CHEM 1110 Lecture Notes

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COURSE ADMINISTRATIVE DETAILS

- My office hours

- Intro to my research

- Introductory Quiz

- Grading details

  - Exams - 40, Final - 15, Quizzes - 15, Achieve Homework - 15, Textbook Homework - 15
  - Online Homework
  - Textbook Homework
  - Daily quizzes

- Importance of reading and learning on your own

- Learning resources

  - My Office Hours
  - PAL group
  - Tutoring services - https://www.suu.edu/academicsuccess/tutoring/

- Show how to access Canvas

  - Calendar, Grades, Modules, etc.
  - Achieve Homework
  - Zoom, etc.

CHAPTER 1

MATTER AND MEASUREMENTS

1.1 Chemistry: The Central Science

- My demos (Ruby, Periodic Table, NaAc hand warmers, Electroplating)
- Chemistry is the "Central Science" because it ties in to other disciplines
- Figure 1.1 – Web of disciplines
- Physical and chemical properties
- Physical and chemical changes
  - Pop balloon vs explode balloon

1.2 States of Matter

- Three phases of matter
- State changes

1.3 Classification of Matter

- Figure 1.3 – flow chart
- Pure substances vs mixtures
- Homogeneous vs heterogeneous mixtures
- Compounds vs Elements
- Physical and chemical changes on the flow chart
- Anatomy of a chemical reaction

1.4 Chemical Elements and Symbols

- Elements are given names and symbols (Some, like Fe, don't match)
- Table 1.2 – Names and symbols of common elements
- Chemical formulas – Type and amount of each element in a compound

1.5   Elements and the Periodic Table

- Metals, non-metals, and metalloids (There are further divisions we will get to later)
- Table 1.4 – Elements essential to human life
- Homework 1.1:
    - 1.38 – Physical vs. Chemical changes
    - 1.41 – Changes of state
    - 1.44 – Mixtures vs. Pure substances
    - 1.46 – Compounds vs. Elements
    - 1.55 – Counting atoms in a chemical formula
    - 1.57 – Classify metals, nonmetals, and metalloids

1.6   Chemical Reactions: An Example of Chemical Change

- Demo: Zinc in copper nitrate solution

1.7   Physical Quantities

- Physical quantities are things that can be measured or calculated
- SI units (Table 1.5)
- Metric prefixes and 10^n conversions (Table 1.6)
    - Each prefix is a multiplication operation
    - Why would you use prefixes?
    - $100m$, $0.1km$, and $10,000cm$
- Derived units
    - Squared and cubed units (cm^3)
    - Combined units (m/s)
- Using units to check your formulas: $B_e = \dfrac{h}{8\pi^2 \mu R_e^2 c} = cm^{-1}$

1.8   Measuring Mass, Length, and Volume

- Weight vs mass
- kg (SI unit) vs g (More convenient in the lab)
- SI vs other systems (Tables 1.7, 1.8, and 1.9)

1.9 Measurement and Significant Figures

- ○ Accuracy in measurements

- ○ Statistics (σ) in many measurements, simple rules for one measurement

- ○ Consider my height: $6ft$ vs $74in$ vs $73.784672in$

    - · More digits implies a more precise measurement
    - · The first one is correct, if we allow that it is just approximate
    - · The last one is only appropriate if I was *actually* measured to the millionth of an inch
    - · If I were precisely $6ft$ (with high precision), how would I write that? ($6.000ft$)

- ○ Significant figures

    - · Zeros in the middle of two numbers are significant
    - · Zeros at the beginning of a number are never significant
    - · Zeros at the end of a number *after* the decimal point are significant
    - · Zeros at the end of a number *before* the decimal point are not significant

- ○ We should keep track of the total number of significant figures, and the position of the *least* significant figure in a number

- ○ Sig-fig practice: Give # of SFs and LSD position for several numbers

1.10 Scientific Notation

- ○ Expressing numbers in scientific notation

    - · Split the number into *quantity* and *magnitude*
    - · Express the quantity as a number with MSD in the ones place
    - · Express the magnitude as a power of 10

- ○ Scientific notation serves two purposes:

    - · It allows for easy expression of very large and very small numbers
    - · It makes it easy to express precision (every 0 in scientific notation is significant)
    - · Consider $2,300kg$ measured to $\pm1kg$ ($2.300 \times 10^3 g$)

- ○ Scientific notation and metric prefixes

    - · For metric to scientific: ($23.6kg \rightarrow 2.36 \times 10^4 g$)
      First convert the metric prefix to a power of 10, then shift the decimal of the quantity to put its MSD in the ones place
    - · For scientific to metric: ($5.60 \times 10^{-7}m \rightarrow 560nm$)
      First shift the decimal to make the power of 10 match a prefix, then replace that power of 10 with the prefix

1.11   Rounding off Numbers

○ Error analysis for addition and subtraction

    · The LSD of the answer will match the LSD of the least precise input

    · # of SFs doesn't matter, and the answer might gain or lose SFs compared to the inputs

    · Practice: $96g + 43g = 139g$ and $23.4g - 18.6g = 4.8g$

    · Consider adding $1.00\mu l$ of water to a $8L$ Home Depot bucket – The precision of the final amount is entirely determined by the low precision of the bucket

○ Error analysis for multiplication and division

    · The # of SFs in the answer will match the smallest # of SFs from the inputs

    · Here, place value does not matter at all

    · Practice: $45.7g \div 8.2cm^3 = 5.6 ^g/_{cm^3}$ and $82.5^{miles}/_{hour} \times 53.24 hours = 4,390 miles$

○ Compound problems

    · Solve compound problems step-by-step, writing the intermediate answers

    · For each intermediate answer, keep track of the # of SFs and LSD position

    · Only round the final answer, don't round intermediates

    · Practice: $\frac{12.3g + 34g}{12.0cm^3 + 7.7cm^3} = 2.4 ^g/_{cm^3}$ (wrong answer with premature rounding)

1.12   Problem Solving: Unit Conversions and Estimating Answers

○ Conversion factors

    · Rearrange conversion equation to make unity ratios: $1 = \frac{1in}{2.54cm} = \frac{2.54cm}{1in}$

    · Multiply a number by the right ratio to cancel out the starting units and leave the converted units

    · "Do I multiply or divide by the conversion factor?" is the wrong question. You are always multiplying by a ratio. The right question is "Where should the units be in my conversion factor?" The numbers go along with the units

    · Most conversion factors have "1" with infinite significant figures, and the other number limits SFs in a calculation

    · Some conversions are defined as *perfect* (such as $60s = 1min$) and won't limit SFs at all

    · Take care to square or cube the first-power ratio for square or cubed dimensions: $1^3 in^3 = 1in^3 = (2.54)^3 \, cm^3 = 16.4cm^3$

○ Dimensional analysis is a method which frames any problem (even complex ones) in terms of multiple unit conversions

- The railroad ties or picket fence method makes it easy to organize a problem
- Set up all the unit conversions before doing any calculations
- Sort out significant figures at the end
- Practice: $\frac{65.3miles}{h} \left| \frac{1h}{60min} \right| \frac{1min}{60s} \left| \frac{1.61km}{1mile} \right| \frac{1000m}{1km} = 28.4^m/_s$

  ○ Homework 1.2:

- 1.60 – Units and metric prefixes
- 1.64 – Scientific Notation
- 1.66 – Counting Significant Figures
- 1.70 – +- error propogation
- 1.71 – */ error propogation

1.13   Temperature, Heat, and Energy

○ All chemical reactions involve a change in energy

○ Energy is the capacity to supply heat or do work. It is in units of J

○ Temperature is the measure of heat energy in an object

- We don't use F (Booooooo, Fahrenheit!)
- $°C$ is split so that $0°C$ is freezing and $100°C$ is boiling
- Absolute temperature, K, has spacing equal to $°C$, but starts at absolute o
- $T(°C) = T(K) - 273.15$

○ The energy needed to raise the temperature of a substance is its specific heat (Table 1.10)

○ $1cal = 4.184J$, based on the heat capacity of water: $4.184^J/_{gK}$

○ $q = mC\Delta T$ (practice problems with q or ΔT as the unknown)

1.14   Density and Specific Gravity

○ $density = \dfrac{mass}{volume}$ (Table 1.11), and can be used to convert mass and volume

○ Density can be affected by temperature changes and phase changes. Water exhibits anomalous behavior in this respect

○ Specific gravity relates the density of a substance to the density of water

○ Homework 1.3:

- 1.82 – Dimensional analysis word problem
- 1.85 – Heat transfer equation
- 1.92 – Density measurement

CHAPTER 2

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ATOMS AND THE PERIODIC TABLE

2.1 Atomic Theory

- DEMO – How many times can we subdivide a crystal of salt and have it still be salt?

- Postulates of atomic theory:

    - All matter is composed of atoms
    - The atoms of a given element differ from the atoms of all other elements
    - Chemical compounds consist of atoms combined in specific ratios (H_2O vs H_2O_2)
    - Chemical reactions change the way that atoms are combined in compounds, but leave the atoms themselves unchanged

- Today we know that atoms are very small ($m = 10^{-23}g$ and $r = 10^{-10}m$)

- Atomic mass unit – ^{12}C has a mass of *exactly* $12 amu$

- Subatomic particles and their interactions

- Atomic structure – Nucleus and empty space

2.2 Elements and Atomic Number

- Atomic number (Z) is the number of protons and it distinguishes one element from another

- Neutral atoms will have the same number of electrons as protons

- Mass number (A) is the number of protons *and* neutrons

2.3 Isotopes and Atomic Weight

- Atoms with the same atomic number, but different mass numbers are called isotopes

- Atomic symbols: ^3_1H

- Hydrogen, deuterium, and tritium

- Atomic weight – Based on natural abundances of the isotopes

- ^{6}Li: 7.59% abundance and $m = 6.015122 amu$. ^{7}Li: 92.41% abundance and $m = 7.016004 amu$

- ○ Homework 2.1:

    - · 2.35 – How elements differ from each other
    - · 2.48 – Atomic symbols and isotopes
    - · 2.52 – Calculate atomic mass from isotopes

2.4   The Periodic Table

- ○ As more elements were discovered, scientists tried to find commonalities and patterns
- ○ Li, Na, and K were similar to each other, as were Cl, Br, and I
- ○ Mendeleev arranged the elements in increasing mass, and saw a pattern
- ○ Mendeleev even correctly predicted the properties of Gallium, Germanium, and Scandium – yet-undiscovered elements
- ○ Rows are called "Periods" and columns are called "Groups"
- ○ The periods get longer further down due to electronic structure – We will talk about this soon
- ○ Main group vs. transition metals vs. inner transition metals

2.5   Some Characteristics of Different Groups

- ○ Periodic trend in atomic radius
- ○ Group 1A – Alkali metals
- ○ Group 2A – Alkaline earth metals
- ○ Group 7A – Halogens
- ○ Group 8A – Noble gases

2.6   Electronic Structure of Atoms

- ○ Electronic structure largely determines physical and chemical properties
- ○ Electrons can only exist in quantized energy levels because of their quantum wavelike nature
- ○ The energy level of an electron can be characterized by four *quantum numbers*
- ○ First is the shell number. Each shell can hold more electrons than the last
- ○ Next is the subshell – s, p, d, and f subshells

◦ Orbitals make up subshells, and each orbital holds 2 electrons

◦ Table 2.2 shows the shells and subshells

◦ Spin differentiates the two electrons which share an orbital

2.7   Electron Configurations

◦ Electrons occupy orbitals according to the following rules:

· Electrons will occupy the lowest energy orbital available. This is the Aufbau principle. The ordering follows a pattern of s, p, d, and f within a shell, but there is some crossover between shells, as shown in figure 2.6

· Each orbital can hold only two electrons, which must have opposite spins. This is the Pauli exclusion principle

· Degenerate groups of orbitals are all filled halfway with electrons of the same spin before any electrons are paired up. This is Hund's rule

◦ Table 2.3 shows electron configurations for the first 20 elements

◦ Drawing electron configuration diagrams with harpoons for electrons

◦ Noble gas notation

2.8   Electron Configurations and the Periodic Table

◦ The periodic table is arranged according to properties, but properties are governed by electronic configuration. Ergo, the periodic table is arranged according to electronic configuration!

◦ Figure 2.7 shows how the blocks of the periodic table represent subshells in electronic structure

◦ Core electrons vs. valence electrons

2.9   Electron Dot Symbols

◦ It is useful to graphically represent the valence electrons with electron dot symbols

◦ Write the symbol for the element, then surround it with the correct number of dots to represent valence electrons

◦ Dots go on the four sides of the symbol, and pair up only when there are more than 4 electrons

◦ Practice Problems

· 2.60 – Classify elements (metal, non-metal, main group, etc.)

· 2.62 – Reactive trends in the periodic table

- 2.70 – How many subshells in each shell
- 2.74 – Drawing electronic structure diagrams
- 2.77 – Electronic configurations of elements

CHAPTER 3

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IONIC COMPOUNDS

3.1 Ions

○ Metals and non-metals will combine to form ionic compounds

○ Ionic compounds have high melting points, stable crystal structures, and a degree of water solubility

○ When ionic compounds dissolve in water, they conduct electricity

○ Positive ions are called *cations* and negative ions are called *anions*

○ Many reactions involve ions which are already charged, while other reactions generate ions through gaining or losing electrons

3.2 Periodic Properties and Ion Formation

○ Ionization energy is the energy required to remove an electron and create a cation

○ Electron affinity is the energy given off when an electron is gained, making an anion

○ Periodic trends in ionization energy and electron affinity

○ Nonmetals such as C and N tend to not form ions

3.3 Ionic Bonds

○ Ionic compounds are held together by ionic bonds

○ Positive cations are attracted to negative anions

○ Ionic solids exhibit a regular crystal lattice on atomic scales

3.4 Some Properties of Ionic Compounds

○ Ionic compounds have different lattice structures, depending on the sizes and charges of the ions

○ Ionic solids are rigid but brittle, easily cleaved along crystal planes

○ Ionic compounds have high melting and boiling points – This is because ionic bonds are very strong

○ Some ionic compounds dissolve in water by making many attractive interactions with water dipoles. Others are just bound too tightly, and are insoluble.

3.5   Ions and the Octet Rule

○ Elements will tend to gain or lose enough electrons to have a full *octet* of valence electrons

○ This means that non-metals will gain electrons to reach the electron configuration of the next highest noble gas

○ Metals will lose electron to reach the next lowest noble gas

○ Electron configurations for ions are written just the same as electron configurations for elements

3.6   Ions of Some Common Elements

○ Use the periodic table to predict the charge an element will take on as it forms an ion

○ Group 1A will be 1+

○ Group 2A will be 2+

○ Group 6A will be 2-

○ Group 7A will be 1-

3.7   Naming Ions

○ For metals with just one ion (1A, 2A, and 3A groups), just add "ion" to the end of the name

○ For metals with multiple ions (transition metals, 4A, and 5A), state the charge and then add "ion"

○ An old rule uses different endings (ferrous and chromous for $2+$, and ferric and chromic for $3+$)

○ Anions change the end of the element name to "-ide"

○ See tables 3.1 and 3.2

○ Practice Problems

· 3.38 – Ion formation by gaining/losing electrons

· 3.46 – Ions and the octet rule

· 3.50 – Ionization energy

· 3.58 – Naming Ions

3.8 Polyatomic Ions

○ Polyatomic ions are composed of more than one atom

○ These polyatomic ions are a covalently bound group, and should be considered as a single ion

○ See table 3.3

○ There are many series of polyatomic ions which differ in only the number of oxygen atoms

○ There are also series of polyatomic ions which differ in hydrogen atoms and charge (each H increases the charge by 1)

3.9 Formulas of Ionic Compounds

○ To find the formula for an ionic compound, we need to balance the charges to give a neutral compound

○ The ratio should be in a mathematically reduced form, making just one *formula unit*

3.10 Naming Ionic Compounds

○ Simply combine the two ion names – Cation first, then anion

○ Do not state the ratio between them, like you do in molecular compounds

○ Table 3.4 – Some common ionic compounds and their applications

3.11 H^+ and OH^- Ions: An Introduction to Acids and Bases

○ The hydrogen cation (really just a bare proton) and hydroxide anion (OH^-) are very important to aqueous chemistry

○ When an acid dissolves in water, it donates a hydrogen cation to a water molecule forming H_3O^+

○ When a base dissolves in water, it will accept a proton from water forming OH^-

○ Some acids can provide multiple H^+ in reactions (HCl, H_2SO_3, H_3PO_4), and some bases can accept multiple protons ($NaOH$, $Ba(OH)_2$)

○ Practice Problems

   · 3.62 – Polyatomic Ions
   · 3.66 – Formulas from names for ionic compounds
   · 3.72 – Names from formulas for ionic compounds
   · 3.78 – Acid/base reactions with water

CHAPTER 4

MOLECULAR COMPOUNDS

4.1 Covalent Bonds

○ A bond formed by the *sharing* of electrons is called a covalent bond. They occur between non-metals

○ A group of atoms held together by covalent bonds is called a molecule (this can include compounds like CO_2 or elements like O_2)

○ Show a water molecule using dot diagrams

○ Repulsive and attractive forces in a covalent bond as a function of bond length

○ Overlap of 1s orbitals for H_2 and 2p orbitals for F_2

4.2 Covalent Bonds and the Periodic Table

○ The periodic table can be used to predict how many bonds an element tends to make

○ Consider HF, H_2O, NH_3, and CH_4

○ This pattern is the result of the octet rule. An atom makes as many bonds as it needs in order to reach a full octet

○ Elements in the third row and below can form more bonds. See Figure 4.3

4.3 Multiple Covalent Bonds

○ Some molecules cannot satisfy the octet rule using only single bonds

○ Consider F_2, O_2, and N_2

○ Single, double, and triple bonds share 2, 4, and 6 electrons

○ Multiple bonds in organic molecules

4.4   Coordinate Covalent Bonds

○ Coordinate covalent bonds are when both shared electrons came from one atom in the pair

○ Identify them in molecular compounds by counting the electrons and comparing them to their valency: N_2O

○ O_3 and NH_3BF_3 are two more examples of coordinate covalent compounds

○ In acid/base chemistry, all acidic hydrogens form coordinate covalent bonds

○ Practice Problems

  · 4.36 – Identify covalent bonds
  · 4.38 – Predict # of bonds from periodic table
  · 4.34 – Definition of a coordinate covalent bond

4.5   Characteristics of Molecular Compounds

○ Molecular compounds have a wide range of melting and boiling points, depending on their intermolecular forces (chapter 8)

○ Few molecular compounds are soluble in water, but many are soluble in organic solvents

○ Most molecular compounds do not conduct electricity, either as a pure substance or dissolved in water

4.6   Molecular Formulas and Lewis Structures

○ Molecular formulas tell the types and numbers of atoms found in a molecular compound

○ Structural formulas use lines to represent covalent bonds and show the structure of a molecule

○ Lewis structures add dots to represent lone pairs of electrons

4.7   Drawing Lewis Structures

○ The first way to draw Lewis structures works for most simple organic molecules

  · If the only atoms are C, H, O, X (Halogen), and H, then this approach might work
  · Arrange the carbon atoms in a chain, or backbone
  · Attach the heteroatoms with the appropriate number of bonds (based on the octet rule)
  · Make any multiple bonds in the carbon chain needed to complete their octets

· If all of the elements have a full octet, and form the correct number of bonds, then it is a valid structure

· Starting from a condensed structure makes this method particularly easy

○ The next method is very useful for molecular compounds with more complex structure, or central atoms which aren't carbon

1. Count up the total number of valence electrons (taking into account any net charge)

2. Draw single bonds from the central atom to all peripheral atoms

3. Fill the octets of the outer atoms with lone-pairs (H doesn't need any)

4. Place any remaining electrons onto the central atom

5. Satisfy the octet rule for the central atom, if needed, by converting outer lone-pairs into multiple-bonds with the central atom

○ Practice: PCl_3, HCN, SO_4^{2-} (sulfate has single bonds for this class)

4.8   The Shapes of Molecules

○ We know that molecules have a particular shape, defined by VSEPR theory

○ We can predict the molecular geometry based on a molecule's Lewis structure

1. Draw a valid Lewis structure

2. Count the number of electron domains

3. Count the number of bonds vs. lone pairs

4. Consult Table 4.2 (or your memory)

○ Bond angles can range from $180°$ to $< 109.5°$

○ If lone pairs are present, they slightly reduce the bonding angles

○ Practice Problems

· 4.54 – Condensed structural formulas

· 4.58 – Drawing Lewis structures

· 4.70 – Molecular geometry

4.9   Polar Covalent Bonds and Electronegativity

○ Any covalent bond between different atoms will share electrons unequally

○ This unequal sharing leads to a polar bond

○ The direction and intensity of the bond polarity is determined by comparing electronegativities of the two elements

○ F is the most electronegative, while Rb is the least

○ Electronegativity differences of 0-0.4 are considered normal covalent bonds

○ Electronegativity differences of 0.5-1.9 are considered polar covalent bonds

○ Electronegativity differences greater than 2 are considered ionic bonds

4.10   Polar Molecules

○ Polar bonds in a molecule can lead to a charge dipole over the entire molecule

○ First find the polar bonds, then add up their dipole vectors to find the net dipole

○ Symmetry can often lead to a non-polar molecule despite having polar bonds

4.11   Naming Binary Molecular Compounds

○ A compound made of only two types of atoms is called a binary compound

○ When naming a binary compound, the less electronegative atom usually comes first

○ The second element will have an "-ide" ending, just like in ionic compounds

○ Use prefixes (Table 4.3) to indicate how many of each element are present

○ Practice Problems

· 4.78 – Polar bonds
· 4.82 – Polar molecules
· 4.86 – Names from formulas
· 4.88 – Formulas from names

CLASSIFICATION AND BALANCING OF CHEMICAL REACTIONS

5.1 Chemical Equations

○ Chemical equations can be thought of as recipes which show what you start with and what you get out

○ Coefficients indicate how many units of each reactant are needed, and how many units of each product are made

○ We find the coefficients by relying on the law of conservation of mass

5.2 Balancing Chemical Equations

○ To balance chemical reactions:

· Write the unbalanced equation, with correct formulas for all reactants and products
· Identify an element which is unbalanced, and balance it with a coefficient
· Repeat the above step until all elements are balanced
· Reduce the coefficients if they share a common factor

5.3 Classes of Chemical Reactions

○ It is sometimes helpful to recognize different classes of chemical reactions

○ Precipitation reactions go from aqueous to solid:

$$Pb(NO_3)_2(aq) + 2\,KI(aq) \longrightarrow 2\,KNO_3(aq) + PbI_2(s)$$

○ Acid-base neutralization reactions result in water and a salt:

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

○ Oxidation-reduction reactions change the oxidation state (or charge) of atoms:

$$Mg(s) + I2(g) \longrightarrow MgI2(s)$$

5.4 Precipitation Reactions and Solubility

○ Solubility is a measure of how much a substance will dissolve in a solvent (usually water)

○ In precipitation reactions, you must determine the solubility of *all* potential products

○ Solubility rules (Table 5.1):

  · A compound is probably soluble if its cation is ammonium (NH_4^+) or group 1A cations (Li^+, Na^+, K^+, Rb^+, or Cs^+)

  · A compound is probably soluble if it contains one of the following anions:
  Halide ions (except when paired with Ag^+, Hg_2^{2+}, or Pb^{2+} ions)
  Nitrate, perchlorate, acetate, or sulfate (except for sulfates paired with Ba^{2+}, Hg_2^{2+}, or Pb^{2+} ions)

  · Most other ionic compounds are not soluble

  · Net ionic equation omit all species which are aqueous on both sides of the reaction

○ Practice Problems

  · 5.28 – Balancing chemical reactions
  · 5.38 – Identify reaction types
  · 5.44 – Solubility rules

5.5 Acids, Bases, and Neutralization Reactions

○ When a strong acid and base react, the product is *neutral* (neither an acid nor a base)

○ The products can include water and a neutral ionic compound

○ Some acid/base neutralization reactions produce gas: $2\,HCl(aq)\ +\ K_2CO_3(aq)\ \longrightarrow\ H_2O(l)\ +\ 2\,KCl(aq)\ +\ CO_2(g)$

○ Many bases yield OH^- as a product, even if they dont have hydroxide as a component themselves: $NH_3\ +\ H_2O\ \longrightarrow\ NH_4^+\ +\ OH^-$

5.6 Redox Reactions

○ Redox reactions involve changes in oxidation state, or the exchange of electrons

○ OIL-RIG – "Oxidation is losing, Reduction is gaining" electrons

○ Oxidation leads to a higher oxidation state, while reduction leads to a lower (reduced) oxidation state

○ Reduction and Oxidation always come together, since the electrons need a source and a destination

○ We use language which can be confusing at first:

  · The oxidizing agent is the chemical which oxidizes its partner. It is reduced itself in the process. It is a destination for electrons

  · The reducing agent is the chemical which reduces its partner. It is oxidized itself in the process. It is a source for electrons

○ Some redox processes we encounter in the world around us: Corrosion, Combustion, Respiration, Bleaching, Metallurgy

5.7 Recognizing Redox Reactions

○ Determining oxidation states can get complicated, but for now we can simply relate oxidation state to charge

○ Elemental states have oxidation states of 0

○ For a covalently bonded compound, the oxidation states sum up to the total charge

○ Elements in a covalent bond will usually take the same oxidation state that they would in an ionic bond (C is an important exception, as well as non-metal oxides)

○ Practice Problems

  · 5.39 – Net ionic equations for acid/base reactions
  · 5.54 – Labeling changes in oxidation number
  · 5.58 – Identify which reactant is oxidized and reduced

CHAPTER 6

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CHEMICAL REACTIONS: MOLE AND MASS RELATIONSHIPS

6.1   The Mole and Avogadro's Number

○ The molecular weight or formula weight of a substance is the sum of atomic weights for its constituent atoms

○ Different substances have different molecular weights, so $1.0g$ of C_2H_4 has more molecules in it than $1.0g$ of HCl

○ If we wanted to make C_2H_5Cl without any left-over reactants, then we would need more grams of HCl than of C_2H_4

○ Find the ratio by comparing the formula masses ($28.0amu$ vs $36.5amu$). If I have $1.0g$ of C_2H_4, how many grams of HCl do I need? ($1.3g$)

○ The mole is an amount that relates amus (atomic scale) to grams (macroscopic scale)

○ The molar mass tells how many grams of a substance is one mole of molecules

○ $N_A = 6.022 \times 10^{23}$, Avogadro's number, is the number of amus per gram

6.2   Gram-Mole Conversions

○ We can convert from grams to moles and vice-versa using the molar mass

○ How many moles is $2.5g$ of C_2H_5Cl? How many molecules?

6.3   Mole Relationships and Chemical Equations

○ A balanced chemical equation relates the ratios of moles of reactants and products

○ The coefficients tell how many moles are needed or produced

○ We call these ratios *stoichiometry*

○ Practice Problems

· 6.32 – Calculating molar mass

· 6.36 – Moles to mass and to molecules

· 6.39 – Stoichiometric ratios in a balanced equation

6.4 Mass Relationships and Chemical Equations

- ○ Mole to mole conversions rely on stoichiometric ratios (coefficients in a balanced chemical equation)

- ○ Mole to mass and mass to mole conversions rely on molar mass

- ○ Mass to mass conversions must follow the path: $mass_1 \longrightarrow moles_1 \longrightarrow moles_2 \longrightarrow mass_2$

- ○ You should always check your calculated answer against a rough estimate to make sure it makes sense

- ○ For $H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2\,H_2O(aq)$:

  - · How many moles of NaOH needed to react with $0.490g$ H_2SO_4?
  - · How many grams of water and sodium sulfate would be produced?

6.5 Limiting Reagent and Percent Yield

- ○ Unless the amounts of reactants are carefully chosen, one will run out first and the other will have left-over remaining after the reaction

- ○ To find the limiting reagent, calculate how much product each one would yield. The reactant which yields less is the limiting reagent and the amount calculated is the theoretical yield

- ○ To find percent yield, you must have an actual mass of product:

$$\%Yield = \frac{ActualYield}{TheoreticalYield} \times 100\%$$

- ○ $CO_2 + H_2O \longrightarrow C_6H_{12}O_6 + O_2$ (unbalanced)

  With $10.0g$ of CO_2 and $5.00g$ of H_2O, how many grams of sugar are produced?

- ○ if $4.8g$ of sugar are actually recovered, what is the % recovery?

- ○ $C_6H_8O_7 + NaHCO_3 \longrightarrow Na_3C_6H_5O_7 + H_2O + CO_2$ (unbalanced)

  With $1.0g$ citric acid and $1.0g$ sodium bicarbonate how many grams of $Na_3C_6H_5O_7$ are produced?

- ○ You actually recover $0.56g$ of citrate. What is the % yield?

- ○ Practice Problems

  - · 6.46 – Finding reactant amounts from desired product yield
  - · 6.52 – Limiting reactant problem

CHEMICAL REACTIONS: ENERGY, RATES, AND EQUILIBRIUM

Some chemical reactions can be balanced, but meaningless because they will never occur naturally.

$$2\,\mathrm{Au(s)} + 3\,\mathrm{H_2O(l)} \longrightarrow \mathrm{Au_2O_3(s)} + 3\,\mathrm{H_2(g)}$$
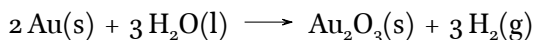
7.1 Energy and Chemical Bonds

○ Potential energy is stored energy. The energy of chemical bonds is a form of negative potential energy

○ Kinetic energy is energy of motion. Motion on the atomic scale is observed as heat

○ In a chemical reaction, energy is required to break chemical bonds, but energy is released when new bonds are formed

○ If the energy released is greater than the energy required to break the bonds, then the reaction produces heat and the products are more stable than the reactants

7.2 Heat Changes During Chemical Reactions

○ To predict if a reaction will give off heat, we must compare the energies of the bonds broken to the energies of the bonds formed

○ Table 7.1 shows average bond enthalpies

○ Consider the reaction: $\mathrm{H_2(g)} + \mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{HCl} - \Delta H_{rxn} \approx -43\dfrac{kcal}{mol}$

○ Since energy is conserved and the molecules lost energy overall, then that energy must have gone somewhere – It is released as heat!

○ Reactions which release heat are called "exothermic," and they occur more readily in nature

○ Reactions which take in heat (make the surroundings colder) are called "endothermic," and they are a bit harder to find in nature

○ Demo – Burning methanol and $\mathrm{Ba(OH)_2 \cdot 8H_2O(s)} + 2\,\mathrm{NH_4SCN(s)} \longrightarrow \mathrm{Ba(SCN)_2(aq)} + 2\,\mathrm{NH_3(g)} +$ 10 $\mathrm{H_2O(l)}$

○ We have a fancy name for the heat of reaction carried under constant pressure conditions: "Enthalpy of Reaction" (ΔH)

7.3 Exothermic and Endothermic Reactions

○ Practice an exothermic reaction: $CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(l) \quad \Delta H_{rxn} = -213\dfrac{kcal}{mol}$

○ Practice an endothermic reaction: $N_2(g) + O_2(g) \longrightarrow 2\,NO(g) \quad \Delta H_{rxn} = 43\dfrac{kcal}{mol}$

○ Note that these quantities are -per mole. Table 7.2 compares combustion enthalpies on a -per gram basis

7.4 Why do Chemical Reactions Occur? Free Energy

○ Chemical reactions which occur in nature are called "spontaneous"

○ Non-spontaneous reactions can be made to occur if driven by some external input of energy

○ Enthalpy is one governing factor in spontaneity, and entropy (S) is the other

○ Entropy can be understood as the amount of disorder in a system

○ A reaction will have $\Delta S_{rxn} > 0$ if the products are more disordered than the reactants

○ Entropy and Enthalpy are combined to form "free energy" (ΔG_{rxn}), which can be thought of as a measure of the spontaneity of a reaction

○ $\Delta G_{rxn} = \Delta H_{rxn} - T\Delta S_{rxn}$

○ $\Delta H_{rxn} < 0$ and $\Delta S_{rxn} > 0$ both trend toward spontaneity – Draw the quadrant diagram

7.5 How do Chemical Reactions Occur? Reaction Rates

○ ΔG_{rxn} and free energy only tell us if a reaction can happen, but doesn't say how quickly a reaction will occur

○ Some spontaneous reactions happen so slowly that they never really happen at all

○ The rate of a reaction depends on different factors, like how the reaction actually occurs (from a molecular perspective)

○ For reactions to occur, molecules must collide with enough energy and in the correct orientation (Figure 7.2)

○ The energy required to start a reaction is called the activation energy, and it limits the rate of reaction

○ Reaction coordinate diagrams can show factors which control both the energetics and the rate of a reaction(Figure 7.3)

- ○ Practice Problems

    - · 7.30 – Reaction enthalpies from bonds
    - · 7.34 – Predicting entropy changes
    - · 7.46 – Reaction coordinate diagrams

7.6 Effects of Temperature, Concentration, and Catalysts on Reaction Rates

- ○ Temperature:

    - · Increasing the temperature increases the frequency and energy of molecular collisions
    - · Both of these factors will increase the rate of any reaction

- ○ Concentration:

    - · Increasing the concentration increases the frequency of molecular collisions
    - · This will increase the rate of any reaction

- ○ Catalysts:

    - · Catalysts change the way a reaction proceeds, requiring less activation energy (Figure 7.4)
    - · This will increase the rate of reaction, since more collisions will have the necessary energy

7.7 Reversible Reactions and Chemical Equilibrium

- ○ Some chemical reactions can go in the forward or reverse direction dependig on the circumstances

- ○ These reactions are called equilibrium reactions because they establish an equilibrium state where both reactants and products are present

- ○ At equilibrium the forward and reverse reactions occur at the same rate, so the concentrations of products and reactants remains constant.

7.8 Equilibrium Equations and Equilibrium Constants

- ○ We can describe the equilibrium conditions for a reaction using math

- ○ The equilibrium constant: $K_{eq} = \dfrac{[M]^m[N]^n}{[A]^a[B]^b}$

- ○ The exponents in K_{eq} are from the coefficients in the balanced chemical equation

- ○ We can therefore find the equilbrium constant if we measure the equilibrium amounts of reactants and products

○ We could also find the equilbrium amount of one species if we know K_{eq} and the concentrations of all other species

○ The value of K_{eq} can indicate whether thee are more products or reactants at equilibrium

○ $K_{eq} > 1$ favors products, while $K_{eq} < 1$ favors reactants

7.9 Le Châtelier's Principle: The Effect of Changing Conditions on Equilibria

○ Le Châtelier's principle describes how a system will respond after being perturbed away from equilibrium

○ Effect of Changes in Concentration:

  · When concentrations are changed, the system will respond to counteract the change
  · Removing product will shift the reaction toward products
  · Adding product will shift the reaction toward reactants
  · etc.

○ Effect of Changes in Temperature

  · To understand the effect of a change in temperature we can consider heat as either a reactant or a product
  · Heat acts as a product in an exothermic reaction, and as a reactant in an endothermic reaction
  · Increasing the temperature is like adding heat, and decreasing the temperature is like removing heat
  · When temperature changes, the value of K_{eq} actually changes

○ Effect of Changes in Pressure (or volume):

  · First, count the number of moles on each side of the reaction (gas for pressure and aqueous for volume)
  · Increasing pressure or decreasing volume will shift the reaction toward the side with fewer moles
  · Decreasing the pressure increasing the volume will shift the reaction toward the side with more moles

○ Practice Problems

  · 7.48 – Catalysts and activation energy
  · 7.56 – Equilibrium constant
  · 7.68 – LeChâtelier's principle

CHAPTER 8

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GASES, LIQUIDS, AND SOLIDS

8.1 The States of Matter and Their Changes

○ The three primary phases of matter: Solid, Liquid, and Gas

○ Phase changes can go either direction at the phase change temperature

○ s $\longrightarrow$ l and l $\longrightarrow$ g are endothermic, while the reverse reactions are exothermic

○ Phase change temperatures can be understood in the context of: $\Delta G = \Delta H - T \Delta S$

○ Sublimation and dry ice

8.2 Intermolecular Forces

○ Intermolecular forces determine the physical phase of a substance

○ London Dispersion Forces

· Figure 8.4

· Effect of molecular shape – More surface area gives stronger forces

○ Dipole-Dipole Forces

· Figure 8.3

· Effect of polarity – Stronger dipoles give stronger forces

○ Hydrogen Bonds

· H-bonds special cases which go beyond normal dipole-dipole bonds

· A H atom is very strongly attracted to another molecule (almost like the H is shared between molecules)

· Page 221 un-numbered figure

· Electron donor must be an electronegative atom (N, O, or F)

· Hydrogen must be bound to an electronegative atom (N, O, or F)

· Acetone is a counter-example. It has O and H, but not bound together

8.3   Gases and the Kinetic-Molecular Theory

○ Gas behavior can be mostly explained by assuming that the individual gas molecules follow these assumptions:

· A gas consists of many particles which move randomly and independently

· The individual gas particles exhibit no intermolecular forces

· The volume occupied by the gas particles is vanishingly small compared to the volume filled by the gas collectively (i.e. the particles are very small)

· The average kinetic energy of the gas particles is proportional to the temperature (in K)

· Gas particles collide elastically

○ Real gases deviate slightly from these assumptions, but we can ignore those deviations and treat them like an *ideal gas*

8.4   Pressure

○ Pressure is the force a gas applies over an area

○ Even ambient air is exerting a fairly high pressure – atmospheric pressure

○ We measure ambient pressure using a mercury barometer – Figure 8.9

○ Pressure can be measured in many different units: $1atm = 760mmHg = 760torr = 14.7psi = 101325Pa$

8.5   Boyle's Law

○ Early physicists experimented with the fundamental state variables of gases (P, V, and T) and noticed simple relationships called gas laws

○ Boyle showed the relationship between pressure and volume

○ Demo a small balloon expanding in a vacuum: $P \propto \dfrac{1}{V}$

○ The pressure increases with decreased volume, which is consistent with the kinetic molecular theory of gases because there are more molecular collisions per second in the smaller volume

○ Solve Boyle's law problems by relating the initial and final conditions: $P_1V_1 = P_2V_2$

8.6 Charles's Law

○ Charles's law relates the volume of a gas to temperature: $V \propto T$

○ Demo a small balloon expanding as it moves from ice to boiling water

○ This behavior is explained by the kinetic molecular theory of gases because the collisions are more forceful, pushing to greater volumes

○ Solve Charles's law problems with the following relation: $\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$

○ Practice Problems

    · 8.34 – Intermolecular forces

    · 8.42 – Converting gas pressures

    · 8.48 – Boyle's Law

    · 8.56 – Charles's Law

8.7 Gay-Lussac's Law

○ Gay-Lussac's law provides the last relation: $P \propto T$

○ This law is redundant, since the previous two laws could be combined to prove the Gay-Lussac law

8.8 The Combined Gas Law

○ Boyle's law, Charles's law, and the Gay-Lussac law all focus on the relationship between two state variables while keeping the third constant

○ These laws can be combined to give a law which allows all three variables to change
$\dfrac{PV}{T} = k$ where k is a constant value

○ Solve these problems with the following relation: $\dfrac{P_1 V_1}{T_1} = \dfrac{P_2 V_2}{T_2}$

○ This law includes all the others, so it is really the only one that you need to know, except for the even better law we will teach you next class

8.9 Avogadro's Law

○ All of the above laws are for a constant (and unknown) number of moles of gas

○ More moles of gas will give greater volume : $\dfrac{V_1}{n_1} = \dfrac{V_2}{n_2}$

○ It is convenient to define a standard temperature ($0°C$) and pressure ($1atm$) or comparing different gas systems

○ At STP, one mole of an ideal gas occupies a standard molar volume of $22.4L$

8.10 The Ideal Gas Law

○ The ideal gas law combines everything we have covered so far: $PV = nRT$

○ R is the gas constant, which can take on different values depending on the units:
$$R = 0.08206\dfrac{L\ atm}{mol\ K} = 62.4\dfrac{L\ Torr}{mol\ K}$$

○ Find the pressure when $0.250mol$ of gas are placed in a $5.20L$ container at $15.0°C$ ($1.14atm$)

○ How many moles are in a $3.15L$ helium balloon at $20.0°C$ and $0.820atm$? ($0.107mol$)

8.11 Partial Pressures and Dalton's Law

○ Since gas particles don't interact at all, a mixture of gases doesn't introduce any new complexity

○ Each gas exerts a partial pressure as if the other gases weren't there, and the total pressure is the sum of the partial pressures

○ $P_{total} = \sum P_{i,partial}$

○ The air is mostly made of O_2 and N_2. In Cedar City, $P_{O_2} = 0.17atm$ and $P_{N_2} = 0.64atm$. What is the total barometric pressure? ($0.81atm$)

○ Practice Problems

· 8.62 – Ideal gas law

· 8.82 – Ideal gas law

· 8.88 – Dalton's law of partial pressures

8.12   Liquids

○ A liquid in a closed container will reach equilibrium with the vapor phase as vaporization and condensation rates come into balance (Figure 8.19)

○ The equilibrium pressure above a liquid is the vapor pressure

○ Different liquids have different vapor pressures, and vapor pressure always increases with temperature (Figure 8.20)

○ When the vapor pressure equals the external pressure ($\approx 1 atm$), boiling will begin

○ The lower external pressure at high altitudes is why water will boil at less than $100°C$ at those altitudes

○ Viscosity is a measure of how freely a liquid flows. Some liquids, like honey, are very viscous

○ Surface Tension is the resistance of a liquid to spreading out and increasing its surface area. This leads to curved droplets of liquids, since this shape minimizes surface area (Figure 8.21)

8.13   Water: A Unique Liquid

○ Water has many unique properties compared to other liquids

○ Water has an unusually high heat capacity and enthalpy of vaporization due to its strong hydrogen-bonding intermolecular forces

○ Water is nearly unique in having a solid phase which is less dense than the liquid phase. This is because ice has an unusually open crystal structure (Figure 8.23)

8.14   Solids

○ Many solids exhibit an ordered structure over long ranges. These are called crystalline solids

○ Crystalline solids can be one of four types:

  · Ionic solids, like NaCl, are made of a lattice of cations and anions
  · Molecular solids, like sugar, are made of individual molecules arranged in a repeating lattice
  · Covalent network solids, like diamond, are extended networks of covalent bonds. A diamond can *kind of* be described as one giant molecule!
  · Metallic solids are made of a lattice of metal atoms which share a sea of electrons across the whole solid. This is why metals conduct electricity

○ Amorphous solids, like glass, show no long-range structure. They are formed when liquids cool too quickly for a proper lattice to form

○ Amorphous solids slowly grow softer over a wide temperature range, rather than melting sharply at a single melting temperature

8.15 Changes of State

○ When a substance undergoes a phase change, it will either give off or take in heat without changing the temperature

○ The amount of heat is called the heat of fusion or heat of vaporization

○ Figure 8.24 shows a heating curve across two phase changes

○ Heat for changing temperature: $q = mC_p\Delta T$

○ Heat for a phase change: $q = m\Delta H_{forv}$

○ $7.5g$ of ice begin at $-6.00°C$. Find how many J are required to convert this ice into steam at $125°C$ $(23,000J)$

For water: $C_s = 2.11\dfrac{J}{g\,°C}$ $\quad C_l = 4.18\dfrac{J}{g\,°C}$ $\quad C_g = 2.00\dfrac{J}{g\,°C}$

$\Delta H_{fus} = 334\dfrac{J}{g}$ $\quad \Delta H_{vap} = 2260\dfrac{J}{g}$

○ Practice Problems

· 8.92 – Pressure and boiling points

· 8.96 – Amorphous and crystalline solids

· 8.98 – Enthalpy of fusion problem

SOLUTIONS

9.1 Mixtures and Solutions

○ Recall the taxonomy of mixtures from chapter 1: Heterogeneous mixtures vs homogeneous mixtures, and solutions vs colloids

○ The primary difference is particle size: Solutions have particles $< 2nm$, colloids have particles $2 - 500nm$, and heterogeneous mixtures have particles $> 500nm$

○ In a solution, the solvent is the dominant component while any minority components are solutes

○ Solutions can be more than just solid-in-liquid – Table 9.2

9.2 The Solvation Process

○ In solvation, solute-solute interactions and solvent-solvent interactions are replaced by solute-solvent interactions

○ To determine solubility, consider the various interactions

· Stronger solute-solute interactions will make it hard to break apart the solute

· Stronger solvent-solvent interactions will make it hard to disrupt solvent to make room for solute

· Similar solute-solute, solvent-solvent, and solute-solvent interactions lead to high solubility

· This leads to the rule of thumb that "like dissolves like"

○ The solvent will form a solvation sphere around the solute

○ The net result can be exothermic or endothermic

9.3 Solid Hydrates

○ Some ionic compounds will naturally incorporate water molecules in their ionic lattice

○ These compounds are called "hygroscopic," and are written with a dot like this: $CaSO_4 \cdot \frac{1}{2} H_2O$

9.4   Solubility

○ Solubility is the amount of one substance that you can dissolve in another substance

○ If two liquids can be mixed in any ratio, then they are said to be miscible

○ If a solution has dissolved as much of a solid as it can, then it is called "saturated"

○ A saturated solution can be identified if excess solute remains undissolved

9.5   Temperature Dependence of Solubility

○ Increasing the temperature will increase the solubility of most, but not all solids

○ If solvation is endothermic, then solubility increases with increasing temperature. If solvation is exothermic, then solubility will decrease with increasing temperature

○ A solution can be heated to dissolve more solute, then cooled to form a supersaturated solution

○ Solute can be spontaneously precipitated from a supersaturated solution – Demo with sodium acetate

9.6   Pressure Dependence of Solubility

○ Pressure will affect the solubility of gases in liquids

○ The solubility is proportional to the gas pressure, according to Henry's Law

9.7   Units of Concentration

○ There are several different units used by chemists to describe concentration

○ Percent concentrations:

· mass/mass percent – mass of solute / mass of solution

· volume/volume percent – volume of solute

· mass/volume percent – mass of solute volume of solution

○ Practice Problems

· 9.38 – Water as a solvent

· 9.42 – Henry's law

· 9.44 – Saturation and solubility

· 9.50 – Preparing a solution of NaCl

9.8   Dilution

○ Dilution is when additional solvent is added to reduce the concentration of all solutes

○ The number of moles of solvent doesn't change, so we can set the number of moles after dilution equal to the number of moles before the dilution

○ $C_1V_1 = C_2V_2$ in general, or $M_1V_1 = M_2V_2$ for molar concentrations

○ $\dfrac{V_1}{V_2}$ is the dilution factor

9.9   Ions in Solution: Electrolytes

○ Substances which produce ions when dissolved in solutions are called electrolytes

○ Electrolytes conduct electricity

○ Substances fall into three categories:

· Strong electrolytes completely ionize (produce one or more ions for each formula unit)
· Weak electrolytes only partially ionize (often only a fraction of a percent)
· Non-electrolytes produce no ions

9.10   Electrolytes in Biology

○ In our biochemistry, there are often many different ions in solution

○ To talk about the concentration of complex solutions, we use two new units:

· Equivalents (Eq): The number of ions that carry 1 mole of charge
· Gram Equivalents (g-Eq): $\dfrac{\mathrm{MolarMassofIon(g)}}{\mathrm{ChargeonIon}}$

○ More realistic concentrations are milliequivalents

9.11   Properties of Solutions

○ Some properties of solutions depend on the concentration but not the identity of the solute

○ These properties are called colligative properties

○ Vapor-Pressure Lowering – Adding a solute will lower the solvent vapor pressure

○ Boiling Point Elevation – $\Delta T_b = \kappa_b C_{molal}$

○ Freezing Point Depression – $\Delta T_f = \kappa_f C_{molal}$

9.12   Osmosis and Osmotic Pressure

○ Osmotic pressure is another important colligative property

○ Osmosis – Pressure exerted across a semipermeable membrane $\pi = \dfrac{nRT}{V}$

○ Isotonic solutions match the salinity of blood cells

○ hypertonic solutions are more salty and will shrivel blood cells

○ hypotonic solutions are less salty and will swell (or lyse) blood cells

○ Practice Problems

- 9.68 – Dilution
- 9.74 – milliequivalents
- 9.82 – Freezing point depression
- 9.86 – Osmolarity

ACIDS AND BASES

10.1   Acids and Bases in Aqueous Solution

○ Acid and base activity is mostly defined in the context of an aqueous solution

○ An acid will react with water to produce hydronium ions

○ A base will react with water to produce hydroxide ions

10.2   Some Common Acids and Bases

○ Acids and bases are common in products around us. Page 292 lists a few

10.3   The Brønsted-Lowry Definition of Acids and Bases

○ There are several different definitions used for acids and bases, we will be using the Brønsted-Lowry definition:

· Remember that a H^+ ion is simply a proton

· An acid is a proton donor – HCl reaction with water

· A base is a proton acceptor – NH_3 reaction with water

○ Under this definition, water acts as a base when reacting with an acid, and acts as an acid when reacting with a base – More on this later

○ Multiprotic acids can donate more than one hydrogen

○ Every acid has a conjugate base, and every base has a conjugate acid – They make up a conjugate pair

10.4   Acid and Base Strength

○ Strong acids and bases will dissociate into H^+ or OH^- completely (and are strong electrolytes)

○ Weak acids and bases dissociate to different extents, less than 100%

○ The relative strengths of a conjugate pair are inverse to each other (i.e. The stronger and acid, the weaker will be its conjugate base, and vice-versa)

○ Table 10.1 shows conjugate pairs over a range of acid and base strengths

○ Acid-base reactions are equilibrium reactions, and the relative strengths determine whether the reaction is reactant or product favored

10.5   Acid Dissociation Constants

○ The strength of an acid is quantified by the acid dissociation constant K_a

○ This is really just th equilibrium constant for the reaction of acid with water

○ $K_a = \dfrac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$

○ Multiprotic acids will have a seperate constant for each stage of dissociation

○ Table 10.2 Gives the K_a values for several common acids

10.6   Water as Both an Acid and a Base

○ Water acts as a base when reacting with an acid, and acts as an acid when reacting with a base

○ Substances like this are called *amphoteric*

○ Water can also react with itself, acting as both an acid and a base to produce both hydronium and hydroxide

○ The water dissociation constant K_w is very important to acid-base chemistry and has the value $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

○ In a neutral solution, $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}$

○ As $[\text{H}_3\text{O}^+]$ goes up, $[\text{OH}^-]$ goes down and vice-versa to obey the K_w expression

○ Practice Problems

    · 10.38 – Complete dissociation of a strong acid

      change me for a better problem

    · 10.46 – Identify acids and bases

    · 10.54 – Equilibrium expression K_a

    · 10.56 – K_w

10.7 Measuring Acidity: pH

○ pH is a way to quantify the acidity of a solution

○ $pH = -\log_{10}[H_3O^+]$

○ Taking the log squeezes the very wide range of possible hydronium concentrations into a reasonably small range.

○ $pH = 7$ is neutral, $pH < 7$ is acidic, and $pH > 7$ is basic

○ We can also define a pOH: $pOH = -\log_{10}[OH^-]$

○ With these definitions we can convert $K_w = [H_3O^+][OH^-]$ into $14 = pH + pOH$

10.8 Working with pH

○ We find pH by taking the negative log of the hydronium concentration

○ We can find the hydronium concentration by taking ten to the power of -pH

○ Draw my box of conversions

10.9 Laboratory Determination of Acidity

○ One simple way to measure acidity is with a color indicator (Figure 10.3)

○ This can take the form of a drop to add to a solution, or a piece of litmus paper to drop the solution onto

○ Indicators are chemicals that change color based on the acidity of their environment

○ More accurate measurements can be made electronically using a pH meter

10.10 Buffer Solutions

○ Pure water will rapidly change pH if any acid or base is added

○ Buffer solutions can withstand changes in pH, even if strong acids or bases are added

○ Figure 10.5 compares pH changes in pure water vs a buffer solution

○ A buffer solution contains a conjugate acid/base pair, which reacts in place of the water to resist pH changes

○ To make an ideal buffer, pick an acid/base pair whose pK_a equals the desired pH you want to buffer at

○ The Henderson-Hasselbalch equation helps to solve for the pH of a buffer solution that doesn't have exactly equal amounts of the conjugate acid and base

○ $pH = pK_a + \log_{10}\left(\dfrac{[\text{A}^-]}{[\text{HA}]}\right)$

○ Our own bodies are regulated by buffer systems (carbonic acid and dihydrogen phosphate, among others)

10.11   Acid and Base Equivalents

○ Just like with ion equivalents, it is often convenient to talk about acids and bases in terms of how many moles of H_3O^+ of OH^- they will produce

○ $1eq = \dfrac{1mol}{\#\ of\ \text{H}_3\text{O}^+\ or\ \text{OH}^-\ ions\ produced}$

○ $1g - eq = \dfrac{MolarMass}{\#\ of\ \text{H}_3\text{O}^+\ or\ \text{OH}^-\ ions\ produced}$

○ Since acids and bases will partition between their conjugates, we can't properly talk about Molarity without doing further calculations

○ What we really want is the number of equivalents in a solution. This is called the *Normal Concentration* (N)

○ $N = \dfrac{eq}{L}$

10.12   Some Common Acid-Base Reactions

10.13   Titration

○ Titration is a technique used to determine the concentration of acids and bases

○ To titrate an acid, a base of known concentration is added dropwise until the the acid is neutralized

○ To titrate a base, an acid of known concentration is added dropwise until the the base is neutralized

○ At the equivalence point, the number of moles of acid and base should be equal

○ The equivalence point can be identified by a color indicator, since the pH swings rapidly once the neutralization has occurred

○ $N_1V_1 = N_2V_2$ – This looks just like the dilution equation because it is based on the same assumption: That the number of moles is the same

10.14   Acidity and Basicity of Salt Solutions

- ○ When a salt dissolves, the cation and anion dissociate and are free to react with water

- ○ The cation may be an acid, and the anion may be a base or an acid, so we need to consider them both

- ○ Remember that a salt is the product of an acid-base reaction, so we can classify salts by the type of reactions which will produce them:

    - · Strong acid + weak base = acidic solution
    - · Weak acid + strong base = basic solution
    - · Strong acid + strong base = neutral solution
    - · Weak acid + weak base = ? (it depends on the specific acid/base strengths)

- ○ Practice Problems

    - · 10.68 – Rough calculations of pH
    - · 10.70 – pH to $\left[H_3O^+\right]$ and $\left[OH^-\right]$
    - · 10.78 – Find pH of a buffer
    - · 10.90 – Titration

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CHAPTER 11

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NUCLEAR CHEMISTRY

</div>

11.1   Nuclear Reactions

○ Atomic symbols keep track of all subatomic particles: $^{12}_{6}C$, $^{13}_{6}C$, $^{12}_{6}C^{+}$, etc.

○ We can describe changes in nuclear reactions using these symbols:

$$^{14}_{6}C \longrightarrow \ ^{14}_{7}N + \ ^{0}_{-1}e^{-}$$
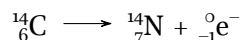
○ In the reaction above, $^{0}_{-1}e^{-}$ (simply an electron) does not really have an atomic number, but the $_{-1}$ is useful for balancing reactions

11.2   The Discovery and Nature of Radioactivity

○ Radioactivity was first observed in 1896 when photosensitive paper stored under a Uranium sample was exposed despite being kept in the dark

○ Radiation is categorized as α particles, β particles, or γ radiation

○ Figure 11.1 shows how different types of radiation respond to an electric field

○ Table 11.1 shows the physical properties of different types of radiation
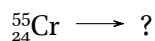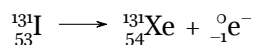
11.3   Stable and Unstable Isotopes

○ Radioactivity comes from unstable nuclei

○ Some unstable nuclei are naturally occurring (U), while others (Pu) are generated artificially

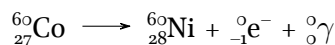○ Figure 11.2 shows the "belt of stability"
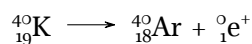
11.4   Nuclear Decay

○ Alpha Emission – Figure 11.3

$$^{238}_{92}U \longrightarrow \ ^{4}_{2}He + \ ^{234}_{90}Th$$

$$^{208}_{84}Po \longrightarrow \ ?$$

- Beta Emission – $^1_0n \longrightarrow \, ^1_1p^+ + \, ^0_{-1}e^-$

  $^{131}_{53}I \longrightarrow \, ^{131}_{54}Xe + \, ^0_{-1}e^-$

  $^{55}_{24}Cr \longrightarrow \, ?$
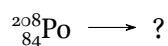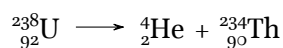
- Gamma Emission – Often present with other decay (Exception is Hafnium battery)

  $^{60}_{27}Co \longrightarrow \, ^{60}_{28}Ni + \, ^0_{-1}e^- + \, ^0_0\gamma$

- Positron Emission – $^1_1p^+ \longrightarrow \, ^1_0n + \, ^0_1e^+$

  $^{40}_{19}K \longrightarrow \, ^{40}_{18}Ar + \, ^0_1e^+$

- Electron Capture – $^1_1p^+ + \, ^0_{-1}e^- \longrightarrow \, ^1_0n$

  $^{197}_{80}Hg + \, ^0_{-1}e^- \longrightarrow \, ^{197}_{79}Au$
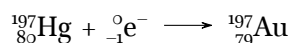
11.5 Radioactive Half-Life

- The half-life is the time required for half of a radioactive sample to decay

- Figure 11.4 shows how the absolute decay rate decreases, but the half-life is constant

- $Fraction\ Remaining = (0.5)^n$

- Table 11.3 shows the half-lives and uses for common radioisotopes

11.6 Radioactive Decay Series

- Radioisotopes often decay into other radioisotopes, creating a cascade of different reactions

- Figure 11.5 shows the Uranium radioactive decay series

11.7 Ionizing Radiation

- Radioactive products are most damaging when they eject electrons off of other materials

- Table 11.4 shows the penetrating distance of different types of ionizing radiation
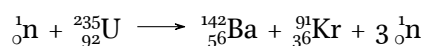
11.8   Detecting Radiation

11.9   Measuring Radiation

11.10   Artificial Transmutation

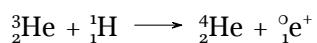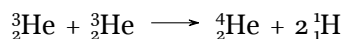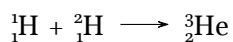## 11.11   Nuclear Fission and Nuclear Fusion

- ○ Fission and fusion both release tremendous amounts of energy, which manifest as lost mass according to $E = mc^2$

- ○ Fission – Figure 11.7

  $${}^{1}_{0}n + {}^{235}_{92}U \longrightarrow {}^{142}_{56}Ba + {}^{91}_{36}Kr + 3\,{}^{1}_{0}n$$

- ○ Note that the production of 3 neutrons can lead to a chain reaction

- ○ Fusion – Combining smaller nuclei to form larger ones

  $${}^{1}_{1}H + {}^{2}_{1}H \longrightarrow {}^{3}_{2}He$$

  $${}^{3}_{2}He + {}^{3}_{2}He \longrightarrow {}^{4}_{2}He + 2\,{}^{1}_{1}H$$

  $${}^{3}_{2}He + {}^{1}_{1}H \longrightarrow {}^{4}_{2}He + {}^{0}_{1}e^{+}$$

- ○ Practice Problems

  - · 11.40 – α and β decay
  - · 11.44 – Predicting products of β decay
  - · 11.48 – Balancing fission reactions
  - · 11.56 – Half-life calculations