrightarrow 3 CO_2(g) + 4 H_2O(g)$    $\left( \Delta H_{rxn} = -2043.9 \dfrac{kJ}{mol} \right)$

**Quiz 5.5 - Enthalpies of Formation**

**Homework 5.5**

- ○ 67: Calculate $\Delta H_f^\circ$ from reaction enthalpies
- ○ 69: Reaction enthalpies from enthalpies of formation

CHAPTER 6

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# ELECTRONIC STRUCTURE AND PERIODIC PROPERTIES OF ELEMENTS

## 6.1  Electromagnetic Energy

○ Light is an electromagnetic wave, which can be thought of like a sound wave or a wave on a lake (Figure 6.2)

○ Light has a wavelength, frequency, and speed according to the equation: $\nu\lambda = c$

○ The speed of light is a constant, $2.998 \times 10^8 \, \dfrac{m}{s}$

○ How far does light travel in $5.00 \, ms$? ($1.50 \, km$)

○ The electromagnetic spectrum is more than just visible light (Figure 6.3)

  · Higher frequencies (shorter wavelengths) are UV light, X-rays, and gamma rays
  · Lower frequencies (longer wavelengths) are infrared light, microwaves, and radio waves
  · TV-remotes are flashy lights, and radio towers are flashlights-on-a-stick
  · There are important technical differences in how we can use these different kinds of light, but they are fundamentally the same thing (an alternating electromagnetic wave)
  · Find the wavelength of your favorite radio station ($MHz$ is a frequency of $10^6 \, s^{-1}$)

○ Light-matter interactions were central to the discovery of modern physics

○ The photoelectric effect was an important matter/light interaction (Figure 6.11)

  · Sometimes light falling on a metal will eject an electron – this is the photoelectric effect
  · The kinetic energy of the ejected electron can be measured
  · The energy depended on the wavelength of light – bluer light ejected electrons at faster velocities
  · There was a threshold where electron ejection stopped, and redder light would have no effect
  · This was surprising, because light intensity had *no* effect on the photoelectron energy
  · Dim blue light would eject fast electrons, bright red light would have no effect
  · This was eventually explained by the idea that light carries energy in small discrete packets
  · These packets of energy are called *photons* and the energy they carry depends on the wavelength
  · The photoelectric effect could be described by the equation: $KE = h\nu - \phi$
  · Here, $\nu$ was a new constant, called Planck's constant, and $\phi$ was the metal's work function
  · The equation for the energy of light was: $E = h\nu$ where $h$ is Planck's constant

- ○ Light can also be *absorbed* or *emitted* by matter

  - · Light is emitted by gases like Ne, Ar, or Na when high voltage passes through it
  - · Light can also be absorbed by gases and other materials
  - · Each substance showed a unique fingerprint of wavelengths of light emitted or absorbed (Figure 6.13)
  - · He was first identified by its absorbance spectrum in sunlight
  - · The unique spectra arise from the particular energy levels of a substance
  - · These spectra showed how matter can only store or release energy in certain, constrained amounts (or *quanta*)
  - · i.e. in addition to the energy of light, the energy of matter was *quantized* as well

- ○ Ultraviolet Catastrophe (Figures 6.10 and 6.9)

- ○ Wave/Particle duality (not in the textbook)

  - · These experiments show that light behaves like both a *wave* (interference) and a *particle* (quantization)
  - · It is wrong to say light is either of those things – rather, it is a new thing with similarities to both (rhinosceros vs dragon + unicorn)
  - · On very small scales, matter behaves like both a wave and a particle as well!
  - · Electrons, in particular, are strongly wave-like, with a characteristic wavelength
  - · The IBM quantum corral image dramatically showed the real physicality of electron waves
  - · This wave-like nature of electrons is important for understanding modern models of atomic structure

- ○ All of these phenomena can be explained in terms of standing waves (Figures 6.7 and 6.8)

## 6.2 The Bohr Model

- ○ The Rydberg Equation: (Figure 6.14)

  - · School teacher Rydberg recognized a pattern in the wavelengths of light in the H spectrum
  - · $\frac{1}{\lambda} = 1.097 \times 10^7 m^{-1} \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$
  - · His equation can be re-written in terms of energy
  - · $E = 2.179 \times 10^{-18} \, J \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$

- ○ The Bohr model of the atom: (Figure 6.15)

  - · There was no real explanation for *why* the absorption and emission spectra of different elements showed different discreet energies

- Niels Bohr proposed that electrons orbit around the nucleus only at fixed distances
- Absorption is when an electron shifts to a higher orbit, using a photon's energy
- Emission is when an electron shifts to a lower orbit, releasing energy as a photon
- The lowest energy state is the *ground state*, all others are *excited states*
- The discrete orbits represent states where the circumference of the orbit is equal to a number of wavelengths for the electron (Figure 6.17)
- $E_n = -2.179 \times 10^{-18} \, J \left( \dfrac{1}{n^2} \right)$

Practice: What is the wavelength of the $2 \leftarrow 4$ transition in the H spectrum? $(486.1 \, nm)$

**Quiz 6.1 - Light and the Bohr Model**

**Homework 6.1**

- 3: Frequency to wavelength

- 5: Wavelength to energy

- 23: Energy of a Bohr state

- 27: Energy of a Bohr transition

## 6.3  Development of Quantum Theory

- Quantum mechanics continued to develop

- Heissenberg Uncertainty Principle: We cannot simultaneously measure the position and velocity of an electron (or any other quantum mechanical particle)

- This means that we generally speak of where an electron is *most probable* to be found, rather than where it *actually is*

- The Schrödinger wave equation describes matter starting from a wave-like perspective

    - DeBroglie began from the equation for the momentum of light: $p = \frac{h}{\lambda}$
    - Light has no mass, but replacing $p$ with $mv$ give us: $mv = \frac{h}{\lambda}$ and rearranging gives: $\lambda = \frac{h}{mv}$
    - This equation describes how anything with momentum has a characteristic wavelength
    - Figure 6.18 illustrates one way electrons behave like waves
    - Schrödinger went even further, adapting Maxwell's wave equations for matter with momentum (required some postulates)
    - The Schrödinger equation gives mathematical functions which describe the electron probability distribution
    - Each solution is called an *orbital*, like the orbits of the Bohr model but 3-dimensional

- ○ Orbitals are organized into *shells* and *subshells*

    - · Subshells are groups of orbitals with similar shapes and the same energy
    - · Subshells are named $s$, $p$, $d$, and $f$
    - · An $s$ subshell has only one orbital (2 $e$s), $p$ has 3 (6 $e$s), $d$ has 5 (10 $e$s), and $f$ has 7 (14 $e$s)
    - · Subshells are grouped into shells, which are indicated by numbers (1, 2, 3, etc.)
    - · These numbers are the numbers in the Rydberg equation, and are the principle energy levels
    - · The first shell only has an $s$ subshell, and each shell beyond that adds one type

    Practice: How many electrons can be placed in a $p$ subshell? (6)

    How many electron can be placed in the $3^{rd}$ shell? (18)

    Which of the following subshells does *not* exist? $2s$, $3f$, $3p$, $5d$ $(3f)$

- ○ Subshells each have orbitals with different *shapes* (Figure 6.21)

    - · Because of the Heissenberg uncertainty principle, we describe region where an electron is likely to be found
    - · These regions have shapes based on the mathematical functions which form them
    - · $s$ orbitals are spherical, $p$ orbitals are dumbells, $d$ orbitals are clover-leafed

- ○ Quantum Numbers

    - · Remember that orbitals are actually mathematical functions
    - · Certain parts of those functions depend on integer numbers (like $n$ in the Rydberg equation)
    - · These integer numbers are called *quantum numbers*
    - · Quantum numbers can be thought as an "address" for each electron (Street, Buildling, Unit, Name)
        - ∗ $n$ – Principal quantum number (1, 2, ...) gives orbital shell, energy, and size
        - ∗ $l$ – Angular momentum quantum number $(0, 1, ..., n-1)$ gives orbital subshell and shape
        - ∗ $l$ is why not all shells have all orbital types
        - ∗ $m_l$ – Magnetic quantum number $(-l, ..., l)$ gives orbital within a subshell (different spatial orientations)
        - ∗ $m_l$ is why subshells have different numbers of orbitals
        - ∗ $m_s$ – Spin quantum number $\left(\pm\frac{1}{2}\right)$ gives "spin" of the electron (up- or down- arrow, Figure 6.23)
    - · Each electron in an atom/ion must have a *unique* set of quantum numbers – Pauli Exclusion Principle
    - · You should be able to point to the right electron given a set of quantum numbers, or give the 4 quantum numbers for an indicated electron in an energy level diagram

        Practice: Give numbers or indicate electrons on an energy level diagram

**Quiz 6.2 - Orbitals and Quantum Numbers**

**Homework 6.2**

- 33: Properties associated with quantum numbers

- 35: Identifying the subshell from quantum numbers

- 41: Orbital shapes and other properties

## 6.4 Electronic Structure of Atoms (Electron Configurations)

- Electronic energy level diagrams

  - The orbitals within each subshell are precisely equal in energy (degenerate)
  - The subshells themselves differ in energy (their order will be explained shortly)
  - Draw the subshells with one line for each orbital
  - Each orbital can hold two electrons, drawn as up- and down-arrows
  - Find the total number of electrons for the element or ion
  - Aufbau Principle – Fill up the orbitals with electrons from the bottom-up
  - Hund's Rule – Fill a subshell with one electron in each orbital before pairing them up (like roommates in an apartment)
  - This is the ground-state configuration of the element

- We can use the periodic table as a cheat-sheet to the order of the subshells and electron configurations (Figure 6.27)

  - The P. T. is actually quite long – Lanthanides and Actinides have been cut and pasted
  - Each region of the P. T. represents a different subshell
  - The rows represent different shells
  - The number of elements in each block is the number of electrons each shell can hold
  - The $d$-block and $f$-block trail the row number by 1 and 2 ($d(-1)$ and $f(-2)$)
  - The order of the subshells is found by simply following the elements and noting in which block they reside

- The arrangements of the electrons can be written as an electron configuration

- Write the subshells, with their number of electrons as a superscript

  Practice: Write the electronic configurations for O, Zr, and Bi

- Especially for large elements like Pb, these configurations are very unwieldy

- We can shorten them by referencing the configuration of the *preceding* noble gas

- The electrons which make up this noble gas configuration are buried inside the atom, and called *core* electrons

- For Bi, we get $[\text{Xe}]\, 6s^2 4f^{14} 5d^{10} 6p^3$

- The outermost electrons (The ones we write) are called *valence* electrons, and are the ones involved in bonding and ion formation

- A few transition metals have anomalous configurations (memorize only Cr and Cu)

- The Lanthanum and Actinium boundary also shows some anomalies (don't memorize them)

- Figure 6.29 gives the valence configuration of every element's ground state

- We can write the configurations of ions as well

  - For most, simply add or remove electrons according to the normal pattern
  - Transition metals will lose the outermost $s$ electrons before they lose any $d$ electrons (This is why so many transition metals have a stable $2+$ ion)
  - Ions and atoms with identical configurations are called *isoelectronic* to each other

  Practice: List several stable ions which are isoelectronic with Ar

## 6.5   Periodic Variations in Element Properties

CHAPTER 7

CHEMICAL BONDING AND MOLECULAR GEOMETRY

CHAPTER 8

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# ADVANCED THEORIES OF COVALENT BONDING

CHAPTER 9

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# GASES

# 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

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