

Check notes

<http://drvanderveen.com/apchemistry.htm>

<http://drvanderveen.com/AP%20Polyatomic%20Ions.pdf>

TABLE OF CONTENTS (turn on headings)

[Chapter 1](#)

[Chapter 2](#)

[Chapter 3](#)

[Chapter 4](#)

[Chapter 5](#)

[Chapter 6](#)

[Chapter 7](#)

[Chapter 8](#)

Chapter 1

1.1 The Study of Chemistry

- a. Matter
 - i. Physical material of the universe
 - ii. Mass and space
- b. Property
 - i. Characteristic allowing us to distinguish between matter
- c. Elements
 - i. Very basic and elementary
- d. Atoms
 - i. Infinitesimally small building blocks of matter
- e. Molecules
 - i. Are formed by atoms when they bond in different shapes
- f. Macroscopic
 - i. Ordinary sized objects (macro=large)
- g. Submicroscopic
 - i. Atoms and molecules

1.2 Classifications of matter

- a. States of matter
 - i. Gas
 - 1. No fixed volume or shape
 - 2. Conforms to volume and shape of its container
 - 3. Can be compressed/expanded
 - 4. Particles move fast, colliding with each other
 - ii. Liquid
 - 1. Distinct volume independent of its container but has no specific shape
 - 2. Assumes the shape of the portion of the container it occupies
 - 3. Can't be compressed
 - 4. Particles are packed more closely together (can still move)
 - iii. Solid
 - 1. Definite shape and volume
 - 2. Can't be compressed
 - 3. Particles are held closely together in definite arrangements
- b. Pure substances
 - i. Matter that has distinct properties and a composition that doesn't vary from sample to sample
 - ii. Elements
 - 1. Cannot be decomposed into simpler substances
 - iii. Compounds
 - 1. Composed of two or more elements (water)

- iv. Mixtures
 - 1. Combination of 2 or more substances in which each substance retains its own chemical identity
- c. Elements
 - i. Varied abundances in the crust and in the body
 - ii. They have chemical symbols in the periodic table
- d. Compounds
 - i. Elements can interact with each other to form compounds
 - 1. Ex:
 - a. H_2 burns in O_2 gas
 - ii. Law of constant composition
 - 1. Elemental composition of a pure compound is always the same
- e. Mixtures
 - i. Pure substances have fixed compositions, but compositions of mixtures vary
 - ii. Heterogeneous
 - 1. Vary throughout
 - iii. Homogeneous
 - 1. Uniform throughout
 - 2. Also called Solutions

1.3 Properties of Matter

- a. Physical properties
 - i. Measured without changing the identity and composition of the substance
 - ii. Color, odor, density, melting, boiling, hardness
- b. Chemical properties
 - i. How a substance may react to form other substances
 - ii. Flammability
- c. Intensive properties
 - i. temperature, melting point, density
 - ii. Do not depend on the amount of sample measured
- d. Extensive properties
 - i. Mass, volume
 - ii. Amount of substance present
- e. Physical and Chemical changes
 - i. Physical change
 - 1. Substance changes physical appearance, but not composition
 - a. Evaporation of water
 - b. Changes of state
 - ii. Chemical changes
 - 1. Substance is transformed into a chemically different substance
 - a. Burn something
- f. Separation of mixtures
 - i. Filtration
 - 1. Put a solid and liquid in a porous item (filter paper or something) and the liquid passes through while the solid does not
 - ii. Distillation
 - 1. Ex: boil a bunch of water until it evaporates, and then you get its components
 - iii. Chromatography

1. Ex: separating ink
 - a. Put some paper with ink in water
 - b. Water moves up the paper
 - c. Water moves past ink spot, which spread the ink at different rates
 - d. Water separates ink into different components

1.4 units of measurement

- a. Quantitative
 - i. Associated with numbers
 - ii. Use SI units
- b. Length and mass
 - i. Length (m)
 - ii. Mass (kg)
- c. Temperature (K)
 - i. $C + 273.15$
 - ii. $F = 9/5(C) + 32$
 - iii. $C = 5/9(F - 32)$
- d. Volume (m^3)
 - i. Also cm^3 for smaller measurements
- e. Density
 - i. $mass/volume = density$
 - ii. g/cm^3
 - iii. g/mL

1.5 Uncertainty in Measurement

- a. Exact numbers (values known exactly)
 - i. $1kg = 1000g$
- b. Inexact numbers (values that have uncertainties)
 - i. Any measurement ever done
- c. Precision and Accuracy
 - i. Precision
 1. Measure of how closely individual measurements agree with one another
 - ii. Accuracy
 1. How closely individual measurements agree with the correct, or "true" value
- d. SigFigs
 - i. Measured quantities are generally reported in such a way that only the last digit is uncertain
 - ii. There should only be one sigfig for the uncertainty
 - iii. Count the sigfigs bro
 - iv. What bout zeroes?????????
 1. Zeros between nonzero digits are significant
 2. Zeros at beginning of any numbers are not
 3. Zeros that are at an end of a number WITH DECIMALS are significant
- e. Sigfigs in calculations
 - i. The least certain measurement limits the certainty of the calculated quantity
 1. Multiplication and division
 - a. Result contains the same number of sigfigs as the measurement with the fewest sigfigs

2. Addition and Subtraction
 - a. Result has the same number of decimal places as the measurement with the fewest decimal places
- ii. Rounding numbers
 1. <5
 - a. Round down
 2. Greater or equal to 5
 - a. Round up
- iii. Don't round till the end

1.6 Dimensional Analysis

- a. Dimensional analysis
 1. Carry units through all calculations
 2. Units are multiplied together, divided into each other, or "canceled"
- b. Conversion factor
 - i. A fraction whose numerator and denominator are the same quantity as expressed in different units
 - ii. $\text{Given unit} * \frac{\text{desired unit}}{\text{given unit}} = \text{desired unit}$
- c. Using two or more conversion factors
 - i. Very possible
- d. Conversions involving volume
 - i. Same concept you just need to know the conversions

Chapter 1 reflections:

Didn't really understand the difference between Physical Change and Chemical change, which really screwed me over. Also missed a very easy density question (water displacement) and failed to count sigfigs on a thermometer correctly (just count the lines lol. So just do a better job of understanding concepts and we should be really fine. D: 16/20

Curved to 15 so nice

Chapter 2

2.1 The Atomic Theory of Matter

- a. Dalton
 - i. Each element is composed of small particles called atoms
 - ii. Atoms of an element are identical, but the atoms of one element are different from the atoms of all other elements
 - iii. Atoms of an element are not changed into atoms of a different element by chemical reactions, atoms are not created nor destroyed
 - iv. Compounds are formed when atoms of more than one element combine
- b. Law of constant composition
 - i. In a given compound, the relative numbers and kinds of atoms are present
- c. Law of conservation of mass/matter
 - i. Total mass of materials present after a chemical is the same as before the reaction
- d. Law of Multiple Proportions
 - i. If elements a and b combine to form more than one compound, the masses of B can combine with a given mass of A in the ratio of small whole numbers

2.2 The Discovery of Atomic Structure

- a. Atom
 - i. Made of smaller subatomic particles
- b. Cathode Rays and Electrons
 - i. Consisted of waves of radiation
 - ii. J.J Thomson
 - 1. Found the electron from cathode rays
- c. Radioactivity
 - 1. Some elements emit high-energy radiation
 - a. Alpha, Beta, gamma radiation
- d. Nuclear Atom
 - i. "Plum-pudding" model discovered by Thomson
 - ii. Rutherford and Gold Foil Experiment

2.3 Modern View of Atomic Structure

- a. Proton, neutron, electron
 - i. Electron charge is -1.602×10^{-19} Coulombs (C)
 - ii. Generally expressed as -1
 - iii. Atoms have same amounts of electrons and protons
 - iv. Atomic mass unit
 - 1. Describes 1.66054×10^{-24} g of atoms
 - v. The nucleus of the cell is very dense
- b. Atomic Numbers, Mass Numbers, Isotopes
 - i. Atomic number

1. Number of protons in the nucleus of an atom
2. Equal to the amount of electrons
3. Know the form of an atom (pg 45)
- ii. Isotopes
 1. Atoms with identical atomic numbers but different mass numbers (amount of neutrons)
 2. Carbon 12- weight of 12.

2.4 Atomic Weights

- a. The Atomic Mass Scale
 - i. Hydrogen is the lightest element so it has the mass of 1. Other elements had their masses determined relative to this value
 1. $1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g}$
 2. $1 \text{ g} = 6.022 \times 10^{23}$
- b. Average Atomic masses
 - i. Many elements have their unique isotopes so the average atomic mass is just the average weight of all of the isotopes
 1. Known as atomic weight

2.5 Periodic Table

- a. Periodic table
 - i. Most significant tool chemists use for organizing and remembering chemical facts
 - ii. Periods
 1. Horizontal rows
 - iii. Groups
 1. Vertical columns
 2. Exhibit similarities in physical and chemical properties
 3. Leftmost group are generally metallic elements (metals)
 4. Metals are generally solids at room temperature
 5. The staircase thing separates metals from nonmetals
 6. Metalloids
 - a. metal/non-metal hybrids

2.6 Molecules and Molecular compounds

- a. Molecule
 1. Assembly of two or more bonded atoms
- b. Molecules and Chemical Formulas
 - i. Chemical Formula
 1. Diatomic molecules start at 7 form a 7, also hydrogen
 - ii. Molecular compounds
 1. Contain more than one type of atom
- c. Molecular and Empirical formulas
 - i. Molecular formulas
 1. Indicate the actual numbers and types of atoms in a molecule
 - ii. Empirical formula
 1. These give the smallest possible whole-number ratios
- d. Picturing models
 - i. Structural formula

1. Shows which atoms are attached to which within the molecule

2.7 Ions and Ionic Compounds

- a. Cation
 - i. Positive charge
- b. Anion
 - i. Negative charge
- c. Polyatomic ions
 - i. Atoms joined together in a molecule, but have a net positive or negative charge
- d. Predicting Ionic Charges
 - i. Group 1, 2, 3 will lose electrons
 - ii. Group 5,6,7 will gain electrons
- e. Ionic Compounds
 - i. Ions form when electrons are transferred from one substance to the other
 - ii. NaCl
 - iii. Cross over trick thing

2.8 Naming inorganic compounds

- a. Chemical nomenclature
 - i. System for naming substances
 - ii. Some substances have common names, but others there is a system to find out
- b. Names and Formulas of Ionic Compounds
 - i. Positive ions (cations)
 1. Cations formed from metal atoms have the same name as the metal:
 - a. Na^+ sodium ion
 2. If a metal can form different cations, the positive charge is indicated by roman numerals (always a positive charge)
 - a. Transition metals
 - b. Fe^{3+} iron(III) ion
 - c. Also can have the ending -ous/-ic
 - i. -ic has a larger charge than -ous
 3. Cations formed from nonmetal atoms have names that end in -ium
 - a. NH_4^+ ammonium ion
 - b. H_3O^+ hydronium ion
 - ii. Negative ions (anions)
 1. The names of monatomic anions are formed by having ending of -ide
 - a. H^- hydride ion
 - b. Can have ending of ide
 - c. OH^- Hydroxide ion
 2. Polyatomic anions containing oxygen can have names ending in -ate or -ite
 - a. Nitrate NO_3^-
 - b. SO_4^{2-} Sulfate
 - c. Prefixes extend to four members
 - d. -ate is most common, -ite are the others.
 3. Anions derived by adding H^+ to an oxyanion are named by adding as a prefix the word hydrogen or dihydrogen
 - a. CO_3^{2-} Carbonate ion
 - b. HCO_3^{2-} Hydrogen Carbonate ion

iii. Ionic Compounds

1. Names of ionic compounds consist of the cation name followed by the anion name

c. Names and Formulas of Acids

- i. Acids containing anions ending in -ide are named by changing the -ide to -ic, adding the prefix hydro- to this anion name, and then following the word acid
- ii. Acids containing anions whose names end in -ate or -ite are named by changing -ate to -ic and -ite to -ous and then adding "acid"

d. Names and formulas of Binary Molecular Compounds

- i. The name of the element farther to the left in the periodic table is written first
 1. Oxygen is always written last except when combined with fluoride.
- ii. If both elements are in the same group, the one having the higher atomic number is named first
- iii. The name of the second element is given and -ide ending
- iv. Greek prefixes are used to indicate the number of atoms in each element, with mono- never used with the same element

2.9 Some Simple Organic Compounds

a. Organic chemistry

- i. Study of compounds of carbon

b. Alkanes

- i. Hydrocarbons
 1. Contain only carbons and hydrogens
- ii. CH_4 , C_2H_6 , $\text{C}_n\text{H}_{2n+2}$

c. Derivatives of Alkanes

- i. You can add functional groups to alkanes that replace an H in the alkane
 1. There is also a prefix as to where the functional group is attached to the molecule

Reflections:

Eh curve probably still fine for me maybe but yeah i really need to memorize more concepts that are in the book to do well on quizzes

Chapter 3

3.1 Chemical equations

- a. Chemical equations
 - i. Plus sign: Reacts with
 - ii. Arrow: produces
 - iii. Left: reactants
 - iv. Right: products
 - v. There are also coefficients
 - vi. Equations must be balanced
- b. Balancing equations
 - i. If we know the formulas of reactants and products we can write and balance the equation
 - 1. Should include smallest possible whole-number coefficients
 - 2. Subscripts aren't changed
 - 3. Balance the elements that occur the least often on each side of the equation
- c. Indicating the states of reactants and products
 - i. Add (g), (l), (s), (aq)
 - ii. Add a Δ above arrow when heat is added

3.2 Simple Patterns of Chemical Reactivity

- a. Combination and decomposition reactions
 - i. Combination reactions
 - 1. 2 or more substances react to form one product
 - 2. Elements combine to form compounds
 - 3. Metal + nonmetal = ionic solid
 - ii. Decomposition reaction
 - 1. One substance undergoes a reaction to produce two or more other substances
 - 2. Heating a substance
 - 3. When it is burning it often only a CO₂ breaks off
- b. Combustion in Air
 - i. Combustion reactions
 - 1. Produce a flame, fast reactions
 - 2. Produces CO₂ and H₂O if you add some O₂
 - ii. Oxidation reactions
 - 1. Combustion reactions that happen in the body

3.3 Formula Weights

- a. Formula and Molecular Weights
 - i. Formula Weight
 - 1. Sum of the atomic weights of each atom in its chemical formula
 - 2. Amu
- b. Percentage composition from formulas

- i. Percentage composition
 - 1. Percentage mass contributed by each element in the substance
 - a. $\% \text{ element} = \frac{(\# \text{ of atoms of that element})(\text{amu})}{\text{formula weight}} * 100\%$
 - i. Mass of an element divided by the total amu of compound

3.4 Avagadro's Number and the Mole

- a. Mole
 - i. Basically just like a dozen/gross
 - ii. The amount of matter that contains as many objects as the number of atoms in 12 g of pure carbon-12
- b. Avogadro's Number
 - i. $6.022 * 10^{23}$
 - ii. 1 mol of carbon 12 atoms: $6.02 * 10^{23}$ atoms
 - iii. 1 mol H₂O atoms: $6.02 * 10^{23}$ H₂O
- c. Molar Mass
 - i. A mol is constant amount of molecule, but has varying mass
 - ii. Count the amu of the compound, which equals the grams
 - iii. Molar mass
 - 1. g/mol
- d. Interconverting masses and moles
 - i. Use conversions
- e. Interconverting Masses and Numbers of Particles

3.5 Empirical Formulas from Analyses

- a. How 2 solve:
 - i. Given the percentages of the mass of elements
 - ii. Assume that it is a 100g sample
 - iii. Find grams of each element
 - iv. Convert it into moles
 - v. Calculate mole ratio by dividing by the smallest value
- b. Molecular formula from empirical formula
 - i. Multiple: $(\text{molecular weight})/(\text{empirical weight})$
 - 1. Multiply the empirical formula subscripts by this multiple
- c. Combustion Analysis
 - i. Helps determine empirical formulas
 - ii. Compound containing carbon and hydrogen is combusted into chambers that absorb the gas
 - iii. Carbon is converted into CO₂, Hydrogen into H₂O
 - iv. From the mass of CO₂ and H₂O we can calculate the number of moles in C and H

Quantitative Information from Balanced Equations

- a. The coefficients in a balanced equation indicate both the relative numbers of molecules (or formula units) involved in the reaction and the relative numbers of moles.
- b. If you have like 5 grams of a reactant and want to see how much grams of a product is then the process is:
 - i. Grams of reactant

- ii. Convert into moles of reactant
- iii. Convert into moles of product
- iv. Convert into grams of product

3.7 Limiting Reactants

- a. Limiting reactant
 - i. Determines/limits the amount of product formed
- b. Excess reactants
 - i. Too much of something, not reacted
- c. Lol how do you solve this????????????????????
 - i. Say you have: $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
 - ii. And you have 15 mol of N_2 and 25 mol of H_2
 - iii. $\text{Mol H}_2 = 15 \text{ mol N}_2 * \frac{3 \text{ mol H}_2}{1 \text{ mol N}_2} = 45 \text{ mols of H}_2$
 - iv. We need 45 mols of H_2 , and the experiment provides 25, so it is limiting
 - v. The N_2 is the excess reactant
 - vi. If you are looking for the mol of final product, then just use the mols of limiting and reactant and convert that into mols of product
- d. Theoretical yields
 - i. $\text{Percent Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} * 100\%$

Additional notes

- a. Converting something into "amount of molecules"
 - i. To do this you just multiply the amount of moles by the avogadro constant.
 - 1. And if it asks you to find the amount of "oxygen molecules in sulfate" or something just multiply the amount of molecules by the amount of oxygen atoms in the molecules.
- b. "What is the molar mass if .5g of bleach is .0012 mol?"
 - i. Just set this up like a proportion
- c. More on limiting/excess reactants
 - i. Make sure that you start with the AMOUNT of mols/grams you have every time that you do an conversion
 - ii. So how much of an excess reactant is left if you put in a certain amount of limiting reactant?
 - 1. Use the limiting reactant equation to find the amount of moles that the excess reactant uses (set up a mol conversion)
 - 2. Then, it is the amount given (moles) minus the amount used (moles)
- d. Theoretical vs actual yield
 - i. If a problem asks you for the the amount in grams if there is 92% yield just multiply the final result by 92/100
 - ii.

Reflections

This chapter wasn't really hard, since the thing wasn't that bad because i've learned to memorize more stuff from the book :D. Also i've learned how to do stoichiometry which help

Chapter 4

Aqueous solution

Water is the dissolving medium

4.1 General properties of aqueous solutions

- a. Solutions
 - i. Homogeneous mixture of two or more substances
 - ii. Solvent
 1. Greatest in quantity, dissolves the solute
 - iii. Solute
 1. Other substances in the solution
- b. Electrolytic properties- tested using "conductivity probe"
 - i. Some solutions may be better at conducting electricity
 1. For example, sodium and water vs sugar and water
 - ii. Electrolyte
 1. Aqueous solution that contains ions
 2. NaCl
 - iii. Nonelectrolyte
 1. Does not contain ions but dissolves in water
 2. Hydrocarbons
- c. Ionic compounds in water
 - i. Dissociates
 1. How ions dissolve in water
 2. Water dissociates other things because the O is negative and the H atoms are positive
 3. Solvation
 - a. Process that prevents dissociated cations and anions from recombining
 - b. Water separates ionic compounds into their cations and anions
 - ii. Water is a very good solvent due to its positive and negative sides
 1. Ions are surrounded by H₂O molecules
 2. Ions
 - a. Memorize all of the common ions and stuff
- d. Molecular compounds in water
 - i. Solution mostly consists of intact molecules that are dispersed throughout the solution
 - ii. Most of these are nonelectrolytes
 - iii. Exceptions:
 1. Acids
 - a. HCl ionizes in water
 - i. Dissociates into H⁺ and Cl⁻ ions
- e. Strong and Weak Electrolytes
 - i. Strong electrolytes
 1. Solutes that exist in solution completely or nearly completely
 2. NaCl and other compounds such as HCl
 - ii. Weak electrolytes

1. Solutes that exist in solution mostly in the form of molecules with only a small fraction in the form of ions
 2. Acetic acid
 3. $\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$
 - a. 99% of the solute is still $\text{HC}_2\text{H}_3\text{O}_2$
 4. This is in equilibrium, so some ions are rejoining, while others are splitting (Chemical Equilibrium)
- iii. Strong vs weak
1. Strong often only has one arrow, whereas the weak ones are in chemical equilibrium and therefore has a double arrow

4.2 Precipitation reactions

- a. Introduction
 - i. Precipitation reaction
 1. Reaction that results in the formation of an insoluble product
 - ii. Precipitate
 1. Insoluble solid formed by a reaction in a solution
- b. Solubility guidelines for ionic compounds
 - i. Solubility
 1. Amount of a substance that can be dissolved in a given quantity of solvent at that temperature
 - ii. Insoluble
 1. Solubility less than .01 mol/L
 - iii. All ionic compounds with nitrate (NO_3^-) are soluble in water
 1. Memorize more of this
- c. Exchange (metathesis) Reactions
 - i. Exchange reactions (metathesis reactions) (double replacement reaction)
 1. $\text{AX} + \text{BY} \rightarrow \text{AY} + \text{BX}$
 - ii. To complete and balance the equation:
 1. Use the chemical formulas of the reactants to determine the ions that are present
 2. Write down the chemical formulas of the products by combining the cation from one reactant with the anion of the other
 3. Balance the equation
- d. Ionic Equations
 - i. Molecular equations
 1. Show the chemical formulas of the reactants and products without indicating their ionic character
 - ii. Complete ionic equation
 1. These show the ions and the strong electrolytes
 - iii. Spectator ions
 1. Ions that appear in identical forms among both the reactants and products of a complete ionic equation
 2. They play no role in the reaction
 3. Net ionic equation
 - a. Omits the spectator ions from the equation
 - b. If every ion in a complete ionic equation is a spectator, then no reaction occurs

4. Writing net ionic equations
 - a. Write a balanced molecular equation for the reaction
 - b. Use solubility rules from table 4.1 to determine which products are soluble and which are insoluble. Only the insoluble ones (precipitates) and the ions from which they came participate in the reaction
 - c. Use the information from step 2 to create a complete ionic equation, placing (aq) next to all the ions that are soluble and (s) next to all the precipitates
 - i. Don't saw the insoluble molecules apart. Furthermore, only saw apart strong electrolytes are dissolved in aqueous solution.
 - d. Cancel out, on both sides of the equation, any ions that are (aq) on both sides. Cancelled-out ions are called spectator ions
 - e. Rewrite your complete ionic equation with your spectator ions removed. The resulting equation is the net ionic equation

4.3 Acid-Base Reactions

- a. Acids
 - i. Acids
 1. Substances that ionize in aqueous solutions to form hydrogen ions, increasing the concentration of H^+ ions (protons)
 2. Proton donors
 3. Acids are surrounded and bound by water molecules
 - ii. Monoprotic acids
 1. Yield one H^+ per molecule of acid
 2. HCl HNO_3
 - iii. Diprotic acid
 1. Yields two H^+ per molecule of acid
 2. H_2SO_4
- b. Bases
 - i. Bases
 1. Substances that accept (react with) H^+ ions
 2. Bases produce hydroxide (OH^-) when they dissolve in water
 3. $NaOH$, KOH , $Ca(OH)_2$, NH_3
- c. Strong acids
 - i. Acids with ions that completely dissociate in water to generally form H^+ (protons)
 1. $HCl(g) \rightarrow H^+ + Cl^-$
 - a. HCl
 - b. HBr
 - c. HI
 - d. $HClO_3$
 - e. $HClO_4$
 - f. HNO_3
 - g. H_2SO_4
- d. Strong bases
 - i. Bases with ions that completely dissociate in water to generally form OH^-
 1. $NaOH(s) \rightarrow Na^+ + OH^-$
 2. Group 1a metal hydroxides
 - a. $LiOH$, $NaOH$, KOH , $RbOH$, $CsOH$
 3. Heavy group 2a metal hydroxides

- a. Ca(OH)_2 , Sr(OH)_2 , Ba(OH)_2
- e. Strong and Weak Acids and Bases
 - i. Strong acids and strong bases
 - 1. Acids and bases that are strong electrolytes
 - 2. More reactive
 - ii. Weak acids and weak bases
 - 1. Weak electrolytes
 - iii. Memorize the common strong acids and bases\
- f. Identifying Strong and Weak Electrolytes
 - i. All ionic are strong
 - ii. For molecular compounds, the strong acids are strong electrolytes
 - 1. The weak electrolytes are weak acids and bases
- g. Neutralization Reactions and Salts
 - i. Acids are sour
 - ii. Bases are bitter
 - iii. Acids can change the color of dyes differently than bases
 - iv. Neutralization reaction
 - 1. When a solution of an acid and a solution of a base are mixed
 - a. They always make H_2O and a salt
 - b. These are a kind of metathesis reaction
- h. Writing balanced neutralization equation
 - i. They look like this:
 - 1. $\text{H}_x\text{B} + \text{C}_y\text{OH} \rightarrow \text{HOH} + \text{C}_w\text{B}_z$
 - 2. Be sure that the charges of C and B balance
 - 3. Alter W and Z to make it so the sum of C and B's charges equal zero
 - ii. Proceed like it is a net ionic equation
 - 1. Determine what is soluble and what is not soluble
 - a. Do this check on the PRODUCTS side
 - b. The soluble gets an (aq) next to it
 - c. The insoluble gets an (s) next to it
 - 2. Saw the equation in half
 - a. The (aq) gets sawed in half, make sure to keep the charge
 - i. If it is a $(\text{OH})_5$ then it is rewritten like 5OH^-
 - b. Do not saw the insoluble solids in half (or the liquid)
 - c. This is the net ionic equation
 - 3. Cancel the things that stay the same in the equation
 - 4. Done
- i. Acid-Base reactions with gas formation

4.4 Oxidation-Reduction (Redox) Reactions

- i. Precipitation reactions
 - 1. Cations and anions come together to form an insoluble ionic compound
- ii. Neutralization reactions
 - 1. H^+ ions and OH^- ions come together to form H_2O molecules
- iii. Oxidation reactions
 - 1. Electrons are transferred from one reactant to another
 - 2. LEO the lion goes GER

- iv. We must keep track of the electrons gained by the substance and the electrons lost by the substance
 - 1. Oxidation numbers track this
- v. How to determine oxidation number?
 - 1. For an atom in its elemental form (an element that isn't bonded to any other kinds of elements but itself) the oxidation number is zero
 - 2. For monatomic ions, the oxidation number is the ion's charge
 - a. K^+ , S^{2-}
 - 3. Nonmetals usually have negative oxidation numbers, although they can sometimes be positive
 - a. The oxidation number of oxygen is usually -2, except in peroxides (H_2O_2), where it is -1
 - b. The oxidation number of hydrogen is usually +1 when bonded to nonmetals and -1 when bonded to metals
 - c. The oxidation number of fluorine is -1 in all compounds. The other halogens have oxidation numbers of -1 in most cases. When bonded to oxygen, however, they have positive oxidation states
 - 4. The sum of the oxidation numbers of all atoms in a neutral compound is zero. The sum of the oxidation numbers in a polyatomic ion equals the charge of the ion
- vi. Oxidation of metals by acids and salts
 - 1. $A + BX \rightarrow AX + B$
- vii. By determining each element's oxidation numbers, we can figure out what elements are oxidized and what elements are reduced
- viii. How?
 - 1. Figure out the oxidation number of every element in the chemical equation
 - 2. Any elements whose oxidation numbers become more positive, going from reactants to products, have lost electrons and are therefore **oxidized**
 - 3. Any elements whose oxidation numbers become more negative, going from reactants to products, have gained electrons and are therefore **reduced**
 - 4. Things can also stay the same, which aren't mentioned
 - a. Note: Elements that are oxidized in a redox reaction are called reducing reagents (or reductants). Elements that get reduced in a redox reaction are called oxidizing reagents (or oxidants)

4.5 Concentrations of Solutions

- j. Intro
 - i. Concentration
 - 1. Amount of solute dissolved in a given quantity of solvent or quantity of solution
 - 2. The greater the amount of solute, the more concentrated the solution is
- k. Molarity (M)
 - i. Is a way of expressing a solution's concentration
 - 1. Moles of solute/(volume of solution in liters)
 - a. Units are moles/liters
 - ii. To do a problem
 - 1. Remember that molarity's units are moles/liters
 - 2. Look at what information -especially which units!- you're given

3. Determine what you're being asked to find - especially what units your final answer should have
4. Use dimensional analysis -focusing on units!- to convert what you're given into what you're being asked to get. In other words, use dimensional analysis to move from your starting point to your destination, focusing on units!!!!!!
- iii. What to start with?
 1. Start with the volume if possible
 - a. If it gives the volume and the molarity
 2. Start with the moles if you are finding the volume
- l. Expressing the concentration of an electrolyte
 - i. The relative concentrations of the ions introduced in the solution depend on the chemical formula of the compound
 1. 1 M solution of NaCl has 1 M of Na⁺ ions and 1 M Cl⁻ ions
 2. 1 M solution of Na₂SO₄ has 2M of Na⁺ ions
- m. Interconverting molarity, moles, and volume
 - i. Like calculating density or something like that
- n. Dilution
 - i. Dilution
 1. Adding water to obtain a solution with lower molarity
 2. Use this equation
 - a. $M_1V_1 = M_2V_2$
 - b. Moles solute in concentration = moles solute in dilution
 - ii. These questions usually involve diluting a more concentrated solution (called a stock solution) to prepare another solution at a lower concentration
 - iii. How much water do you need to add to reach X molarity?
 1. This is generally set up into a proportion
 - a. $m/v = m/v$
- o. Another equation
 - i. Weight of solute required= molarity*molecular mass*volume of solution (ml)/1000

4.6 solution stoichiometry and chemical analysis

- a. Neutralization Reactions
 - i. We can add acids to bases to neutralize them, As discussed earlier, acid base reactions are called neutralization reactions
 1. To neutralize a base, we add the same number of moles of acid as there are moles of base in the solution

Didn't memorize the solubility rules so idk ٩(ツ)٩

Chapter 5

Thermochemistry

Some stuff:

- a. Polyatomic elements
 - i. These cannot exist as single uncharged atoms in their elemental form
 - 1. Start at 7, form a 7, these are the diatomic elements
 - ii. They can form as single charged atoms

5.1 The Nature of Energy

- a. Intro
 - i. Energy
 - 1. The capacity to do work or to transfer heat
 - ii. Work (w)
 - 1. Energy used to cause an object with mass to move
 - a. $\text{Work} = \text{Force} \times \text{Distance}$
 - iii. Heat (q)
 - 1. **Energy** used to cause the temperature of an object to increase
- b. Kinetic energy and potential energy
 - i. Kinetic energy
 - 1. Energy of motion
 - a. A system actively producing or exerting energy is called kinetic energy
 - b. $\frac{1}{2}mv^2 = E$
 - ii. There is also potential energy
 - 1. Energy from its position relative to other objects
 - 2. System that is positioned to be able to produce energy
 - a. $E = mgh$
 - iii. Force
 - 1. Any kind of push or pull exerted on an object
 - iv. Electrostatic potential energy
 - 1. $E = \frac{k Q_1 Q_2}{d}$
 - a. $k = 8.99 \times 10^9 \text{ J}\cdot\text{m}/\text{C}^2$
 - b. Q_1 and Q_2 are the order of magnitude of the charge of the electron
 - i. $1.60 \times 10^{-19} \text{ C}$
 - ii. When the 2 elements have the same sign, the sign is positive
 - iii. When they have opposite charges, then E is negative
 - c. d = distance between the 2 particles
 - v. What is energy transfer?
 - 1. Energy is transferred from one system to another as heat, work, or a combo of both
- c. Units of Energy

- i. Joule (J)
 - 1. Not a lot of energy $\sim 10^7$
- ii. calorie
 - 1. The amount of energy required to raise the temperature of 1 g of water by one degree
 - 2. 4.184 J
 - a. Calorie with a capital C is 1000 calories
- d. System and Surroundings
 - i. System
 - 1. The portion we single out for study
 - a. Reactants and products
 - ii. Surroundings
 - 1. Everything else
 - a. Container
 - iii. Closed system
 - 1. Nothing can get in or out or in
 - iv. Open system
 - 1. Matter can get in or out
 - v. Reactions can exchange energy with its surroundings in the form of work and heat
 - vi. Isolated system
 - 1. Neither energy nor matter may be exchanged with the surroundings
- e. Transferring Energy: Work and Heat
 - i. Work
 - 1. Energy used to cause an object to move against a force
 - a. $W = F \cdot d$
 - i. Work equals force times the distance
 - ii. Heat
 - 1. Energy transferred from a warmer object to a colder one
 - 2. Energy transferred between a system and its surroundings as a result of their difference in temperature

5.2 The First Law of Thermodynamics

- a. Intro
 - i. First law of thermodynamics
 - 1. Energy is conserved
 - 2. Any energy that is lost by a system is gained by the surroundings, vice versa
 - ii. A system can gain/lose energy from surroundings
- b. Internal energy
 - i. Internal energy
 - 1. Is the sum of all the kinetic and potential energies of all its components
 - ii. Change in internal energy:
 - 1. $\Delta E = E_{final} - E_{initial}$
 - 2. Positive value of ΔE
 - a. $E_{final} > E_{initial}$
 - i. Energy gained from surroundings (endothermic)
 - 3. Negative value of ΔE
 - a. $E_{final} < E_{initial}$

- i. Loss of energy to surroundings (exothermic)
- c. Relating ΔE to Heat and Work
 - i. A system may exchange energy with its surroundings as heat or as work
 - 1. When heat is added to a system or work is done on a system, its internal energy increases
 - ii. First law of thermodynamics equation
 - 1. $\Delta E = q + w$
 - a. Q is the heat given/liberated from the system
 - i. If q is negative, the system loses heat
 - ii. If q is positive, the system gains heat
 - b. W is the work done/ done by the system
 - i. If w is negative, work has been done by system
 - ii. If w is positive, work has been done on system
 - c. E is change in energy
 - i. If e is positive, the system gains energy
 - ii. If e is negative, there is a net loss of energy
- d. Exothermic and Endothermic Processes
 - i. Endothermic
 - 1. The system absorbs heat
 - 2. Ex: melting of ice
 - a. Heat flows into the system from its surroundings
 - ii. Exothermic
 - 1. System loses heat
 - a. Heat exits or flows out of the system and into the surroundings
- e. State functions
 - i. State function
 - 1. A property of a system that is determined by specifying the system's condition, or state
 - 2. The value of a state function depends only on the present state of the system, not on the path of the system took to reach that state
 - ii. All systems have some amount of internal energy. It doesn't matter how you made it

5.3 Enthalpy

- a. Intro
 - i. Pressure-volume work
 - 1. Work involved in the expansion or compression of gases (when gasses do work)
 - a. $W = -P \Delta V$
 - i. P is pressure
 - 1. ¿pressure of surroundings?
 - ii. V is change in volume of the system
 - ii. Enthalpy
 - 1. Heat flow in processes occurring at constant pressure when no forms of work are performed other than P-V work
 - a. $H = E + PV$
 - i. Enthalpy equals internal energy plus the P-V work
 - 2. When a change occurs at a constant pressure:
 - a. $\Delta H = \Delta E + P\Delta E$
 - 3. There's also this one lol, rearranged from the other equations

- a. $(qp + w) - 2 = qp$
- 4. The change in enthalpy equals the heat gained or lost at constant pressure
- 5. Positive H
 - a. Endothermic
- 6. Negative H
 - a. Exothermic

5.4 Enthalpies of Reaction

- a. Intro
 - i. $\Delta H = H_{\text{products}} - H_{\text{reactants}}$
 - 1. Enthalpy of reactions (heat of reaction)
 - a. This is the enthalpy change that accompanies a reaction
 - b. If ΔH is negative, then the reaction is exothermic
 - c. ΔH is reported at the end of a balanced equation, measured in kilojoules
 - 2. Thermochemical equations
 - a. Show the enthalpy change
- b. Guidelines when using thermochemical equations and enthalpy diagrams
 - i. Enthalpy is an extensive property
 - 1. The magnitude of ΔH , therefore, is directly proportional to the amount of reactant consumed in the process.
 - 2. 2 times more reactants (2 times more moles)= 2 times more heat generated
 - ii. The enthalpy change for a reaction is equal in magnitude, but opposite sign, to ΔH for the reverse reaction
 - 1. This is what happens if you flip a reaction around
 - iii. The enthalpy change for a reaction depends on the state of the reactants and products
 - 1. Different if a reactant is like water or a solid

Calorimetry 5.5 POG

- a. Intro
 - i. Calorimetry
 - 1. A measure of heat flow
 - ii. Calorimeter
 - 1. A device used to measure heat flow
- b. Heat capacity and specific heat
 - i. The more heat an object gains, the hotter it gets
 - ii. Heat capacity (C)
 - 1. The heat required to raise its temperature by 1 K
 - iii. Molar heat capacity
 - 1. The heat capacity of one mole of a substance
 - 2. You solve this by multiplying the specific heat capacity by the weight of the substance in grams
 - iv. What is the specific heat capacity of 185 grams of a substance?????
 - 1. Just multiply the 185 grams by the specific heat capacity
 - v. Specific heat capacity
 - 1. Heat capacity of one gram of a substance
 - 2. $Q = MCT$

- c. Constant Pressure Calorimetry
 - i. Help us to measure the enthalpies of specific reactions by measuring the heat that they produce
- d. Bomb Calorimetry

Hess's Law

- a. Intro
 - i. If a particular reaction can be carried out in one step or in a series of steps, the sum of the enthalpy changes associated with the individual steps must be the same as the enthalpy change associated with the one step process
- b. Hess's law
 - i. If a reaction is carried out in a series of steps, the change in enthalpy for the overall reaction will equal the sum of the enthalpy changes for the individual steps
- c. Calculating change in enthalpy for multistep reactions
 - i. If it gives you 2-3 equations and tells you a target equation, just manipulate your given equations so that they can add up and cancel so they can get into the form of the target equation

Enthalpies of formation 5.7

- a. Enthalpy of formation
 - i. Energy needed for vaporization, fusion, or combustion
- b. Standard enthalpy change
 - i. The enthalpy change when all reactants and products are in their standard states
- c. Standard enthalpy of formation
 - i. Change in enthalpy for the reaction that forms one mole of the compound from its elements, with all substances in their standard states.
- d. Using enthalpies of formation to calculate enthalpies of reaction
 - i. Just products-reactants
- e. Calculating Reaction enthalpy values from final enthalpy values
 - i. Write out a chemical reaction describing the formation of each component in your overall reaction
 - ii. Make sure that all the chemical equations you've written out in step 1 add up to give you the overall reaction in the problem
 - iii. Use appendix c to write the individual final enthalpy values for each equation in step 1. If the equation are written backwards, then change the sign of the final enthalpy value.
 - iv. Add up all the formation enthalpy values from your reactions to obtain your final enthalpy for the total reaction

5.8 Foods and Fuels

- a. Intro
 - i. Fuel value
 - 1. Energy released when one gram of a material is combusted
- b. Foods
 - i. Glucose is an example of a carbohydrate
 - 1. It is rapidly combusted in the body
 - 2. The value of carbohydrate is 4 cal/g

- ii. Excess energy is stored as fat
 - 1. They are insoluble in water, so they can be stored easily
 - 2. They have higher energy density 9 cal/g
- c. Fuels
 - i. Coal
 - 1. High fuel value because of the large carbon content
 - 2. Large shipping costs
 - ii. Fossil fuels
 - 1. Coal, petroleum, natural gas
 - iii. Natural gas
 - 1. Gaseous hydrocarbons
 - iv. Petroleum
 - v. Syngas
 - 1. Gas that is made from coal when it is pulverized and treated with steam. This produces a mixture of gasses that have high fuel values
- d. Other energy sources
 - i. Nuclear energy
 - 1. Free of polluting glasses but have radioactive waste
 - ii. Renewable energy
 - 1. Solar, wind, geothermal, hydroelectric, biomass
 - 2. Solar energy is the largest source of renewable energy

Hard problems to know

<https://www.chegg.com/homework-help/questions-and-answers/don-t-understand-different-parts-this-problem-consider-following-reaction-ch3oh-co-2h2-h-90-q351657>

- a. When it is asking for the "enthalpy of formation" of a molecule its just the value of the formed molecule
- b. If it asks for delta H with a degree sign, you will have to do the

$$\text{Products} - \text{Reactants} = \Delta H^\circ$$

Chapter 6

6.1 The Wave Nature of Light

- a. Electromagnetic radiation
 - i. Includes the light that we can see with our eyes (visible light)
 - ii. This carries energy through space, so it is known as radiant energy
 - iii. Other forms of electromagnetic radiation
 - 1. X-ray
 - 2. Radio waves
 - 3. Infrared radiation (heat)
- b. Speed of light
 - i. 3×10^8 m/s
- c. Waves
 - i. Wavelength
 - 1. The distance between 2 adjacent peaks
 - ii. Frequency
 - 1. The amount of cycles that pass through a point in a given second
 - iii. Amplitude
 - 1. How deep the peaks are
 - iv. Relationship between wavelength and frequency
 - 1. $v\lambda = c$
 - a. v is the frequency
 - b. λ is the wavelength
 - c. c is the speed of light
 - v. Hertz (Hz)
 - 1. The cycles per second

6.2 Quantized Energy and Photons

- a. Intro
 - i. What does the wave model of light not explain?
 - 1. The emission of light from hot objects (blackbody radiation)
 - 2. Emission of electrons from metal surfaces on which light shines
 - 3. The emission of light from electronically excited gas atoms
 - b. Hot objects and the quantization of energy
 - i. Heated solids emit radiation
 - ii. Max planck made up an assumption:
 - 1. $E = hv$
 - a. h is the plancks constant
 - i. 6.626×10^{-34} joules seconds
 - b. E is the energy
 - c. v is the frequency of radiation
- c. The photoelectric effect and photons
 - i. Photoelectric effect

1. When photons of sufficiently high energy strike a metal surface, the electrons are emitted from the metal

6.3 line spectra and the bohr model

- a. Line spectra
 - i. Most radiation sources produce radiation containing many different wavelengths
 - ii. Spectra
 1. When radiation from such sources is separated into its different wavelengths, a spectrum is produced
 - iii. Continuous spectrum
 1. The rainbow of colors that contains light of all wavelengths
 - a. Rainbow
 - iv. Line spectra
 1. Spectrum containing radiation of only specific wavelengths
 - v. Rydberg equation
 1. Allowed the calculation of the wavelengths of all the spectral lines of hydrogen
 - a. $(-2.18 \times 10^{-18}) \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$
- b. Bohr's model
 - i. 3 postulates
 1. Only orbits of certain radii, corresponding to certain definite energies, are permitted for the electron in a hydrogen atom
 2. An electron in a permitted orbit has a specific energy and is in an "allowed" energy state. An electron in an allowed energy state will not radiate energy and therefore will not spiral into the nucleus
 3. Energy is emitted or absorbed by the electron only as the electron changes from one allowed energy state to another. This equation is $E = h\nu$
- c. The energy states of the hydrogen atom
 - i. Big ass equation
 - ii. Ground state
 1. The lowest energy state of an electron
 - iii. Excited state
 1. When an electron is anywhere except the first orbital
 - a. Another big ass equation
 - iv. Zero-energy
 1. State in which the electron is removed from the nucleus
 - v. Electrons can jump from one energy state to another by absorbing or emitting photons
 1. $E = hc/\lambda$
 - vi. Another big ass equation
- d. Limitations to the Bohr Model
 - i. Cannot explain spectra of other elements (other than the hydrogen) in any way but a very crude way

6.4 the wave behavior of matter

- a. Intro

- i. Louis de Broglie
 - 1. $\lambda = h / mv$
 - a. h is Planck's constant, m and v are mass and velocity
- ii. Mv
 - 1. Is the momentum
 - 2. Broglie used matter waves to describe the wave characteristics of material particles
- b. The uncertainty principle
 - i. Trying to use physics to calculate the physics of an electron???????
 - ii. Uncertainty principle
 - 1. The science of nature becomes iffy when looking at a subatomic level
 - a. RANDOM ASS EQUATION

6.5 Quantum Mechanics and Atomic Orbitals

- a. Wave functions
 - i. Ψ (psi)
 - ii. Ψ^2 tells us the electron's location when the electron is in an energy state
- b. Probability density/electron density
 - i. Pitchfork squared
- c. Orbitals and quantum numbers
 - i. Orbitals
 - 1. Describe a specific distribution of electron density in space, given by the orbital's probability density
 - ii. Steps:
 - 1. The principal quantum number (N) is a positive integer. As it increases, the orbital becomes larger, so the electron is less tightly bound to the nucleus
 - 2. The second number (azimuthal quantum number (l)) can have numbers from 0 to $n-1$ and shows the shape of the orbital (s, p, d, f, 0, 1, 2, 3)
 - 3. Magnetic quantum number (m_l) has values from $-l$ to l and describes the orientation of the orbital in space
 - iii. Electron shell
 - 1. Collection of all the orbitals with the same value of n
 - 2. Subshell
 - a. The set of orbitals that have the same n and l values

6.6 representations of orbitals

- a. S orbitals
 - i. This is spherically symmetric
 - ii. It is just a sphere
 - iii. Radial probability function
 - 1. A graph that shows the probability of where an electron is from a nucleus.
 - iv. Node
 - 1. An intermediate point at which a probability function goes to zero
- b. P orbitals
 - i. These are not symmetric
 - ii. There are 3 dumbbell-shaped orbitals that have 2 lobes

- iii. These dumbbells just sit on the x, y, and z planes of the atom
- c. D and f orbitals
 - i. The d orbital backfills once
 - 1. It has a 4 leaf clover shape
 - a. They are on the x, y, z planes, and also one where the lobes lie between the x and y axes. The final one has a doughnut shape on the xy plane
 - ii. The f orbital backfills twice
 - 1. Don't need to know the shape

6.7 many-electron atoms

- a. Orbitals and their energies
 - i. The presence of more electrons can greatly change the energy of the orbitals
 - ii. In a many-electron atom, for a given value of n, the energy of an orbital increases with increasing value of l (there are more subshells per additional energy level)
 - iii. Degenerate
 - 1. Orbitals with the same energy
- b. Electron spin and the Pauli exclusion principle
 - i. Electrons spin, which helps create a magnetic field
 - ii. Spin magnetic quantum number (m_s)
 - 1. This is just the spin of an electron
 - iii. Pauli exclusion principle
 - 1. No two electrons in an atom can have the same set of four quantum numbers n, l, m_l , and m_s
 - 2. Because m_s must be different that means an orbital can hold can only hold 2 electrons and they must have opposite spins

6.8 electron configurations

- a. Intro
 - i. Electron configuration
 - 1. The way in which the electrons are distributed among the various orbitals of an atom
 - 2. The ground state is the most stable electron configuration of an atom
 - 3. Orbital diagram
 - a. Each orbital is denoted by a box and each electron a half arrow (which the 2 arrows face the opposite directions)
 - b. Electrons are paired when there are 2 arrows in a box
 - c. An unpaired electron is one not accompanied by a partner of opposite spin
- b. Hund's Rule
 - i. For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized
 - 1. Electrons will occupy orbitals singly to the maximum extent possible and that these single electrons in a given subshell will all have the same spin magnetic quantum number

- a. Ok basically this states that because electrons repel each other, they will want to remain as far apart from each other as possible.
 - i. Aka you fill the boxes of the orbital diagram with 1 electron until you have to double up the electrons in each box (for p orbital, 3 electrons in the 3 boxes)
 - c. Condensed electron configurations
 - i. You can use the noble gasses to represent parts of the electron configuration of an atom
 - ii. Core electrons
 - 1. What the noble gasses represent
 - iii. Outer-shell electrons
 - 1. Basically what you write, also contains valence electrons
 - d. Transition metals
 - i. Contain the elements with d orbitals
 - 1. Make sure to remember the Hund's rule when filling out the orbital diagram
 - e. Lanthanides and actinides
 - i. Have the f orbital
 - ii. And eh might add some more but most are radioactive

6.9 electron configurations and the periodic table

- a. Intro
 - i. Know the spdf blocks
 - ii. Representative elements/main-group elements
 - 1. These are the S and P block of the periodic table
 - iii. F-block metals
 - 1. These are the lanthanides/actinides
 - 2.
- b. Anomalous electron configurations
 - i. Chromium takes an electron from its S shell to put into its D shell to become more stable
 - ii. This is the same for Copper, which takes an electron from the S shell to put in the D shell in order for the D shell to become more stable

Yeah not too bad

Chapter 7

Valence Electrons

- a. Outer orbitals of the shell, remain the same in every column

7.1 Development of the Periodic table

- a. Intro
 - i. Obviously there are many properties of each element that are different, so chemists tried to find a way to classify the elements
 - ii. Dmitri Mendeleev and Lothar Meyer made the first modern table of elements
 - iii. However Mendeleev is more credited because he had better ideas and thus was able to predict what several elements would be (such as Ga and Ge)
 - iv. Henry Moseley improve the table

7.2 Effective Nuclear Charge

- a. Coulomb's Law (introduction to it i think)
 - i. Force of attraction increases as the nuclear charge increases
 - ii. Force of attraction decreases as the electron moves farther from the nucleus
- b. Intro
 - i. Properties of atoms depend on their electron configurations and how strongly their outer electrons are attracted to the nucleus
 - ii. The attraction between an electron and the nucleus depends on the magnitude of the net nuclear charge acting on the electron and on the average distance between the nucleus and the electron
 - 1. This is called the effective nuclear charge (Z_{eff})
 - 2. A is the nuclear charge from the atom
 - 3. There is also a shielding effect which is the S constant
 - a. $Z_{\text{eff}} = Z - S$
 - iii. Because electrons repel each other, we will just take the net attraction
 - iv. Effective nuclear charge
 - v. Inner electrons shield the ones toward the outside

7.3 Sizes of atoms and ions

- a. Intro
 - i. Atomic size
 - 1. Cannot be measured directly
 - 2. Can be defined as:
 - ii. Nonbonding atomic radius
 - 1. Closest distance separating the nuclei between collisions of non-bonded atoms
 - a. Known as the "van der Waals radii"
 - iii. Bonding atomic radius
 - 1. Shorter than the nonbonding atomic radius
 - 2. Distance separating the nuclei of atoms when they are bonded to each other

- iv. Atomic radius is represented with a number and A with a circle on top
- b. Trends in Atomic radii
 - i. Atomic radii decreases when going across a period
 - 1. There is an increasing effective nuclear charge (Z_{eff}) as we move across a row. This charge draws the electrons into the atom, decreasing the size of the atom
 - ii. Atomic radii increases when going down a group
 - 1. This happens because there is an increase in the principal quantum number (n) of the outer electrons
 - 2. The outer electrons are further away as they are in a further energy level
 - 3. Z_{eff} increase slowly down a group, but there is additional shielding.
 - a. However outermost electrons are still held less tightly
- c. What is bond length?
 - i. Add the A circle thing (atomic radii) of both elements together
- d. What are the sizes of these atoms ranked?
 - i. Either look at a figure that tells you are use the thing from above
- e. Periodic Trends in Ionic Radii
 - i. Cations are smaller than parent atoms
 - 1. Farthest-out occupied orbitals are emptied, and there are fewer electron-electron repulsions
 - ii. Anions are larger than their parent atoms
 - 1. More electrons, so more electron-electron repulsions, so electrons spread out more in space
 - iii. For ions carrying the same charge, size increases as we go down a column in the periodic table
 - 1. Literally the same thing as the size (discussed above)
 - iv. Isoelectronic series
 - 1. A group of ions all containing the same number of electrons
 - a. O^{2-} , F^- , Na^+ , Mg^{2+} , Al^{3+}
 - i. All of these have 10 electrons
 - b. O^{2-} has 2 additional electron, it'll be the largest

7.4 Ionization energy

- a. Intro
 - i. Ionization energy
 - 1. Minimum energy required to remove an electron from the ground state of the isolated gaseous atom or ion
 - ii. First ionization energy
 - 1. Energy needed to remove first electron from neutral atom
 - iii. Second ionization energy
 - 1. Energy needed to remove the second electron
 - iv. The higher the ionization energy, the harder it is to remove the electron
- b. Variations in successive ionization energies
 - i. It takes more energy to remove each successive electron
 - 1. This is because the atom becomes more positive, so there is a larger attraction between the nucleus and the electron

- ii. There is a large energy jump between the outer electrons and the inner-shell electrons (if you are writing the formula for an atom the inner-shell is just the [noble gas] basically just what the noble gas has)
 - 1. This is because they are more likely to be found closer to the nucleus
- c. Periodic trends in first ionization energies
 - i. Trends
 - 1. In each period the ionization energy increases with increasing atomic number, with some irregularities
 - 2. Within each group of the table, the ionization energy decreases with increasing atomic number
 - 3. The representative elements show a larger range of values of ionization than do the transition-metal elements. F-block elements only show very small changes in ionisation energy
 - ii. Basically increases with the smaller ionization
 - iii. Ionization energy depends on the effective nuclear charge and the average distance of the electron from the nucleus
 - iv. Irregularities
 - 1. Ionization energy is decreased via going down rows when you are adding a P orbital to an S orbital. The P orbital is often at a higher energy level in comparison to the S orbital, so it is held less tightly
 - 2. Ionization energy is decreased when going down rows when there is are 4 electrons in a P orbital. There is repulsion in the paired electrons of the P orbital
- d. Electron Config of Ions
 - i. When writing these just take off electrons from the highest energy level
 - 1. For iron, if you take electrons off, take them off from the S orbital (instead of the D orbital)
 - ii. When adding electrons, add them to the lowest energy level (n)

7.5 Electron Affinities

- a. Intro
 - i. Most atoms can gain electrons to form ions
 - ii. Electron affinity
 - 1. Measures the attraction of the atom of an added electron
 - 2. Energy change that occurs when an electron is added to an isolated, gaseous atom
 - 3. This number is generally negative, in comparison to ionization energy, which is positive
- b. Trend
 - i. This has the exact same trend as for ionization energy
 - ii. Halogens have the most negative electron affinities because they want to achieve noble gas state
 - iii. Metals tend to have slightly negative or even positive electron affinities
 - iv. Noble gasses have a value of ~0 and there are a bunch of exceptions
 - v. Electron affinities do not change greatly down a group

7.6 Metal/nonmetal/metalloid

- a. Intro
 - i. Metallic behavior tends to increase when going down a group
 - ii. Metallic behavior tends to decrease when going across a period
 - iii. Metallic behavior
 - 1. The characteristic of metal
- b. Metals
 - i. These are shiny, conduct heat and electricity, malleable (hammered into sheets, and ductile (thin wires)
 - ii. Generally always solids
 - iii. Metals have low ionization energies and tend to form positive ions easily
 - iv. Compounds of metals with nonmetals tend to be ionic substances
 - v. Most metal oxides are basic
 - 1. React with water to form metal hydroxides
- c. Nonmetals
 - i. These vary a lot
 - ii. They are not lustrous and are poor conductors of heat and electricity
 - iii. Generally have lower melting points than those of metals
 - iv. Many exist as diatomic molecules
 - v. Nonmetals tend to gain electrons when they react with metals (since they have high electron affinity)
 - vi. Compounds composed entirely of nonmetals are molecular substances
 - vii. Most nonmetal oxides are acidic
- d. Metalloids
 - i. Have a mix of both metal and nonmetal characteristics

7.7 Group Trends for the Active Metals

- a. Group 1A: The Alkali Metals
 - i. These are silvery, metallic luster, and high thermal and electrical conductivity
 - ii. Low densities and melting points
- b. Group 2A: The Alkaline Earth Metals
 - i. These are harder, denser, and have a higher melting point in comparison with the alkali metals
 - ii. Also less reactive (higher IE)

7.8 Group Trends for nonmetals

- a. Hydrogen
 - i. Diatomic gas
 - ii. Pretty high ionization energy, but still loses its electron really frequently
- b. 6A: Oxygen Group
 - i. Down the group, there is a change from non metallic to metallic character
 - ii. Ozone: O₃
 - 1. Toxic
- c. 7A: The Halogens

- i. Their melting and boiling points increase with increasing atomic number
- d. 8A: The noble gasses
 - i. They are all nonmetals that are gases at room temperature

<http://www.chemmybear.com/apch07sc.pdf>

<http://www.chemmybear.com/stdycrds.html>

UNIT 1 Test pog?

Adklajflk so bad xdddasjdfkajsflksjxl

Chapter 8

8.1 Chemical Bonds, Lewis symbols, and the octet rule

- a. Intro
 - i. Chemical bonds
 - 1. Whenever 2 atoms or ions are strongly attached to each other
 - ii. Ionic bond
 - 1. Electrostatic forces that exist between ions of opposite charge
 - a. Generally between substances of the left and right sides of the periodic table
 - iii. Covalent bond
 - 1. Sharing of electrons between 2 atoms
 - 2. Nonmetallic substances
 - iv. Metallic bonds
 - 1. Found in metals, with each atom bonded to several neighboring atoms
- b. Lewis Symbols
 - i. Consists of the chemical symbol for the element plus a dot for each valence electron
 - 1. Each side of the symbol can have 2 electrons
- c. The Octet Rule
 - i. Atoms tend to lose, gain, or share electrons until they are surrounded by 8 valence electrons
 - 1. Consists of a full S and P subshell in an atom

8.2 Ionic Bonding

- a. Intro
 - i. NaCl is made up of Na^+ and Cl^- ions
 - 1. Na gives an electron to the Cl
- b. Ionic bond:
 - i. Typically form between element with low ionization energy (which would easily lose an electron) and an element with high electron affinity (which would readily gain an electron)
- c. Energetics of Ionic Bond Formation
 - i. Ionic bond formation is very exothermic
 - 1. For NaCl:
 - a. A electron is removed from Na (endothermic)
 - b. An electron is added to Cl (exothermic)
 - c. Putting the 2 together is highly exothermic
 - ii. Lattice energy
 - 1. The energy required to completely separate a mole of a solid ionic compound into its gaseous ions
 - 2. For a given arrangement of ions, the lattice energy increases as the charges on the ions increase and as their radii decrease
 - 3. Electron configurations of ions of the representative elements
 - a. Addition of electrons to nonmetals is either exothermic or slightly endothermic
- d. transition -metal ions
 - i. Lattice energy of ionic compounds are large enough to compensate for the loss of up to 3 electrons from atoms
 - ii. Many transition metals can give off a variable amount of electrons
- e. Polyatomic ions
 - i. Common ions (generally bonded by covalent bonds)
 - ii. The group of atoms as a whole acts like a charge species when the ion forms an ionic compound with an ion of opposite charge
- f. Electron configurations of ions of the representative elements
 - i. (1) Group 1 elements form +1 ions because loss of additional electrons from an inner shell requires a very large amount of energy;
 - ii. (2) Group 2 elements form +2 ions; group 13 elements form +3 ions
 - iii. (3) Group 14 metals usually form +2 ions
 - 1. (a) Group 15 elements typically form -3 ions
 - 2. (b) Group 16 elements typically form -2 ions
 - 3. (c) Group 17 elements typically form -1 ions
- g. i) Transition Metal ions
 - i. (1) Transition metals generally do not form ions that have a noble-gas configuration.
 - ii. (2) Transition metals form the valence shell s electrons first, than as many d electrons as are required to reach the charge of the ion.
- h. ii) Properties of Ionic Compounds
 - i. (1) Brittle, with high melting points
 - ii. (2) Crystalline structure—rigid, well defined 3D arrangement

8.3 Covalent Bonding

- a. Ionic substances
 - i. Brittle, high melting points
 - ii. Crystalline (flat surfaces that make characteristic angles with one another)
- b. Covalent bonds
 - i. H₂
 - ii. i) Properties of Covalent Materials
 - 1. Low melting points, may vaporize readily; may be pliable in their solid forms
 - 2. Both [repulsive & attractive forces](#) are present
 - 3. Attractions between the nuclei & the electrons cause the electron density to concentrate between the nuclei leads to a net attractive force
 - 4. If the nuclei are too close, the repulsions will be stronger and the atoms will separate
- c. Multiple Bonds
 - i. Single bond- single covalent bonds
 - ii. Double bonds- there are 2 electron pairs shared between 2 elements
 - iii. Triple bonds
 - 1. The more bonds, the shorter the **bond length** between the 2 atoms

8.4 Bond Polarity and Electronegativity

- a. Intro
 - i. When there are 2 of the identical atoms in a bond, the electrons must be shared equally
 - ii. In ionic compounds there is no sharing of electrons
 - iii. Bond polarity
 - 1. Describes the sharing of electrons between atoms
 - iv. Non polar covalent bond
 - 1. Electrons are shared equally
 - v. Polar covalent
 - 1. One atom has greater electronegativity
- b. Electronegativity
 - i. Defined as the ability of an atom in a molecule to attract electrons to itself
 - 1. The greater it is, the greater its ability to attract electrons
 - ii. Trends
 - 1. Within a period electronegativity increases left to right
 - 2. Decreases down a group
- c. Electronegativity and bond polarity
 - i. Finding the difference in electronegativity can help see the degree of polarity of the bonding between them
 - 1. The more electronegative element will have a negative charge (δ^-)
- d. Polar, nonpolar, ionic bonding
 - i. (1) Nonpolar bond = very small difference in EN
 - ii. (2) Polar bond = small to moderate differences in EN result in δ^+ and δ^- regions of bond
 - iii. (3) Ionic bond = large difference in EN results in electron transfer
- e. Dipole Movements
 - i. The difference in electronegativity in a polar bond creates a polar molecule
 - ii. One element will have a positive dipole and the other will be a negative dipole

- i) Remember to name molecules properly!
 - (1) Ionic compounds: use Stock notation NO PREFIXES!
 - (2) Covalent molecules: use prefixes
 - (3) Note that compounds containing metals with high oxidation numbers (usually above 3+) often have properties more similar to molecular compounds than ionic compounds!

8.5 Drawing Lewis structures IMPORTANT LKADJLAKD

a. Rules

- i. Sum of all the valence electrons, add one for a negative ion, decrease one for a positive ion
- ii. Write a skeleton structure, placing the least electronegative (not H though) in the middle
- iii. Use dashes to connect bonded atoms. Each dash is a bonded shared pair
- iv. Complete octets around all the atoms bonded to the central atom
- v. Place leftover electrons in the central atom
 - 1. N, O, C, F always follow this rule
 - 2. P and S might have expanded octets, while B might be electron deficient
- vi. If there are not enough electrons to give the central atom an octet, use multiple bonds

b. Formal charge

- i. The charge of the atom would have if all the atoms in the molecule had the same electronegativity
 - 1. $FC = \text{valence electrons} - \text{lone pair electrons} - \text{number of bonds}$
 - a. In general, choose the lewis structure in which the atoms have formal charges closest to zero
 - b. Choose the lewis structures which places any negative formal charges on the most electronegative atoms
- ii. Formal charges do not represent the real charges on atoms

8.6 Resonance structures

- a. Sometimes a single lewis structure is inadequate for describing a molecule or polyatomic ion
 - i. Sometimes it is possible to draw several correct, equivalent structures for the same molecule, in which only the placement of the electron differs
 - ii. These structures are called resonance structures, and they are separated by double-headed arrows
- b. The "true" structure is somewhere in the middle- a blend of the equivalent structures
- c. In some cases, the possible lewis structures for a species may not be equivalent to each other; instead, one or more may represent a more stable arrangement than the others
- d. Benzene is a resonance structure
 - i. It is an aromatic molecule C_6H_6
 - ii. There are alternating single/double bonds, so all C-C bond lengths are the same with a bond order of 1.5

8.7 Exceptions to the Octet rule

- a. Some molecules have an odd number of electrons
 - i. You can't really draw a lewis structure of these that well
- b. Electron-deficient molecules
 - i. Less than an octet around the central atom

1. B, Be
 2. BF₃
- c. Expanded Octets
- i. More than an octet around the central atom
 1. P and S
 2. Elements must be in period 3 or lower in order to have expanded octets
 - ii. Old viewpoint: the central atoms seem to be using their empty d orbitals to accommodate additional electrons
 1. Current thinking: computational data don't seem to support this idea of involving d orbitals. (DON'T HAVE TO KNOW)
 - iii. The larger the central atom, the larger the number of atoms that can surround it. Small surrounding atoms (F, Cl, and O) are frequently found in expanded octet situations
- d. Sometimes Lewis structures are written with an expanded octet even though structures can be written with an octet
- i. Rule of thumb: go with an expanded octet if it reduces formal charges but in general satisfy the octet rule if possible

8.8 Strengths of Covalent Bonds

- a. Bond Enthalpy: the enthalpy change for the breaking of a particular bond in one mole of a gaseous substance
 - i. $\Delta H_{(\text{bond type})}$ information is tabulated.
 - ii. Often, average bond enthalpies are used
 1. Aka bond strength or bond energy
 - iii. Bond enthalpy is always a positive quantity (energy is always needed to break bonds)
 - iv. The greater the enthalpy, the stronger the bond
 - v. A molecule with strong chemical bonds is less reactive than one with weak bonds
- b. Bond enthalpies can be used as a means to calculate the enthalpies of reactions
 - i. An application of Hess's law!
 - ii. Enthalpy change = (enthalpies of bonds broken) - (enthalpies of bonds formed)
- c. As the number of bonds between two atoms increases, the bond becomes shorter.
 - i. Single bonds are longer than double bonds, etc

Chapter 9

9.1 Molecular Shapes

- a. Intro
 - i. Bond angles
 - 1. The angles made by the lines joining the nuclei of the atoms in the molecule
 - ii. Linear: 180
 - iii. Trigonal planar: 120
 - iv. Tetrahedral: 109.5
 - v. Trigonal bipyramidal: 90 AND 120
 - vi. Octahedral: all 90

9.2 The VSEPR Model

- a. VSEPR or valence shell electron-pair repulsion
- b. Intro
 - i. Bonding pair
 - 1. Region in which the electrons will most likely be found in
 - ii. non bonding pair (lone pair)
 - 1. Defines an electron domain that is located principally on one atom
 - iii. The best arrangement of a given number of electron domains is the one that minimizes the repulsions among them
 - iv. Electron domain geometry
 - 1. The arrangement of electron domains about the central atoms
 - v. Molecular geometry
 - 1. Arrangement of only the atoms in a molecule or ion
- c. Using vsepr to estimate the shapes of molecules
 - i. Draw the lewis structure of the molecule or ion and count the total number of electron domains around the central atom
 - ii. Describe the electron-domain geometry by arranging the electron domains about the central atom so that the repulsion among them are minimized
 - iii. Use the arrangement of bonded atoms to determine the molecular geometry
- d. The effect of nonbonding electrons and multiple bonds on bond angles
 - i. Tetrahedral bond orders:
 - 1. 109.5, 107, 104.5
 - ii. Nonbonding pairs experience less nuclear attraction, so they are more spread out in comparison to bonding pairs.

1. Therefore, electron domains for nonbonding electron pairs exert greater repulsive forces on adjacent electron domains and thus tend to compress the bond angles
- iii. Furthermore, domains for multiple bonds will also exert a greater repulsive force on adjacent electron domains
- e. Molecules with expanded valence shells
 - i. These deal with the atoms with 5 or 6 bonding domains
 - ii. Axial positions
 1. Point up and down
 - iii. Equatorial positions
 1. Point to forward and to the sides :P
 - iv. Axial and equatorial positions are shown in a trigonal bipyramidal
- f. Shapes of the larger molecules
 - i. There are molecules with more than one central atom
 1. Use VSEPR model for each individual atom
- g. Why does VSEPR happen?
 - i. The bonds want to be in a location that minimizes the repulsion between the electron domains

just know the handout thing forester gave out

9.3 Molecular shape and molecular polarity

- a. Intro
 - i. Bond polarity
 1. Difference in how electrons in a bond are shared
 - ii. The dipole moment depends on both the polarities of the individual bonds and the geometry of the molecule
 - iii. Bond dipole
 1. The dipole moment that is due only to the two atoms in that bond
 - a. This can be drawn as a bond vector
- b. As the electronegativity difference between the atoms in the bond increases, the dipole moment of the bond increases
- c. For a molecule containing more than two atoms, the dipole moment of the molecule depends on the polarity of the bonds and their geometric arrangement
 - i. The overall dipole moment of a molecule is the vector sum of its bond dipoles
 - ii. CO₂ is nonpolar, even though the CO bond is polar
 1. The 2 dipoles cancel each other out as the sum of the vectors is zero
 - iii. H₂O is polar, because the bond dipoles reinforce each other
 - iv. For symmetrical shapes, if all the atoms bonded to the central atom are identical, there will be no dipole moment

9.4 Covalent bonding and orbital overlap

- a. Valence bond theory
 - i. Bonds form when a valence atomic orbital of one atom overlaps with that of another atom

1. The orbitals share a region of space
2. Now electrons of opposite spin can share the common space between the nuclei. Electrons are simultaneously attracted to both nuclei
- ii. The optimum distance between the nuclei puts the system at a potential energy minimum
 1. Too far apart: no attractive forces
 2. Too close together: lots of repulsion between the nuclei

9.5 Hybrid orbitals

- a. Hybrid orbitals
 - i. A way to reconcile valence bond theory and observed molecular geometries
- b. Assume that atomic orbitals on an atom mix to form new "hybrid orbitals" in a process called hybridization

Hybridization	Electron domain geometry	Bond angles
sp	linear	180
sp ²	Trigonal planar	120
sp ³	tetrahedral	109.5
sp ³ d	Trigonal bipyramidal	
sp ³ d ²	octahedral	

Domain: paired electron/ nonbonding electron

OneNote is Better.