Honors Chemistry Notes

anastasia

Fall 2022 & Spring 2023

Contents

1 Nature of Science

Chemistry Theme Songs? Sodium-Sodium-Sodium-Sodium Sodium-Sodium-Sodium-Sodium Batman

1.1 Lab Safety & Equipment

- Goggles must be worn over your eyes at all times! Wearing safety goggles correctly is required to protect your eyes during laboratory investigations.
- Your clothing should cover your legs - shorts are not appropriate for the laboratory. Lab aprons can be used to protect good clothing. No loose clothing! It can dip into chemicals or fall into a flame and catch fire.
- Sandals and open-toed shoes do not protect your feet from broken glass that is frequently found in the lab. Also, leather shoes protect your feet from chemical spills, shoes do not.
- Dangling hair can fall into the Bunsen burner and catch fire or can fall into a chemical solution.
- Do not apply cosmetics, eat, or drink in the lab - these activities are ways by which you can accidentally ingest harmful chemicals.
- Do not taste chemicals. Do not small any chemicals directly. If you smell chemicals, use your hand to waft vapors to your nose.
- Heat test tubes at an angle and away from you and others.
- Handle hot glassware with the appropriate tongs.
- Never work along in the lab - in the case of a problem, you may need another person to prevent injury or even save your life!
- Don't assume you dispose waste down the sink. Dispose of all waste materials according your instructional procedure.
- Never remove chemicals from the laboratory.
- Wash your hands with soap and water before leaving - this rule applies even if you have been wearing gloves!
- Report any accidents or unsafe conditions immediately!
- Remember that the lab is a place for serious work! Careless behavior may endanger yourself and others and will not be tolerated!

Know the safety equipment and how to use the following safety equipment.

- Eye wash fountain
- Safety shower
- Fire extinguisher
- Emergency exits

NFPA Chemical Hazard Label

- Blue - Health
- Red - Flammability
- Yellow - Reactivity (Stability)
- White - Special

Hazard Ratings

- 4 - Severe
- 3 - Serious
- 2 - Dangerous
- 1 - Minor
- 0 - Slight

MSDS

- Material Safety Data Sheet (now often just called Safety Data Sheet, SDS)
- On file for all purchased chemicals.
- Includes all information shown on a chemical label and more.

Lab Equipment

- Beakers hold liquids. They don't precisely measure.
- Test tubes hold small amounts of liquid.
- Erlenmeyer flasks are used to hold liquids and swirl mixtures.
- Test tube racks hold test tubes.
- Bunsen burners heat with intensity.
- Hot plates heat at a wide variety of temperatures, from low to high.
- Plastic pipettes transfer small, approximate amounts. Not for measuring.
- Volumetric flasks are used for making solutions of a specific volume. They only have one line for measuring.
- Beaker tongs are used to pick up a hot beaker.
- Test tube tongs are used to hold one test tube.
- Crucible tongs are used to pick up a crucible or hold something in flame.
- Ring stand & rings are used for holding items over flame for a long period of time or filtering.
- Wire gauzes are used to put hot beakers on to prevent shattering.
- Balances are used to measure the mass of an object.
- Glass pipettes measure small amounts of liquid by suction.
- Graduated cylinders measure the volume of a liquid.

Exercise List the Laboratory Safety DOs and DON'Ts

Exercise What are four pieces of safety equipment and how do you use them?

Exercise What does a Safety Data Sheet (SDS) tell you?

Exercise Which lab equipment measures chemicals?

1.2 Matter, Energy, & Change

Chemistry is the science that investigates structures and properties of matter.

- Matter - anything composed of atoms
- Mass - a measure of how much matter is in an object
- Weight - measure of gravity's pull on matter
- Volume - measure of how much space is taken up

Exercise What is the difference between mass and weight, and what instruments are used to measure mass and weight?

There are two types of data

- Qualitative (qualities)
- Quantitative (quantities)

Graphs

- Independent Variable - the one that is controlled or consistent; found on the x-axis
- Dependent Variable - the result; found on the y-axis

Measurable Properties

- Extensive - property that depends on HOW MUCH matter you have
- Intensive - property that is INDEPENDENT of the amount of matter

Physical and Chemical Properties

- A physical property can be observed without a chemical change occurring.
- A chemical property can be observed only when a chemical change occurs. In physical changes:
  - atoms are not rearranged into new substances
  - include all changes of state
  - changes in size, shape, or dissolving

  In chemical changes:
  - bonds are broken between atoms and new bonds are formed to make new substances.
  - Chemical changes are usually more interesting than physical changes

Four Indicators of a Chemical Change

1. Energy change - heat or light is produced, or a decrease in temperature occurs
   - Exothermic - gives off heat, feels hot
   - Endothermic - absorbs heat, feels cool
2. Production or evolution of a gas
3. Precipitate - a solid is formed when two liquids are mixed together
   - The clue that a precipitate has formed is that the liquid turns cloudy, it could be any color.
4. Color Change

Exercise Is tarnishing a physical of chemical change/property?

Exercise Is breaking a physical of chemical change/property?

Exercise Is combustion a physical of chemical change/property?

Classification of Matter

- Mixture: two or more pure substances that can be separated by physical changes.
- Homogeneous Mixture: two or more pure substances mixed evenly. When you look at it, you can't see separate parts.
- Heterogeneous Mixture: two or more pure substances mixed unevenly.
- Element: one of the 118 pure substances that cannot be separated by chemical change or physical change. Represented by a symbol on the periodic table.
- Allotrope: same element with different bonding of atoms (different properties)

- Compound: made from atoms that are chemically bonded together. Can be separated by chemical change, but not physical change. Represented by a formula.

Exercise How is bonding different than mixing?

The Law of Definite Proportions (sometimes called Law of Constant Composition) states that all samples of a compound contain the same elements in the same proportion.

The Law of Multiple Proportions states that if elements combine to make more than one compound, the masses will be small, whole number ratios.

The Law of Conservation of Mass states that matter cannot be created or destroyed in any type of change. What you start with is what you end up with, just in a different form.

The Law of Conservation of Energy states that energy cannot be created or destroyed (but it can change forms).

The Periodic Table

- Find the zig-zag line.
- Metals are to the left of the zig-zag line (except for H)
- Non-metals are to the right of the zig-zag line
- Elements touching the line are called metalloids
- The vertical columns are called groups (or families)
- The horizontal rows are called periods.

1.3 Measurement

Accuracy vs. Precision

- Accuracy - how close a measurement is to the accepted value
- Precision - how close a series of measurements are to each other

Accuracy is correctness, precise is consistency.

Percent error indicates the accuracy of a measurement

$$\%error = \left| \frac{accepted - experimental}{accepted} \right| \times 100$$

Example

Juan calculated the density of aluminum three times: 2.75 g/cm^3, 2.68 g/cm^3, and 2.84 g/cm^3.

Aluminum has a density of 2.70 g/cm^3. Calculate the average percent error for the three trials.

The percent error for the trials are 1.85%, 0.74% and 5.19% respectively. The average of the three is 2.59%.

Exercise Suppose you calculate your semester grade in chemistry as 90.1, but you receive a grade of 89.4 on your report card. What is your percent error?

Exercise On a bathroom scale, a person always weighs 2.5 lbs less than on the scale at the doctor's office. What is the percent error of the bathroom scale if the person's actual weight is 125 pounds?

Significant Figures

- Indicate accuracy of a measurement.
- Sig figs in a measurement include the known digits plus a final estimated digit.

- It is important to be honest when reporting a measurement so that is does not appear to be more accurate than the equipment used to make the measurement.

Counting Sig Figs

Count all numbers except

- Leading zeroes
- Trailing zeros without a decimal point

Rules for Counting Sig Figs

- All nonzero digits are significant
- Sandwiched zeroes are significant
- Zeroes at the beginning are never significant
- Zeroes at the end are significant only if you can see the decimal point

Note:

- Non significant does mean unaccounted for
- Sig Figs keep track of the accuracy of our measurements

Exercise Count the number of sig figs in each number

1. 98
2. 0.0098000
3. 980.0

Scientific Notation Converting into scientific notation:

- Move decimal until there's 1 digit to its left. The places moved is the exponent.
- A number greater than 1 gets a positive exponent and a number less than 1 gets a negative exponent.

Exercise Write in scientific notation and keep the same number of significant figures:

1. 0.00007
2. 422000.

Exercise Write in standard notation and keep the same number of significant figures:

1. 3.1 $\times 10^4$
2. 1.00 $\times 10^2$

Mathematical Operations with Sig Figs

- When combining measurements with differing degrees of accuracy and precision, the accuracy of the final answer can be no greater than the least accurate measurement.
- This principle can be translated into simple rules for mathematical operations.
- Remember the order of operations and always include units in your answer if units are given in the problem.

When adding or subtracting, the answer cannot go beyond the last significant place of the least precise measurement.

When multiplying or dividing, the # with the fewest sig figs determines the # of sig figs in the answer.

Exact numbers do not limit the number of significant figures.

Exercise

1. 150.0 grams + 0.507 grams
2. 98.0 grams ÷ 2.33 liters

Tips:

- Determine which rule you are dealing with first! Add/Sub = least decimal places. Mult/Div = least number of sig figs.

Density

- Density is the measure of how much mass is contained in a given unit of volume.
- It depends on what the matter is, not how much you have.
- Density is an intensive property.

Density depends on two things:

1. How tightly packed the atoms are
2. What kind of atoms they are

Density is calculated with the formula

$$D = \frac{m}{V}.$$

This can be arranged to solve for mass or volume.

Exercise Use algebra to rearrange the density formula to solve for volume.

When working density problems, use the following steps:

1. Write the correct formula you'll be using
2. Substitute in the correct values with units
3. Work the problem with your calculator and give the answer with the correct number of sig figs and correct units

Exercise A metal cylinder is placed into a graduated cylinder with a 24.0 mL of water. After the cylinder is added, the volume of water rised to 30.4 mL. The density of the cylinder is known to be 8.9 g/mL. What is the mass of the cylinder?

Exercise A metal cylinder has a diameter of 4.4 cm and a height of 10.5 cm. If the cylinder is silver, which has a density of 10.5 g/cm^3, what is its mass? The volume of a cylinder is $\pi r^2 h$. (Use 3.14 for pi.)

Proportions In a direct proportion, the relationship should be linear.

In an inverse proportion, the relationship will be non-linear and decreasing.

1.4 Dimensional Analysis

First, off the metric system!

S.I. or metric units are: Mass in grams (g), Length in meters (m), Volume in liters (L)

Prefixes to know: kilo = 1000, centi = 1/100, milli = 1/1000

Memorize these conversions!

- 1 kg = 1000 g
- 1 g = 100 cg
- 1 g = 1000 mg
- 1 km = 1000 m
- 1 m = 100 cm
- 1 m = 1000 mm
- 1 cm = 10 mm
- 1 L = 1000 mL

Dimensional analysis is the method that chemists (and other scientists) use to solve conversion problems.

Exercise

1. Convert 23.9 km to m
2. If 1 inch = 2.54 cm, convert 3.00 cm to inches
3. Convert 25 inches to cm
4. If 1 gallon = 4.1 L, convert 2.5 gal to L

Exercise

1. What is the length of a football field in cm if there are 2.54 cm in an inch and 36 inches in a yard?
2. Diamonds are measured in units called a carat. One carat equals 200 mg. If a diamond is 0.600 carat, what is the mass of the diamond in ounces?

Chapter Problems

1. How many significant figures are in 1.003?
2. Write 0.00007 in scientific notation, keeping the same number of significant figures.
3. Write 1.00×10^2 in floating decimal notation (standard notation), keeping the same number of significant figures.
4. Round 0.003008 to 3 significant figures.
5. Calculate 45.0 cm - 9.2 cm and round to the correct number of significant figures and include the correct units.
6. Round 400×600 to the correct number of significant figures.
7. If candy bars are 3 for one dollar, how much money will you need to buy 46 candy bars?
8. What volume will be occupied by 7.0 kg of helium if 4.003 g of helium occupies 22.4 L?
9. Twenty five paper clips are dropped into a graduated cylinder and the water level rises from 10.8 mL to 12.2 mL. If the density of the paper clips is 7.87 g/mL, what is the mass of the 25 paper clips? What is the mass of one paper clip?
10. A block of wood with a density of 0.548 g/cm^3 has a mass of 34.49 g. If two dimensions of the block are 2.5 cm and 7.8 cm, what is the 3rd dimension?

2 Atomic Structure and Energy of Electrons

You can't trust atoms. They make up everything.

2.1 Atomic Theory & Structure

Theories vs. Laws

- A theory is an explanation based on many observations.
- A law is a fact of nature that is observed so often it is accepted as truth.
- Theories EXPLAIN laws
- Both a scientific theory and scientific law are accepted to be true by the scientific community as a whole.
- A theory is like a car. Components of it can be changed or improved upon, without changing the overall truth of the theory as a whole.

What is atomic theory?

- The idea that matter is made up of atoms, the smallest pieces of matter.
- Over the years, atomic theory has evolved and changed to better explain scientific observations about atoms.
- Ancient Greeks believed all matter was made up of four basic elements: fire, earth, water and air.
- Democritus
  - Greek philosopher
  - Idea of 'democracy'
  - Idea of 'atomos'
    - Atomos = 'indivisible'
    - 'Atom' is derived
  - No experiments to support idea

Democritus's model of the atom consisted of a solid and indestructable atom with no protons, electrons, or neutrons.

- Lavoisier 18th century
- Proposed the law of conservation of mass/matter.
- Observed that the mass of the reactants equaled the mass of the products in a chemical reaction.
- Proust - Proposed the law of definite proportions for compounds.

Dalton's Atomic Theory

- All matter is made of tiny indivisible particles called atoms.
- Atoms of the same element are identical, those of different elements are different.
- Atoms of different elements combine in whole number ratios to form compounds.
- Chemical reactions involve the rearrangement of atoms. No new atoms are created or destroyed.

Thomson

- J.J. Thomson - English physicist. 1897
- Made a piece of equipment called a cathode ray tube. It is a vacuum tube - all the air has been pumped out.
- Thomson's Model - Plum Pudding Model (also called Chocolate Chip Cookie Model)
  - Atoms are composed of charged particles (subatomic particles).
  - The particles that were attracted to the positive plate were negative.
    - These were called "electrons"
    - Protons were discovered the same way.

Rutherford 1895

- Experiment: Gold Foil Experiment
- Most particles pass through, but some are bounced back towards the source.
- Model: Rutherford explained that atoms must be mostly empty space with a small, concentrated center of positive charge.

Chadwick

- Discovered the neutron.
  - Neutron is a subatomic particle roughly the size of a proton (large compared to electrons).

Bohr

- Model: proposed the "electron cloud" in which electrons orbit at a given distance from the nucleus.
- Small orbits = low energy
- Big orbits = high energy

Quantum Mechanical Model
Modern atomic theory describes the electronic structure of the atom as the probability of finding electrons within certain regions of space (orbitals).

Modern View

- The atom is mostly empty space
- Two regions
  - Nucleus
    - protons and neutrons
  - Electron cloud
    - region where you might find an electron

2.2 Structure of Atom & Isotopes

Major Parts of the Atom

- Nucleus: dense, central part of the atom
- Protons and neutrons are found in the nucleus
- Electron cloud: large area outside of the nucleus
- Electrons occupy the electron cloud

Protons are located in the nucleus with positive charge and have a large relative size.

Nuetrons are located in the nucleus with 0 charge and with a large relative size.

Electrons are located in the electron cloud, have a negative charge and have a tiny relative size.

Atoms and the Periodic Table

- Atomic Number - the whole number in an element's box on the periodic table.
  - Atomic # = # protons = # electrons
  - The atomic number determines an element's identity!

Exercise An atom has 24 protons. What element is it?

- Mass Number - the sum of the protons and neutrons
- This number isn't on the periodic table, because the number of neutrons can vary (these are called isotopes)
- Atomic Mass - the decimal number on the periodic table. The weighted average mass of all isotopes of that element.
- Isotopes - atoms of the same element that have different mass numbers.
- This means the number of protons is the same, and the number of neutrons is different.

Isotopes of Hydrogen

- Protium - 1 proton, 1 electron, mass number of 1
- Deuterium - 1 proton, 1 neutron, 1 electron, mass number of 2
- Tritium - 1 proton, 2 neutrons, 1 electron, mass number of 3

How to write isotopes

- Method 1: Subscript/Superscript Method
- The atomic # is your subscript (below) and the mass # is the superscript (above), both on the left side of the symbol
- Method 2: Hyphen-notation method
- This symbol is written, then hyphen, then mass #

Exercise Given ruthenium and the super/sub method of $^{101}_{4}4$Ru, write the atomic number, mass number, number of protons, neutrons, and electrons and the hyphen method for this element.

2.3 Average Atomic Mass

- Atoms can't be easily measured in grams because they are so small.
- Scientists devised "atomic mass units" - a carbon-12 isotope is 12.000000 amu's.

Average Atomic Mass

- A different kind of average - a "weighted" average.
- This means that we take into account the abundance of each isotope found in nature.

Formula to memorize:

$$[(\text{mass})(\text{abundance})+(\text{mass})(\text{abundance})+(\text{mass})(\text{abundance})]/100.000$$

- That's for 3 isotopes. Use the (mass)(abundance) for as many isotopes as there are.
- The 100 won't limit sig figs in your answer. Your answer is limited by whichever mass or abundance has the fewest sig figs.

Exercise Argon has three isotopes with the following percent abundances: Ar-36 with a mass of 35.968 amu and an abundance of 0.3337%. Ar-38 with a mass of 37.963 amu and an abundance of 0.063%. Ar-40 with a mass of 39.962 amu and an abundance of 99.600%. Calculate the average atomic mass.

Exercise The atomic weight of gallium is 69.72 amu. The masses of naturally occurring isotopes are 68.92 amu for Ga-69 and 70.92 amu for Ga-71. Calculate the percent abundance of each isotope.

2.4 Moles

- A mole is the amount of substance that contains the same number of atoms as 12 grams of Carbon-12.
- It is a counting unit just like a dozen.
- A mole is 6.02×10^{23} of something.
- 6.02×10^{23} is called "Avogadro's Number" because Amedeo Avogadro discovered it.
- 1 mole of any element has a mass (in grams) equal to its average atomic mass.

Molar Mass

- When we write out the average atomic mass in "grams" we call this the molar mass - it is literally the mass of one mole.

Conversions

1.0000 mole of any substance equals 6.02×10^{23} atoms of that element equals molar mass in grams of that element.

To do a molar conversion problem:

- Do dimensional analysis.
- Start with what you're given.
- Bring that unit down and over.
- Put the unit you want on top.
- Fill in the numbers.
  - Put a "1" in front of moles in a conversion.
  - Put "6.022×10^{23}" in front of atoms in a conversion.
  - Put the molar mass in front of grams in a conversion.

Exercise What is the mass in grams of 7.00 moles of iron?

Exercise What is the number of moles in 4.50×10^{-20} grams of oxygen?

Exercise How many atoms are in 1.00×10^{-10} grams of gold?

2.5 Electron Configuration

Energy Level

- The region surrounding the nucleus where an electron is likely to be found.
- Think of rungs on a ladder, fixed levels with space in between.
- Sublevel - smaller part of an energy level indicated by letters (1s, 2s, 4d, etc.)
- Orbital - smaller part of a sublevel, each orbital holds 2 electrons, moving in opposite direction... (4 possible shapes)
- "Electron configuration" describes the location of electrons in a given atom. This determines how an element behaves chemically, and thus is the core of chemistry.

We'll learn three ways to show electron configuration

- Orbital Notation
- Electron Configuration
- Lewis Dot Structures

Aufbau Principle - electrons enter orbitals of lowest energy first. Low energy orbitals are closer to the nucleus.

Pauli Exclusion Principle - no two electrons can be in the same orbital moving the same way. Each electron is unique.

Hund's Rule - when electrons are filling up orbitals of equal energy (say for instance 3 orbitals, which is 6 electrons), one electron enters each orbital until they're half-filled with electrons spinning in the same direction, then they fill with the opposite-spin electrons

Orbital Notation

- Numbers represent energy levels and letters represent sublevels
- Lines represent 1 orbital each (can also use boxes)

Electron configuration notation

- Write coefficient & letter for each energy sublevel.
- Superscript (number on top) shows # of electrons at that sublevel.
- This method simply takes less space.

Shorthand Notation

- If you had to show the electron configuration for bismuth, it would be long. There is a way to shorten what you have to write.
- Use the symbol for the noble gas before the element you are using and put it in brackets. That represents all the electrons up until that noble gas. Then continue with the rest of the electron configuration.

f-block issues

- Period 6
  - f-block includes elements La to Yb
  - d-block includes elements Lu to Hg
- Period 7
  - f-block includes Ac to No
  - d-block includes Lr to Uub

Lewis Dot Notation

- Lewis Dot diagrams show electrons available for bonding. These are the outermost electrons (valence electrons).
- Valence electrons are the total electrons in the last energy level (highest coefficient).
- Notice that electrons do not pair up until all four sides have one electron already.

Exercise Use Electron Configuration Notation to write the electron configuration for oxygen.

Exercise Use the Noble Gas Core Method to write the electron configuration for nickel.

Exercise Use Lewis Dot Notation for argon.

2.6 Ion Electron Configurations

- How do positive ions (cations) form? Atoms (typically metals) lose electrons.
- How to negative ions (anions) form? Atoms (typically nonmetals) gain electrons.

When representative elements (s and p block) become ions, they take on the electron configuration of the nearest noble gas. This gives them 8 valence electrons.

- Transition metals (Groups 3-12) often have variable charges
- Use these guidelines to help figure out their electron configuration
  - Transition elements usually lose their s and p electrons first.
  - Completely full, half-full, or empty sublevels are stable.
  - Electrons can move from s-sublevels to d-sublevels if it makes the atom more stable.

Memorizing Monatomic Ions

- Monatomic cations - attach "ion" to the element name
- Monatomic anions - change the element ending to "-ide"
- The systematic name just uses a Roman numeral to indicate the charge. Used for transition metals (variably charged)

Memorizing Polyatomic Ions

- Help with formulas
  - Does the polyatomic ion contain an element in "the elbow"?
  - If so, the ion "-ate" 3 oxygen atoms
  - If not, the ion "-ate" 4 oxygen atoms
  - "-ite" ions contain 1 less oxygen atom than "-ate"
  - "hypo-x-ite" ions have 1 less oxygen atom than "-ite"
  - "per-x-ate" ions have 1 more oxygen atom than "-ate"
- Help with charges
  - There are only two polyatomic cations; the rest are anions
  - If the polyatomic ion contains oxygen, look at what group the other element is in
    * If it's an even # group, the ion charge is even
    * If it's an odd # group, the ion charge is odd

2.7 EM Spectrum

The Wave-Particle Theory

- A theory that attempts to explain how electrons can behave in two different ways
  - as waves (like light)
  - as particles (like a ball)

First, we will look at wave behavior.

Light consists of electromagnetic waves that travel 3.00×10^8 m/s.

- That's 670,616,629 miles per hour!
- This is the "speed of light", also known as "c"

Electromagnetic Waves

- The electromagnetic (EM) spectrum is a series of waves that have different wavelengths.
- Visible light is small portion of the EM spectrum, with mid-energy.
- EM waves are also called radiation.

EM Wave Characteristics

- Amplitude - height from origin to crest
- Frequency - number of cycles that pass a given point in a given amount of time
  - Measured in Hertz (Hz)
  - 1 Hz $= 1$ wave passes per second
  - 1 Hz $= 1/$s $=$ s^{-1}
  - Symbol is nu, ν
- Wavelength - distance between crests of a wave
  - Symbol is lambda, λ

All EM Waves move at the speed of light

$$c = \lambda\nu$$

As wavelength increases, frequency decreases. They are inversely proportional.

Exercise What is the vavelength of a wave with a frequency of 7600 Hz?

Important conversions:

- 1 m $= 1 \times 10^9$ nm
- 1 MHz $= 1 \times 10^6$ Hz

Exercise - What is the frequency of a wave with a wavelength of 467 nm? (6.42×10^{14} Hz)

- The visible spectrum of ROYGBIV is continuous; there are no breaks and the colors blend together.
- White light is a combination of all colors of light. A prism breaks up white light into the separate colors so we can see them.
- Each color has a definite frequency and wavelength.
  - The speed these colors of light are traveling never changes; it's always the speed of light, c

Low energy colors have a long wavelength and low frequency, while high energy colors have a short wavelength and high frequency.

- Remember that electrons occupy energy levels.
- When electrons are in the lowest energy level, they are said to be in their ground state
- It is possible for electrons to jump from ground state to a higher energy level (called excited state) by absorbing energy.
- When electrons lose energy they will fall back down to their ground state and release energy, and some of it is released as waves we can see - LIGHT!
- With many electrons jumping to energy levels and falling back, many different shades of light are released and blended.
- We can use a prism to separate the light to see the individual shades.
- This is called an atomic emission spectrum.

Types of Spectra

- Continuous Spectrum - no breaks
- Atomic Emission Spectrum - a lot of black space, aka "bright line" spectrum
- Absorption Spectrum - small dark regions, aka "dark line" spectrum

Spectroscopy is the science of producing atomic spectra and studying them.

Particle Model

- The idea that light can act as a particle

- Particles of light are called photons, or quanta (plural for quantum)
- A quantum behaves like a particle, and can move other matter

The Photoelectric Effect

- The particle model was needed to explain why when you shine a high energy light on some metals, electrons are ejected (moved) from the metal
- Einstein proposed in 1905 that light can behave as both a wave and a particle.
- He defined a photon as a particle of electromagnetic radiation with no mass that carries a quantum of energy.
- For this, he won the Nobel Prize.
- The energy contained in a photon (a quantum) depends on its frequency

$$E_{\text{photon}} = h\nu$$

  $E =$ energy in joules [J]
  $h =$ Planck's constnat $= 6.626 \times 10^{-34}$ J·s $\nu =$ frequency (nu), [Hz]

- According to Planck, matter can emit or absorb energy only in whole quanta ($1h\nu$, $2h\nu$, etc.)

Exercise Calculate the frequency of a photon with 7.2×10^{-34} J of energy.

Exercise Calculate the wavelength of a photon with 5.32×10^{-33} J of energy.

Chapter Problems

1. Give the atomic number, mass number, number of protons, neutrons, electrons, and super/sub method for arsenic.
2. How many moles are equal to 5.098×10^{23} atoms of hydrogen?
3. How many moles of nickel are equal to 2.04×10^{-4} grams?
4. What is the wavelength of radiation that has a frequency of 5.00×10^{12} Hz?
5. In order to eject one molre of photons from gaseous cesium atoms 382 kJ of energy is absorbed. Calculate the wavelength associated with this energy.
6. How many Joules of energy are there in one photon of yellow light whose wavelength is 630 nm?
7. Which orbital has the lowest energy?
8. The element magnesium consists of three naturally occurring isotopes with masses 23.98504, 24,98584, and 25.98259 amu. The relative abundances of these three isotopes are 78.70, 10.13, and 11.17 percent, respectively. From these data calculate the average atomic mass of magnesium.

3 Periodicity

How often do I like jokes about chemistry? Periodically.

3.1 Introduction to Periodic Table & Activity

Dmitri Mendeleev

- Many people had arranged the known elements of their day, and Dmitri Mendeleev arranged them by increasing atomic mass.

In 1869 when created, he left gaps and predicted some elements that had yet to be discovered. Later, they were and they fit into his table perfectly.

Henry Moseley

- Moseley's periodic table was similar, but he arranged them in order of increasing atomic number, not mass
- Remember that atomic number is the same as number of protons.
- This is the periodic table we use today.

Mass vs. Number

- Increasing atomic mass and atomic number are not exactly the same
- On our modern periodic table (Moseley's) there are a few atomic masses out of order. That's okay because we organize it by atomic number.

Modern Periodic Law

- States that the physical and chemical properties of elements repeat when they are arranged by increasing atomic number.

Classification of Elements

- The zig-zag line divides periodic table into two parts.
- Left of zig-zag line are metals.
- Right of zig-zag line are nonmetals.
- The elements touching the line are metalloids.

Properties of Metals

- Usually silver-gray in color, except gold & copper
- Solid at room temperature, except mercury
- Lustrous or shiny appearance
- Malleable
- Ductile
- Good conductors
- Usually react with acids
- High melting points

Properties of Nonmetals

- Dull

- Brittle (nonmalleable)
- Poor conductors of heat and electricity
- Usually no reaction with acids
- Gases, liquids, or low-melting-point solids

Properties of Metalloids

- All elements touching zig-zag line, except aluminum which is a metal
- Exhibit properties of both metals and nonmetals
- Not good conductors alone

The metals can be divided up into smaller groups.

Alkali Metals

- Group 1 of the periodic table
- Have one valence electron
- Very reactive
- Form $+1$ ions
- The exception in group 1 is hydrogen, which is not an alkali metal

Alkaline Earth Metals

- Group 2
- Two valence electrons
- Form $+2$ ions
- Less reactive than group 1

Blocks:

- The periodic table is divide up into four blocks, the s block, the p block, the d block, and the f block, based on electron arrangement.
- The s-block is all elements in groups 1 and 2.
- Groups 3-12 have transition metals and are called the d-block. They do not follow patterns as well as groups 1, 2, and 13-18. The number of valence electrons are harder to predict and they can have a variety of charges.
- Al, Ga, In, Sn, Tl, Pb, Bi are sometimes called "poor metals" because they don't have perfectly metallic properties
- Metalloids are the elements touching the zig-zag line, except aluminum which is a metal. These are commonly used in electronics as a semiconductor

Rare Earth Elements:

- The Lanthanide and Actinide series
- The Lanthanide series is part of Period 6
- The Actinide series is part of Period 7
- These are found in the f-block and are also called rare earth elements

There are also a few groups of elements that are nonmetals.

Halogens

- Group 17 of the table
- Have 7 valence electrons
- Form -1 ions

- Very reactive, especially with the alkali metals.

Noble Gases (Inert Gases)

- Group 18 of the PT.
- Octet of valence electrons (full valence shell)
- Tend not to form ions
- Inert (do not react)

p-block

- Groups 13-18 are called the p block
- The p-block has a few metals: Al, Ga, In, Sn, Tl, Pb, Bi, Po
- The p-block also contains metalloids and nonmetals

Once you know which group an element is in, the number of valence electrons that element has is predictable.

Once you know which group an element is in, the charge of the ion that element forms is likewise predictable.

Exercise Calcium is in which block?

Exercise Uranium is in which block?

Exercise Silicon is in which block?

3.2 Periodic Trends

Periodic Trends are patterns that appear on the periodic table.

4 factors that cause the trends

- Nuclear Pull (Z) - the number of protons
  - The protons pull on the outer electrons. The more protons, the more pull exerted by the nucleus on the outer electrons.

  Exercise Which of the following elements has the most nuclear pull? Carbon or Fluorine?
- Electron repulsion - size of the e^- cloud.
  - The more electrons in an atom's electron cloud, the more they are pushed away from each other, making a bigger cloud.
- Shielding electrons - all inner e^- shield the valence electrons from nuclear pull
  - Electrons on the inner shells feel the nuclear pull stronger than the valence electrons, which are farther from the nucleus
- Z_{eff} - the "effective" nuclear pull on outer electrons. This takes into account the shielding electrons which are taking most of the force.

Atomic Radius Trend

Atomic radius increases down a column because the valence electrons are in a farther energy level and decrease across a period because the nuclear pull is increasing and pulling the energy levels in.

Ionic Size

Metals ions are smaller than their atoms because metal ions lose electrons causing electron repulsion and smaller size.

Nonmetal ions are larger than their atoms because they are gaining electrons, causing more electron repulsion, and larger size.

Ionization Energy

The energy needed to pull an electron from an atom.

The greater the ionization energy, the more difficult it is to remove an electron.

This decreases down a group because there are more shielding electrons, so it takes less energy to "steal" and electron. This increases across a period because the nuclear pull on those electrons is increased with no extra shielding, so its takes more energy to get the electrons away.

Electronegativity

The ability of an atom to take an electron from another atom.

This decreases down a group because there are more electrons to shield the nucleus. This increases across a period because of increased Z.

Electron Affinity

The energy change that occurs when an atom acquires an electron.

Most atoms give off energy when gaining an electron, the more attracted an atom is to the new electron, the more energy released.

Therefore, the trend correlates with electronegativity.

Z_{eff}

The effective nuclear charge - the nuclear pull as felt by the valence electrons.

Equal to the number of protons in the nucleus minus the number of electrons that are between the nucleus and the valence electrons.

No change down a group, because even though nuclear pull has increased, you have more shielding e^-'s.

Increases across a period because nuclear pull is increasing and no additional shielding.

Reactivity

Most reactive corners of the PT are lower left and upper right.

This is because metals tend to donate electrons to obtain their octet. The most reactive metals are therefore the ones with the lowest ionization energy.

Nonmetals tend to gain electrons to obtain their octet. The most reactive nonmetals are on the upper right because they have the highest electronegativity.

Exercise Which is the smallest atom? Na, Li, or Be?

Exercise Which has the highest electronegativity? As, Sn, or S?

Exercise In the following pairs, which have the larger atomic radius? Mg or Ba, Cu or Cu^{2+}, S or S^{2-}.

Exercise In the following pairs, which has the higher ionization energy? Li or Cs, Ca or Br.

Chapter Problems

1. Explain why it takes more energy to remove the second electron from a lithium atom than it does to remove the first electron from a lithium atom.
2. How does the ionic radius of a nonmetal compare with its atomic radius? Explain why the change in radius occurs.
3. Experiments show that the electronegativity of phosphorus is 2.5, and the electronegativity of chlorine 3.5. Explain the difference in electronegativity using principles of atomic structure.
4. Explain how ionization energy changes as you move left to right across a period using principles of atomic structure.

4 Bonding and Compounds

Do you know any jokes about sodium hypobromite? NaBrO

4.1 Types of Bonds Overview

Chemical compounds are formed by the joining of two or more atoms. When atoms bond, their valence electrons are redistributed in ways that make the atoms more stable. The way the electrons are redistributed depends on the type of bond formed.

A chemical bond is a mutual attraction between the nuclei and valence electrons of different atoms that binds atoms together.

Ionic Bonds

- These bonds are the result of the electrical attraction between positive ions and negative ions.
- The ions are formed because atoms completely give up their electrons to other atoms.

Ionic Bonding Process:

1. In an ionic bond, electrons are transferred from one atom to another.
2. The transfer creates a positive ion and a negative ion.
3. Cations and anions are attracted to each other due to the electrostatic attraction between positive and negative ions, so they are bound together.

- These bonds usually occur between a metal and a nonmetal, creating an ionic compound, also known as a salt.
- Both ions end up with an octet of electrons in their valence shell.
- Salts are neutral because they have an equal positive and negative charge.
- Metals lose electrons and nonmetals gain electrons in an ionic compound.

Covalent Bonds

- These bonds are the result of the sharing of electron pairs between two atoms.
- In a covalent bond, the electrons are "owned" by both of the two bonded atoms.

The Covalent Bonding Process:

1. Covalent bonds are the result of sharing electrons between two atoms.
2. Because the atoms must stay together to share, molecules are formed.
3. The molecules are neutral because they have the same number of protons and electrons.

- Covalent bonds usually occur between two nonmetals.
- Covalent bonding results in individual molecules.

Metallic Bonding

- In pure metals or alloys, there are usually vacant valence orbitals. The vacant orbitals overlap from one atom to another, allowing the outermost electrons to roam freely throughout the entire metal.
- These are called delocalized electrons. These mobile electrons, a "sea of electrons", move throughout the entire metal.
- Metallic bonds are a result of the attraction between metal nuclei and the surrounding sea of electrons.

The Metallic Bonding Process:

1. Metal atoms have overlapping empty orbitals.
2. Each metal atom loses its valence electrons to roam freely throughout the metal.
3. The metal is held together because the free-floating electrons and positive metal cores are attracted to each other.

Exercise What type of bonding is present in phosphorus decoxide?

4.2 Ionic Nomenclature

3 main types of compounds

- Type 1 - Ionic Compounds
  - A positive ion and a negative ion.
  - A metal and a nonmetal.
  - A metal and a polyatomic ion.
- Type 2 - Covalent Compounds
  - Made of more than one nonmetal atom
- Type 3 - Acids
  - Made up of positive hydrogen ions paired with negative ions.
  - These compounds appear to be covalent but behave like ionic compounds.

Exercise What type of compound is Copper (II) bromide?

Ionic Formulas and Nomenclature

- Chemical formulas show the number of atoms in a compound.
- "Nomenclature" is a naming system.
- The "nomenclature" system for people is first and last names.
- The typical "nomenclature" system for marriages is that the man keeps his last name and the bride changes her last name.

How to name ionic compounds

1. Make sure the compound is ionic by looking for a metal.
2. Name the cation and then the anion. Remember that nonmetal monatomic ions end in "-ide".
3. Write the ion symbols with their charges.
4. Cross the charges over and take the absolute value. These numbers become the subscripts. Reduce the subscripts if they are divisible by an integer.

Exercise Write the chemical formula for Zinc nitride.

If a polyatomic ion needs a subscript, put parenthesis around the polyatomic ion to show that more than one polyatomic ion is present.

Exercise - Write the chemical formula for Tin(IV) Sulfate.

To name an ionic compound, name the positive ion then name the negative ion. Easy peasy. Remember nonmetal ions end in -ide.

Exercise What is $Zn(OH)_2$?

Transition Metal Names

- If the metal has more than one possible charge you must know which one it is.
- Do a reverse cross of the subscripts to determine the charge of the metal.

Exercise What is $NiPO_4$?

4.3 Covalent & Acid Nomenclature

Covalent compounds contain only nonmetals (also called molecular compounds).

To name a covalent compound

- name the first element
- then name the second one and change its ending to -ide
- Use prefixes to show how many atoms of each element you have.

Exercise What is the formula for Tetraphosphorus decasulfide?

Exercise What is PCl_5?

- Acids are an important class of hydrogen-containing compounds and are named in a special way.
- Acids are defined as substances whose molecules produce hydrogen ions when dissolved in water.
- When we encounter acids, it will be written with H as the first element.

Writing and Naming Acids

- composed of an anion connected to enough H^+ ions to totally neutralize or balance the anion's charge. Criss-cross!
- name of an acid is related to the name of its anion... Three acid naming rules:
  - If the anion ends with ide, the acid is hydro______ic acid.
  - If the anion ends with ate, the acid is ______ic acid.
  - If the anion ends with ite, the acid is ______ous acid.

Exercise What is HCN?

Exercise What is the formula for Dichromic Acid?

4.4 Mole Problems

1 mol is 6.022×10^{23} of anything

Molar Mass The molar mass concept works the same way with compounds as it did with pure elements. You simply add the molar mass of each atom within the formula.

We call this "molar mass" for molecular compounds and "formula mass" for ionic compounds. Generally, though, we use the term molar mass for atomic mass in grams of any compound, ionic or covalent, or any element.

Exercise Calculate the molar mass of H_2O

Exercise Calculate the molar mass of calcium chloride.

If parenthesis appear ina formula, the number outside the parenthesis multiplies by every atom inside the parenthesis, just like a coefficient in math.

Exercise Calculate the molar mass of $Ca(NO_3)_2$

Exercise How many moles are equal to 3.905×10^{23} formula units of calcium hydroxide?

Exercise How many nitric acid molecules are in 4.20 g of HNO_3?

4.5 Percent Composition

Percent composition of a compound:

$$\text{\% composition} = \frac{\text{\# of atoms of element} \times (\text{MM of element})}{\text{MM of compound}} \times 100$$

where "MM" is molar mass

You must know the CORRECT formula of the compound to calculate percent composition.

Exercise What is the % copper in copper(II) carbonate?

Round percent composition answers to 4 sig figs.

Hydrates are ionic compounds that can trap water in their crystalline structure when they form. The water is part of the structure, and it is a definite ratio of the compound.

Anhydrous compounds have no water in their crystalline structure.

To calculate the % water in the hydrate, use the same formula as before, but water is the part on top.

Exercise Calculate the % water in copper(II) sulfate pentahydrate.

Exercise Calculate the % water in magnesium sulfate heptahydrate.

4.6 Empirical & Molecular Formulas

Questions typically look like this:

- You are given the percent composition of a compound.
- You determine the formula based on the percentages.

Let's rhyme to solve these:

- Percent to mass
- Mass to mole
- Divide by least
- Multiply 'til whole

Exercise What is the formula of a compound that is 25.9% nitrogen and 74.1% oxygen?

So far you have calculated the simplest formula. We will now take this further. The empirical formula is the just the lowest possible ratio. The molecular formula, the actual makeup of a molecule, may be different.

If possible, you want to give the molecular formula. It is more descriptive of the actual molecular makeup. To do this, you must know the molar mass of the molecule.

Note: Ionic compounds never have molecular formulas, since the definition of the formula of an ionic compound is the lowest possible ratio. Only molecular, or covalent compounds, can have a molecular formula.

How to determine molecular formula:

1. Divide the true molar mass by the empirical formula's molar mass to get an integer.
2. Multiply the subscripts of the empirical formula by this integer.

Exercise A compound has the empirical formula CH. The molar mass of the compound is 78.110 g. What is the molecular formula of the compound?

4.7 Oxidation Numbers

"Oxidation numbers" are an accounting system used to keep track of electrons in a chemical reaction.

The oxidation state of a free element is 0.

The oxidation state for a monatomic ion is equal to its charge.

The algebraic sum of the oxidation numbers of all the atoms in a compound must be zero.

Similarly, the algebraic sum of the oxidation numbers of all the atoms in a polyatomic ion must equal the charge of the polyatomic ion.

Really useful rules:

- In compounds, the more electronegative element is always negative.
- In compounds, hydrogen is usually $+1$, unless it is bonded to a metal. In that case it is a hydride and the number is -1.
- In compounds, oxygen is usually -2. However, if it is a peroxide, it is -1. If it is bonded to fluorine, oxygen will be $+2$. This is rare.
- The oxidation number for alkali metals in compounds is always $+1$. The oxidation number for alkaline earth metals in compounds is always $+2$.

Oxidation numbers do not have to be the same ones found on the periodic table. In fact, they will not always be whole numbers! Rule 3 cannot be violated! Remember, oxidation numbers are just an accounting system for keeping track of electrons.

Exercise Write the oxidation numbers for each element in I_2.

Exercise Write the oxidation numbers for each element in MnO_4^-.

Chapter Problems

1. Write the correct chemical formula for manganese (III) oxide
2. What is the correct name of the ionic compound CsCl
3. Name the compound $Cd(C_2H_3O_2)_2$
4. Write the formula for silver hydrogen carbonate
5. Write the formula for pentaphosphorus decoxide.
6. Write the name of the compound $Au_2(SO_4)_3$
7. Write the formula for the compound sodium chlorate
8. Write the formula for hypochlorous acid
9. What is the ionic/covalent name for H_3PO_4 and the acid name.
10. How many molecules are in 25.5 g of iron(III) hydroxide?
11. How many ions are in 5.58×10^{24} formula units of tin(IV) nitrate.
12. What is the percent sulfur in sodium thiosulfate?
13. What mass of oxygen would you have in a 7.5 g sample of zin sulfite?
14. What is the empirical formula of a compound that is 24.3% carbon, 4.1% hydrogen, and 71.6% chlorine?
15. What is the molecular formula of a gas that is 30.5% nitrogen and 69.5% oxygen? The molecular mass of the compound is 91.8 grams.
16. Determine the oxidation number of the sulfur atom in SO_3
17. Give oxidation numbers for Pt in $PtCl_6^{2-}$.

5 Reactions

I told a chemistry joke. There was no reaction.

5.1 Balancing Equations

When chemical reactions occur:

- Bonds are broken and new bonds form.
- Energy is produced or absorbed.
- New compounds are formed, or compounds decompose to their elements.
- The Law of Conservation of Mass is obeyed.

Symbols in Equations:

- yield $\rightarrow$
- Sometimes a reaction occurs then stops
- reversible reaction $\leftrightarrow$
- Sometimes the reaction goes back and forth between product and reactant.
- solid or precipitate (s)
- gas (g)
- liquid (l)
- water solution (aq)
- heat Δ

A catalyst is a chemical that speeds up a reaction, but is not actually used up.

Exercise Put the reaction $Ca(OH)_2$ (s) $\rightarrow$ CaO (s) + H_2O (l) in words.

Balancing Chemical Equations is necessary so that the correct amount of reactants can be determined and the amounts of the products can be predicted.

It also satisfies the law of conservation of mass.

How to balance a chemical reaction:

- Write the equation with the correct formulas and symbols.
- Add coefficients to the formulas to make the number of atoms of each element on both sides of the equation the same. A coefficient is a whole number before a chemical formula.
- You may not add coefficients to the middle of a formula.
- You may not change the subscript of a correctly written formula.

Exercise Balance FeS + HCl $\rightarrow$ $FeCl_2$ + H_2S

Exercise Write the balanced equation for lithium chlorate decomposing into lithium chloride and oxygen gas.

5.2 Synthesis & Decomposition

Synthesis Reactions - more than one reactant and only one product.

Decomposition Reactions - one reactant and more than one product.

Synthesis Rules

- Reaction between 2 nonmetals produces a common covalent compound.
- Reaction of a metal and a nonmetal produces an ionic compound.
- Reaction of a metal oxide and water produces a metal hydroxide.
- Reaction of a metal oxide with carbon dioxide produces a metal carbonate.
- Reaction of a metal chloride with oxygen produces a metal chlorate.
- Reaction of a nonmetal oxide with water produces an acid in solution.

Exercise Write the equation for carbon burning.

Decomposition Rules

- Decomposition of a binary compound produces two elements.
- Decomposition of a metal carbonate produces a metal oxide and CO_2
- Decomposition of a metal hydroxide produces a metal oxide and water
- Decomposition of a metal chlorate produces a metal chloride and oxygen
- Decomposition of an oxyacid produces a nonmetal oxide and water. The oxidation number of the nonmetal remains the same.

Exercise Write the equation for sodium carbonate decomposing when heated.

5.3 Single Replacement, Double Replacement, & Combustion

Single replacement is when an element replaces another element in a compound.

An element will replace another element in a compound if the lone element is more reactive than the element in the compound.

Single Replacement Rules

- Replacement of a metal by a more reactive metal.
- Replacement of hydrogen in water by a group 1 metal produces a metal hydroxide and H_2
- Replacement of hydrogen in water by a group 2 metal produces a metal oxide and H_2
- Replacement of hydrogen in an acid by a metal more active than H. Metal replaces hydrogen as if it were a metal.
- Replacement of a nonmetal (usually a halogen) in a compound by a more reactive nonmetal.

Exercise Write the equation of Fluorine + Potassium Bromide

Double replacement reactions occur when elements in two compounds exchange places to make two new compounds.

These reactions occur between ions in aqueous solutions and produce at least one of the following - a precipitate, a gas, or water

If a product is insoluble, it is a precipitate.

Note that hydrogen sulfide is a gas.

H_2CO_3 decomposes to carbon dioxide and water and H_2SO_3 decomposes to sulfur dioxide and water.

Ammonium hydroxide decomposes to form ammonia gas (NH_3) and water.

Exercise Write the equation for Sodium bicarbonate + hydrochloric acid

The burning of a hydrocarbon in O_2 to produce heat is combustion.

When hydrocarbons burn in excess oxygen, the products are always carbon dioxide and water.

If there is too little oxygen, carbon monoxide is produced. Carbon monoxide is highly toxic!

5.4 Reaction Rates

Reversible reactions - some reactions continue until the products being to react and form the reactants again.

Equilibrium - the point in a reaction when the rate of the forward reaction is equal to the rate of the reverse reaction.

Reaction rate is how fast a chemical reaction will occur.

Molecular collisions are necessary for two substances to react. Many factors affect how often molecules collide, and therefore affect the reation rate.

There are several factors that affect the speed of a reaction:

- Temperature of reactants
- Concentration of reactants
- Presence of a catalyst or an inhibitor
- Surface area of reactants

Raising the temperature of a substance causes its molecules to move faster. Faster molecules will collide more often, increasing the speed of the reaction. Therefore, higher temperature results in a faster reaction.

Increasing the concentration results in more reactants in a given space, so you will have more collisions per unit time.

A catalyst is a substance that helps molecules come together. It is not used up in a reaction, it just speeds the reaction.

An inhibitor prevents molecules from reacting with each other, thus slowing the reaction rate.

Reactions depend on collisions. The more surface area on which collisions can occur, the faster the reaction.

Some reactions would never happen unless energy is added to the system.

5.5 Redox Reactions

Redox reactions are reactions in which elements' oxidation numbers (charges) change due to moving electrons.

Redox stands for reduction-oxidation reactions. Electrons move from one atom to another or from one ion to another. This means the oxidation numbers of elements change from the reactant to the product side of an equation.

Many types of reactions classify as redox. This isn't a totally separate type of reaction.

Oxidation is loss of electrons and reduction is gain of electrons.

Loss of electrons means the charge goes up and gain of electrons means that the charge goes down.

The element that is oxidized comes from the reactant that is the reducing agent, and the element that is reduced comes from the reactant that is the oxidizing agent.

Exercise What is being oxidized, reduced, and the reducing agent, and the oxidizing agent in $2Na + Cl_2 \rightarrow 2NaCl$

5.6 Net Ionic Equations

A net ionic equation does not show the ions that don't change (i.e do not include the ions that stay aqueous)

Steps for Writing Net Ionic Equations:

1. Write the balanced equation with all states labeled.
2. Split any aqueous ionic or strong acids into ions.
3. Cancel out any ions that appear on both sides of the arrow.

Note that the diatomic elements in aqueous form are no longer diatomic.

Exercise Write the net ionic equation for a solution of lead(II) nitrate is added to hydrochloric acid.

Chapter Problems

1. Balance the equation $H_2 + O_2 \rightarrow H_2O$
2. Write a balanced equation for ammonium bicarbonate yielding ammonia, water, and carbon dioxide
3. Identify the reaction of magnesium oxide plus water as synthesis or decomposition and complete and balance the equation
4. Make up a decomposition equation
5. Use the activity series to determine if the single replacement reaction for mercury and nickel (II) carbonate occurs. If it occurs, write the equation and balance, otherwise write no reaction.
6. Write the balanced equation for lead (II) nitrate + sodium chloride
7. Determine the type of reaction for propane (C_3H_8) oxygen and write the balanced equation for this reaction
8. Determine the type of reaction for sodium and water and write the balanced equation.
9. Write the net ionic equation for chlorine gas reacting with a solution of potassium iodide.
10. Write the full ionic equation for a solution of copper (II) chloride reacting with a solution of lead (II) acetate.

6 Stoichiometry

What is a chemist's favorite plant? Stoichiome-tree.

6.1 Stoichiometry

Stoichiometry is the measurement and calculation of the amounts of reactants and products in chemical reactions.

Balanced chemical equations represent the relationship between the number of moles of reactants and the number of moles of products.

Mole ratio is the conversion factor for any two reactants or products in a chemical reaction.

Mole to Mole Problems

1. Write the balanced equation
2. Write the strategy (molar road map)
3. Set up the correct calculation

Stoichiometry Road Map - Grams of A $\leftrightarrow$ Moles of A $\leftrightarrow$ moles of B $\leftrightarrow$ grams of B

Exercise Iron reacts with carbon dioxide to form iron(III) oxide and carbon monoxide. How many moles of carbon dioxide are needed to produce 2.2 moles of iron(III) oxide?

Exercise How many grams of magnesium chloride are produced when 0.500 moles of magnesium reacts with an excess of hydrochloric acid?

Exercise How many moles of zinc sulfate are produced when 4.55 g of zinc reacts with an excess of sulfuric acid?

Exercise Calcium carbonate reacts with phosphoric acid to produce calcium phosphate, carbon dioxide, and water. Calculate the number of grams of CO_2 formed when 0.47 g of water is produced.

6.2 Percent Yield, Limiting Reactant, & Gas and Solution Stoichiometry

- Stoichiometric calculations are based on ideal reactions.
- Many reactions do not go to completion and not as much product is produced as expected.

There are two different yields:

Theoretical Yield: what stoichiometry predicts.

Actual Yield: What is actually produced and measured in the lab.

Percent Yield:

$$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Exercise What is the percent yield for the reaction

$$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$$

in which 214 g of chlorine react with an excess of potassium bromide to produce 412 g of bromine?

The limiting reactant is the chemical that is used up first in a chemical reaction. It limits the amount of product that can be made.

The other reactant(s) is/are called the excess reactant(s).

If you are given the amounts of both reactants and asked to predict the amounts of products, you must base your answer on the limiting reactant.

Exercise The balanced equation for the reaction between 50.0 g of silicon dioxide and 50.0 g of carbon is?

$$SiO_2 + 3C \rightarrow SiC + 2CO$$

Assuming the reaction is 100% efficient, what is the excess reactant and how much in excess is it?

We can also involve gases in stoichiometry:

- Liters are used for a gas as STP
- STP means "standard temperature and pressure", which we'll define as 0˚C and 1 atm
- Use 22.4 L/mol. This applies to ANY GAS AT STP.

Exercise How many liters of 3.4 M copper(II) sulfate are needed to react fully with 2.00 grams of zinc?

Exercise How many atoms of oxygen are there in a 3.0 mole sample of $Mg(ClO_3)_2$?

Chapter Problems

1. Ammonium nitrate decomposes into nitrogen gas, water, and oxygen gas. Write and balance the equation. How many moles of oxygen gas are produced when 8.14 moles of ammonium nitrate decompose?
2. Write the balanced equation for the decomposition of sugar, $C_{12}H_{22}O_{11}$, into carbon and water. How many moles of carbon would be produced if you started with 2 moles of sugar?
3. C_5H_{11} reacts with oxygen in a combustion reaction. Write the balanced equation for this reaction. How many grams of C_5H_{11} must be burned to produce 1.25 moles of water?
4. Sucrose, $C_{12}H_{22}O_{11}$, decomposes into carbon and water vapor. Write the balanced equation for this reaction. If 4.00 g of sucrose decomposes, how many grams of water vapor are produced?
5. What volume does 0.00353 moles of He gas occupy at 0°C and 1.0 atm?
6. A 3.7 mole sample of Na_2SO_4 would have how many atoms of sodium?
7. Calcium bromide reacts with 5.100 g of aluminum oxide in a double replacement reaction. How much calcium bromide must be present for all of the aluminum oxide to react?
8. What is the percent yield if 4.30 g of potassium is reacted with an excess of iodine and 13.78 g of potassium iodide is formed?
9. The compound cisplatin, $PtCl_2(NH_3)_2$ is effective in treating some types of cancer. It can be synthesized using the following skeleton equation

   $$K_2PtCl_4 + NH_3 \rightarrow KCl + PtCl_2(NH_3)_2$$

   Suppose a researcher needs exactly 5.00 grams of cisplatin for an experiment. How much K_2PtCl_4 would she need to start with, assuming she has an excess of ammonia?
10. How many liters of oxygen gas can be obtained from the thermal decomposition of 500.0 g of potassium chlorate? How many grams of the other product are formed?

7 VSEPR/IMFs

Names Bond, Ionic Bond. Taken not shared.

7.1 Types of Bonding

Chemical components are formed by the joining of two or more atoms. When atoms bond, their valence electrons are redistributed in ways that make the atoms more stable. The way the electrons are redistributed depends on the type of bond formed.

A chemical bond is a mutual attraction between the nuclei and valence electrons of different atoms that binds atoms together.

Ionic bonds are the result of the electrical attraction between positive and negative ions.

The ions are formed because atoms completely give up their electrons to other atoms.

Ionic Bonds

- These bonds usually occur between a metal and a nonmetal, creating an ionic compound, also known as a salt.
- Both atoms end up with an octet of electrons in their valence shell.
- Salts are neutral because they have an equal positive and negative charge.
- Metals lose electrons and nonmetals gain electrons in an ionic compound.

Covalent Compounds

- These bonds are the result of the sharing of electron pairs between two atoms.
- In a covalent bond, the electrons are "owned" by the two bonded electrons.

Covalent bonds usually occur between two nonmetals and results in individual molecules.

Metallic Bonding

- In pure metals or alloys, there are usually vacant valence orbitals. The vacant orbitals overlap from one atom to another, allowing the outermost electrons to roam freely throughout the entire metal.
- These are called delocalized electrons. These mobile electrons, a "sea of electrons", move throughout the entire metal.
- Metallic bonds are the result of the attraction between metal nuclei and the surrounding sea of electrons.

Exercise What type of bonding is present in carbon dioxide?

7.2 Bonding

A chemical bond is an attractive force between atoms or ions that binds them together as a unit. Bonds form in order to decrease potential energy and increase stability.

What is Chemical Bonding?

- A chemical bond is formed when electrons are shared or given between two or more atoms.
- The electrons involved are only the outermost electrons - the valence electrons.
- Chemical Bond - a link between atoms that holds them together.

Keeping Track of Electrons:

- The electrons responsible for the chemical properties of atoms are those in the outer energy level.

- Valence electrons - The s and p electrons that are in the highest energy level.
- Core (or shielding) electrons - those in the energy levels below.

Remember atoms in the same column have the same outer electron configuration and have the same number of valence electrons.

In the s block, the number of valence electrons is the group number, in the d block the number of valence electrons varies and isn't always predictable, and in the p block the number of valence electrons is the group number minus 10.

Exercise How many valence electrons does phosphorus have?

Atoms typically bond to form an octet in their valence level. All atoms want this stability. This is also called "noble gas configuration".

When an atom gains or loses electrons, it is an ion. Loss of electrons is a cation and is positively charged. Gain of electrons is an anion and is negatively charged.

Intramolecular bonds hold atoms to atoms - they are your ionic, covalent, and metallic bonds.

Intermolecular Bonds hold two or more molecules/ions together. They are your hydrogen, dipole-dipole, ion-dipole, and london dispersion forces.

Ionic Bond

- An ionic bond is formed when electrons are transferred from one atom to another. This creates positive and negative ions.
- When one or more electrons is transferred, you get both a positive and negative ion.
- Since they have opposite charges, they are attracted to one another. This is called an "electrostatic attraction".
- Ionic bonds typically form with a metal and a nonmetal.
- Ionic substances are sometimes caleld salts.
- Overall, salts are neutral. They have equal amounts of positive and negative charge.

What are ionic compounds?

- Because of their valence electron structure, metals lose their electrons, and nonmetals gain electrons.
- This is why metal ions have a positive charge and nonmetal ions have a negative charge.

Formula Unit

- A formula that tells the ratio of ions in an ionic compound.
- The smallest part of an ionic compound that still has the composition of the compound.

Lewis Dot notation can be use to visualize ionic compounds and how they form.

Properties of Ionic Compounds

- High melting points/boiling points - it takes a lot of energy to break strong bonds.
- Hard, brittle solids
- Many are soluble inw ater
- When dissolved, free ions float and conduct electricity
- Form crystalline solids

Do they conduct?

- Conducting electricity is allowing charges to move.
- In a solid, the ions are locked in place - ionic solids are insulators.
- When melted, the ions can move around.
- Melted ionic compounds conduct.

- Dissolved in water they can conduct.

In order for electrons to be transferred, one element must be much more electronegative than the other. They must have an EN difference of more than 1.7. In general, most combinations of metal+nonmetal will have this great ΔEN.

Covalent bonds are a bond that results from the sharing of electrons. They are made of molecules instead of crystal lattice and usually occur between two nonmetals. When two atoms do not have a big ΔEN, they will share electrons. There are varying degrees of how electrons can be shared.

- Shared equally: nonpolar covalent bond. The EN values are almost equal, a difference less than 0.5
- Shared unequally: polar covalent bond. The EN values are not equal, but not different enough to form an ionic bond.

In nonpolar covalent bonds, electrons are shared equally, the molecule overall is neutral.

In polar covalent bonds, electrons are not shared equally. The more electronegative atom attracts the electrons more, forming a partially negative region of the atom. The less electronegative atom becomes partially positive.

Properties of Covalent Bonds

- No ions, no charges, do not conduct electricity.
- Weak attraction between molecules.
- Usually liquids or gases at room temperature.
- If solid, have low melting points.
- Amorphous Solid - do not have a regular/repeating pattern.

Most bonds are a blend of ionic and covalent characteristics. Difference in electronegativity determines bond type.

Metallic bonds occur between metal atoms. Bonding due to a "sea of electrons" - electrons that are not bound to one specific atom, they are able to move around the substance from atom to atom. Accounts for properties of metals and metal alloys.

Metals are

- Malleable
- Ductile
- Good at conducting heat and electricity

Properties are due to the free-floating electrons.

Exercise What type of bond is CH_4?

Lewis Structures:

- Lewis structures can be drawn for both ionically and covalently bonded compounds.
- Just keep in mind ionically bonded salts will contain ions.
- Covalently bonded molecules will show shared electrons.

Types of Covalent Bonds:

- Single Bond - one pair of electrons is shared; represented by a single line drawn between two atoms.
- Double Bond - two pairs of electrons shared; represented by two lines drawn connecting the two atoms.
- Triple Bond - three pairs of electrons shared; represented by three lines drawn connecting the two atoms.

Multiple Bonds: usually formed by C, N, O, P, S

Triple bonds are stronger than double bonds and double bonds are stronger than single bonds. It takes more energy to break a double bond than a single bond, and more energy to break a triple bond than a double or single bond.

Multiple bonds increase the electron density between two nuclei. As the electron density increases, the repulsion between the two nuclei decreases. An increase in electron density also increases the attraction each nucleus has for the additional bonding electron pairs. The nuclei move closer together and the bond length is shorter for a double bond than a single bond.

Predicting the Arrangement of Atoms within a molecule:

- H is always a terminal atom. H is ALWAYS connected to only one other atom.
- The element with the lowest electronegativity is the central atom in the molecule. Put other atoms around the central atom.
- Find the total # of valence electrons by adding up group #'s of the elements. For ions add electrons for negative charges and subtract electrons for positive charges. Divide by two to get the number of electron pairs available to go around.
- Use a pair of electrons to connect each terminal atom to the central atom.

Usually central atoms will have 4 things around them, so spread atoms at 90 degree angles.

- Place lone pairs about teach terminal atom to satisfy the octet rule.
- Left over pairs are assigned to the central atom. If the central atom is from the 3rd of higher period, it can accommodate more than four electrons pairs.
- If the central atom is not yet surrounded by four electron pairs, convert one or more terminal atom lone pairs to pi bonds pairs. Not all elements form pi bonds! Only C, N, O, P, and S.

Remember, only C, N, O, P, S are able to form multiple bonds.

Exceptions to the octet Rule:

- Electron Deficient: less than 8 electrons
  - Hydrogen: 2 in outer energy level
  - Boron: 6 in outer energy level
  - Beryllium: 4 in outer energy level
- Exceed Octet: more than 8
  - anything in 3rd period or heavier
  - because d-orbitals are available and add extras to the middle atom.

Often times there is more than one possible way for atoms to bond together in a given molecule.

VSEPR:

- Valence Shell Electron Pair Repulsion
- We've already discussed this - areas of electrons around a central atom tend to spread out to reduce electrostatic repulsion
- Can be used to predict 3-D shape of molecules.

Areas of Electron Density:

Bonded electrons or unbonded electrons (lone pairs). These areas spread as far apart from each other as possible.

Molecules are nonpolar if they have only one kind of terminal atom and no lone electron pairs on the center atom. Molecules are polar if they have more than one kind of terminal atom or at least one lone electron pair on the central atom.

Hybridization is the mixing of different types of atomic orbitals to produce a set of equivalent hybrid orbitals. For assigning hybridization, we tell what type of orbitals are mixed.

We will use the following key now for describing shapes: "A" represents the central atom, "X" represents the atoms attached to the central atom, and "E" represents a lone pair of electrons on the central atom.

- 2 bonding regions, 0 lone pairs - AX_2 - linear - usually nonpolar - hybridization: sp - bond angle: 180°

- 3 bonding regions, 0 lone pairs - AX_3 - trigonal planar - usually nonpolar - hybridization: sp^2 - bond angle: $120°$
- 2 bonding regions, 1 lone pair - AX_2E - bent - always polar - hybridization: sp^2 - bond angle: $< 120°$
- 4 bonding regions, 0 lone pairs - AX_4 - tetrahedral - usually nonpolar - hybridization: sp^3 - bond angle: $109.5°$
- 3 bonding regions, 1 lone pair - AX_3E - trigonal pyramidal - always polar - hybridization: sp^3 - bond angle: $107°$
- 2 bonding regions, 2 lone pairs - AX_2E_2 - bent - polar - hybridization: sp^3 - bond angle: $104.5°$
- 5 bonding regions, 0 lone pairs - AX_5 - trigonal bipyramidal - usually nonpolar - hybridization: under debate - bond angle: $90°, 120°, 180°$
- 4 bonding regions, 1 lone pair - AX_4E - seesaw - polar - hybridization: under debate - bond angles: $< 90°, < 120°, < 180°$
- 3 bonding regions, 2 lone pairs - AX_3E_2 - T-shaped - polar - hybridization: under debate - bond angles: $< 90°$
- 2 bonding regions, 3 lone pairs - AX_2E_3 - linear - polar - hybridization: under debate - bond angle: $180°$
- 6 bonding regions, 0 lone pairs - AX_6 - octahedral - usually nonpolar - hybridization: under debate - bond angles: $90°, 180°$
- 5 bonding regions, 1 lone pair - AX_5E - square pyramidal - polar - hybridization: under debate - bond angles: $< 90°, < 180°$
- 4 bonding regions, 2 lone pairs - AX_4E_2 - square planar - polar - hybridization: under debate - bond angles: $90°$
- 3 bonding reginos, 3 lone pairs - AX_3E_3 - T-shaped - polar - hybridization: under debate - bond angles: $< 90°$

Hybridization: The mixing of different types of atomic orbitals to produce a set of equivalent hybrid orbitals.

Polarity - bonds can be polar while the molecule isn't and vice versa.

Molecular Polarity - if a central atom has no lone pairs of electrons and all surrounding bonds are identical, then the molecule is nonpolar. Even though a molecule might have polar bonds within it, if those polar bonds cancel each other out, the molecule is nonpolar.

Molecules that have lone pairs of electrons on the central atom and.or different types of terminal atoms attached to the central atom are considered polar molecules. The charge is unevenly distributed throughout the molecule.

Exercise Is O_3 polar or nonpolar?

There are many types of bonds that hold that hold molecules and molecules or molecules and ions together. These forces are incredibly important.

London Dispersion Forces

- These are the forces that exist among non-ionic and non-polar substances.
- They exist among noble gases and nonpolar molecules.
- These forces are weak.

They are caused by an instantaneous dipole formation in which electron cloud becomes asymmetrical, and the molecules are slightly attractive to each other. This is the weakest intra- and intermolecular forces.

When you are comparing two substances that both have dispersion forces:

- The substance with more electrons has stronger dispersion forces since its electron cloud is larger and more polarizable.

Heavy noble gases have stronger dispersion forces than lighter noble gases.

Dipole-dipole forces

- These forces exist between molecules that have permanent dipole moments.
- Look for molecules with lone electron pairs on the central atom or different types of terminal atoms.
- These molecules have an uneven distribution of charge and therefore have attraction to each other.

Dipole-dipole forces are stronger than dispersion forces since the polarity in these molecules is permanent.

Hydrogen bonding is a special subset of dipole-dipole forces that exist only in H-N, H-O, and H-F.

This results in a partially positive pole and partially negative pole.

- When two or more water molecules are near each other, the weak positive hydrogen atom of one molecule will be attracted to the weak negative oxygen atom of the other molecule.
- This attraction between molecules is called hydrogen bonding.
- After many hydrogen bonds are formed, you have a weak force holding all the water molecules to each other.
- Hydrogen bonding is the reason water freezes into ice crystals of a certain repeating shape.
- Remember that hydrogen bonding can occur between any molecules that contain O-H, N-H, F-H bonds.

Hydrogen bonds have a high boiling point. It takes a large amount of energy to boil water into vapor because of the bonds holding the molecules together. The bonds must be broken in order for the liquid to change to a gas.

Hydrogen bonds also have high surface tension.

Water molecules are attracted to the glass because glass molecules are also polar, but not attracted to nonpolar plastic, so there is no meniscus in a plastic graduated cylinder.

Water can be drawn up into a thing glass tube with no effort because of the attraction between the water and glass molecules.

Surface tension can be decreased by adding a surfactant - this type of substance interferes with hydrogen bonding.

The Last IMF - ion-dipole forces.

- Attraction that helps ionic compounds dissolve in a polar substance.
- Think of salt water.

Exercise What IMF is present in CI_4?

Chapter Problems

1. Classify the bond between Mg and H and indicate if it is a full or partial charges, if any are present.
2. Explain the difference between dipole-dipole forces and hydrogen bonds.
3. Given data that phosphorus trichloride has a higher melting point than phosphorus pentachloride, explain why.
4. Why is silicon tetrabromide a liquid at room temperature while silicon tetraiodide is a solid at room temperature?
5. Draw a lewis structure for CS_2 and give the number of valence electrons.
6. Draw a lewis structure for CO_3^{2-} and give the number of valence electrons.

8 States of Matter

You has mass, you have volume. You matter.

There are three states of matter.

Solid:

- Matter that has both a definite shape and definite volume.
- Molecules or atoms are very close together and can only vibrate a little.
- They do not move past each other.

Liquid:

- Matter that has a distinct volume but no specific shape.
- Molecules or atoms are close together but have the ability to slide across one another very easily.

Gas:

- Matter that has no fixed volume or shape. It conforms to the volume and shape of its container.
- Its molecules or atoms are very far apart from each other and move very fast.

Compression - forcing a substance into a smaller volume.

- Gases are very compressible because of their empty space.
- Liquids have very little compressibility.
- Solids have almost no compressibility.

Density Comparison:

- If you consider the solid, liquid, and gas state of one particular substance, this rule holds true in most cases:
- Solid is more dense than liquid and liquid is more dense than gas.

Two Types of Solids:

Crystalline Solids

- molecules are packed together in a predictable way. THey are arranged in an orderly, geometric, three dimensional structure. The smallest repeating part of a crystalline structure is called a unit cell.

  Atomic Solids
- Unit particles are atoms.
- Noble gases are atomic solids when they are cooled to solid state. Usually very soft because they have weak IMFs.

  Molecular Solids
- Units are molecules, held together by weak IMFs. Low melting points.

  Covalent Network Solids
- Form a 3-D covalent network, very strong. High melting points.

  Ionic Solids
- Crystal lattice is formed from alternating cations and anions.
- High melting point and hardness.

- Always solids at room temperature.

  Metallic Solids

- Atoms are surrounded by mobile valence electrons.
- Malleable, ductile, conductors.

Amorphous Solids:

- particles are not arranged in a regular repeating manner.
- Amorphous means "without shape"

Liquids:

- Fluidity - liquids have the ability to flow
- Viscosity - the measure of the resistance of a liquid to flow.
- Liquids with big, complex molecules tend to be very viscous.
- Viscosity decreases with increasing temperature.

Buoyancy:

- The upward force a liquid exerts on an object.

Phase Changes - matter can change from one phase to another by adding or removing energy. There are six phase changes.

Phase Changes That Require Energy

- Melting - solid changing to liquid.
- Vaporization - liquid to gas, occurs when molecules have enough energy to escape the pull of the other molecules.
- Sublimation - solid changing directly into gas

Phase Changes That Release Energy

- Condensation - gas to liquid.
- Freezing - liquid to solid, achieved by removing heat.
- Deposition - gas directly to solid - achieved by removing heat.

Boiling is heating a liquid to the temperature at which all molecules have enough energy to escape and vaporize. Evaporation is the vaporization of surface molecules; very slow. This does not occur at high temperatures.

A phase diagram shows what phase a substance will be in at a certain temperature and pressure. Pressure is usually measured in atmospheres.

Triple point - the point on a phase diagram that shows the temperature and pressure combination at which three phases of a substance can coexist.

Critical point - temperature and pressure combination above which a vapor cannot be liquefied under any circumstances.

When energy/heat is added to or removed from a substance, two things could happen: temperature changes or phase change.

How do you know how much energy is needed for a change?

First off.

- Q for heating/adding energy is always positive.
- Q for cooling/releasing energy is always negative.

For a single phase, use the formula $q = mc\Delta T$, where q is heat in Joules, m is mass in grams, c is the specific heat, and ΔT is the change in temperature.

For a single temperature, use the formula $g = mol \cdot \Delta H$.

The heat needed to melt or freeze is the latent heat of fusion and the heat needed to boil/condense is the latent heat of vaporization.

Exercise How much energy is needed to convert 153 grams of ice at -15° C to steam at 125° C? The molar mass of water is 18.016 g/mol.

Chapter Problems

1. Which liquid is more viscous at room temperature, water of molasses? Explain your reasoning.
2. How is it possible that a pile of snow can slowly shrink even on days when the temperature never rises above the freezing point?
3. How much heat is required to warm 225 g of ice from $-46.8°$C to $0.0°$C, melt the ice, warm the water from $0.0°$C to $100.0°$C, boil the water, and heat the steam to $173.0°$C?

9 Gas Laws

A funny gas? He He He.

9.1 Kinetic Molecular Theory, Temperature, and Pressure

- A gas has no definite shape or volume.
- They adapt to the shape and volume of their container.
- Ideal Gases are imaginary gases that comply with all the postulates of the Kinetic Molecular Theory.
- Gas Laws attempt to explain the behavior of gases under certain conditions.

The Kinetic Molecular Theory

- Gases are made up of tiny particles.
- Gas particles move randomly, in straight lines in all directions and at various speeds.
- The forces of attractions or repulsion between two gas particles are extremely weak or neglibible, except when they coolide.
- When gas molecules collide, the collisions are elastic.
- The average kinetic energy of a molecule is proportional to the Kelvin temperature. Gases at higher temperatures have higher kinetic energies.

Characteristics of Gases

- Expansion - gases will expand to fill their containers since they have no definite shape or volume.
- Fluidity - gases have the ability to flow and be poured as liquids are.
- Low Density - gases have low density because the particles are spread far apart.
- Compressibility - gas particles can be made to occupy a smaller space by decreasing the volume of the container.
- Diffusion - gases spread out and mix with each other without agitation.

Avogadro's Principle

- Equal volumes of gases contain equal numbers of moles of those gases if the temperatures and pressures are the same.
- The volume occupied by one mole of any gas is 22.4 liters at standard temperature and pressure. This is called the molar volume/

Temperature Conversions:

- $T_K = T_C + 273$
- $T_C = T_K - 273$
- $T_F = (9/5)T_C + 32$
- $T_C = (5/9) \cdot (T_F - 32)$

Exercise Convert -20° C to K.

- Absolute zero is 0 K.
- At absolute zero, matter stops moving.
- Atoms/molecules in a solid, which usually vibrate, come to a complete stop.

Pressure (the force of a gas acting on the walls of its container) is measured in several different units.

- atm - atmospheres
- mm Hg - millimeters of mercury
- torr - torr
- Pa - pascals
- kPa - kilopascals
- psi - pounds (force) per square inch

Atmospheric pressure varies from day to day and is measured with a barometer.

1 atm of pressure is equal to 760 mm Hg, 760 torr, 101.325 kPa, 14.7 psi.

Exercise Convert 800. mm Hg to atm.

9.2 Gas Laws & Density

Boyle's Law states that the volume of a gas varies inversely with the pressure if the temperature is held constant.

- Boyle discovered that for any given ideal gas, the product of pressure and volume is always an exact constant.
- $P \cdot V =$ constant
- So, even if you change the pressure and volume of a gas, the product will still be the same.
- $P_1V_1 = P_2V_2$
- Remember, in Boyle's Law, temperature is held constant.

Exercise A syringe has 10.0 mL of gas inside and the pressure is 1.00 atm. If pressure is applied and the volume decreases to 4.8 mL, what is the final pressure of the gas inside?

Charles' Law states that the volume of a gas varies directly with the Kelvin temperature if the pressure is held constant.

- Charles discovered that volume divided by temperature is a constant.
- V/T = constant
- So, if you change the volume and temperature of a sample of gas, V/T will always be the same number.
- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
- Remember, this only applies when pressure is held constant and temperature is in Kelvin.

Exercise To what temperature Kelvin must 7.98 cm^3 of oxygen be cooled, to reduce its volume to 5.00 cm^3 if it is initially at STP and pressure does not change?

Gay-Lussac's Law states that the pressure of a gas varies directly with temperature if the volume is held constant. Just like Charles' Law, the temperature must be in Kelvin.

- Gay-Lussac discovered that for any given mass of an ideal gas, the pressure divided by temperature (in Kelvin!) was always a constant.
- P/T = constant
- So if you change the pressure and temperature of a gas, the press/temp will still be the same.
- $\frac{P_1}{T_1} = \frac{P_2}{T_2}$
- Remember, now volume is held constant.

Exercise If you cap a 2 L coke bottle containing air, and the temperature changes from 25°C to 35°C, what is the pressure on the inside wall of the bottle? Assume the initial atmospheric pressure when you capped the bottle was 728 mmHg.

The three laws can be combined into one law that can always be used when conditions are changed. We use this equation to figure out the new pressure, temperature, or volume of a gas if the initial conditions are known.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Exercise How much pressure must be applied to 68 L of a gas at STP to reduce its volume by half if the temperature is raised to 20.°C?

The Ideal Gas Law describes the conditions of an ideal gas in terms of pressure, temperature, volume, and the number of moles of a gas. Ideal gas law does not involve changes in conditions.

$$PV = NRT$$

R is the "Ideal Gas Constant" and is equal to 0.08206 L· atm/ mol · K

Exercise At which temperature would 0.0828 moles of hydrogen have a pressure of 1.00 atm and a volume of 55.0 L?

Real Gases do have forces of attraction and the molecules do have volume. Real pressure is lowered than what is predicted due to IMFs, especially for polar molecules or when hydrogen bonding is present. Real volume is higher than what is predicted due to molecular volume being significant, especially for larger molecules of gas. Gases act "most ideally" at high temperatures and low pressure.

We can relate the molar mass of a gas with density.

$$MM = \frac{DRT}{P}$$

Exercise If the density of a gas is 1.2 g/L at 745 torr and 20.°C, what is its molar mass?

Dalton's Law of Partial Pressures shows the pressure of a mixture of gases is simply equal to the sum of the partial pressure of each gas.

$$P_T = P_1 + P_2 + P_3 + \dots$$

When you collect a gas by water displacement, the collected gas also contains water vapor. There is more water vapor at higher temperatures.

Exercise A student collects 89 mL of oxygen gas by bubbling it through water. The pressure reading that day is 103.2 kPa and the temperature is 20.°C. Determine the number of moles of gas collected. At 20.°C, the partial P of water vapor is 2.3 kPa.

Graham's Law of Diffusion shows the rate of diffusion of gases is inversely proportional to their molar masses.

$$\frac{r_A}{r_B} = \sqrt{\frac{MM_B}{MM_A}}$$

where r is the rate (speed) and MM is the molar mass of the gas.

Exercise A molecule of oxygen gas has an average speed of 12.3 m/s at a given temp and pressure. What is the average speed of hydrogen molecules at the same conditions?

Chapter Problems

1. At atmospheric pressure in Denver, Colorado, is usually about 84.0 kPa. What is this pressure in atm and torr units?
2. What would happen to life on earth if ice were denser than liquid water? Explain your answer in detail and use examples.

3. Why is it important not to puncture an aerosol can? Explain WHY.
4. How-air balloons rise because the hot air inside the balloon is less dense than the cooler air outside. Calculate the volume an air sample will occupy inside a balloon at $43.0°$C if it occupies a 2.50 L at the outside air temperature of $22.0°$C, assuming the pressure is the same in both locations.
5. The lowest pressure achieved in laboratory is about 1.0×10^{-15} mm Hg. How many molecules of gas are present in a 1.00 L sample at that pressure and temperature of $22.0°$C?
6. Determine the pressure inside an old-fashioned television picture tube with a volume of 3.50 L that contains 2.00×10^{-5} g of nitrogen gas at $22.0°$C.
7. Propane, C_3H_8, is a gas commonly used as a home fuel for cooking and heating. Calculate the volume that 0.540 mol of propane occupies at STP. Think about the size of this volume compared to the amount of propane it contains. Why do you think propane is usually liquefied before it is transported?
8. At a given temperature, gaseous ammonia molecules (NH_3) have a velocity that is ______times ______than gaseous sulfur dioxide molecules (SO_2).
9. Write the balanced reaction for the complete combustion of octane, C_8H_{18}. If air is 20.9% oxygen by volume, how many liters of air are needed to complete the combustion of 25.0 L of octane vapor at STP? What volume of each product is produced?
10. A 3.25 gram sample of solid calcium carbide (CaC_2) reacts with water to produce acetylene gas (C_2H_2) and aqueous calcium hydroxide. If the acetylene was collected over water at $17°$C and 740.0 mmHg, how many milliliters of acetylene were produced? The vapor pressure of water at $17°$C is 14.5 torr.

10 Solutions

Where can you find the answers to solubility problems? In the solutions manual.

10.1 Solutions, Colloids, Suspensions, Electrolytes & Solubility

- Mixtures that are mostly liquid can be classified as one of three different things -
  - Solutions (homogeneous mixtures)
  - Same makeup throughout - evenly mixes on the molecular level
- Heterogeneous Mixtures
  - Different makeup in different parts of the mixture.
  - Includes colloids and suspensions.
- Simple tests can help determine what type of mixture you have.

Solutions are made up of two parts -

- Solute
- Solvent

The solute dissolves into the solvent.

Properties of solutions:

- Clear, but not necessarily colorless.
- Solute particles are ions or molecules wit ha size less than a nonmeter - very small.
- Particles cannot be seen even with a microscope, and the mixture doesn't scatter light.
- Cannot be separated by filtering, settling, or centrifuging.
- Solutions can be separated by evaporation.
- Can conduct electricity if an ionic compound is dissolved. Do not conduct if covalent/molecular compounds are dissolved.

Colloids:

- A heterogeneous mixture. Some can appear homogeneous with just your eyes, but the particles dispersed throughout the liquid can be seen with a microscope.
- The particles in a colloid are bigger than the particles dispersed in a solution.
- Can appear clear, slightly cloudy, or very cloudy.
- They scatter light. This is called the Tyndall Effect.
- Dispersed particles are about 10-100 times bigger than the particles dissolved in solutions.
- Will not conduct electricity.
- Will not separate into separate parts by settling, standing, or filtering.
- Can be separated by centrifuge or heating, depends on the specific colloid.

Suspensions:

- Large particles dispersed in liquid - can be seen with a light microscope and sometimes the naked eye.
- Cloudy when shaken, but the dispersed particles settle upon standing.

- Can speed separation by filtering or centrifuging.
- Will not conduct electricity.

Exercise A mixture doesn't leave residue on paper. Identify what this is.

The solvent is what is doing the dissolving. It is usually water because it is a non-ionic polar molecule that can dissolve anything else that is non-ionic or polar. Water can also dissolve most ionic compounds due to its hydrogen bonding.

The solute is what is dissolved in something else.

Electrolytes:

- Some solutions can conduct electricity. These are called electrolytes.
- Either ionic compounds or strong acids can act like electrolytes.
- Dissociated ions carry an electric current.

There are three categories of strong electrolytes: strong acids, strong bases, and soluble salts.

The strong acids are - HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_3$, and $HClO_4$.

Strong bases are hydroxides of group I and heavy group II metals (Car, Sr, Ba)

The soluble salts are compounds with ions of NO_3^-, Group I, NH_4^+, $C_2H_3O_2^-$, ClO_4^-, ClO_3^- with no exceptions. Cl^-, Br^- and I^- are soluble except with Pb^{+2}, Ag^+, and Hg_2^{+2}. SO_4^{-2} is soluble except with Pb^{+2}, Ag^+, Hg_2^{+2}, Ca^{+2}, Sr^{+2}, and Ba^{+2}.

Solubility Factors:

First we need to know if something will dissolve.

- Miscible - two substances that are miscible are soluble together and will mix.
- Immiscible - two insoluble liquids; they will not mix together.
- Polar things will dissolve in other polar things.
- Ionic substances will dissolve in water if the ionic bonds aren't too strong.
- Polar and Nonpolar substances will not dissolve together.

The process of dissolving: Solvation

1. Solvent is attracted to the solute.
2. Solvent particles surround the solute particles and pull them into solution.

Factors affecting rate of solution:

- Surface area - more contact between solute and solvent increases rate of solution.
- Agitation - mixing the mixture causes more contact between solute & solvent, increases rate of solution.
- Temperature - solvent particles are moving faster at higher temperatures, increases rate of solution.

Sometimes substances will not fully mix together no matter how much work is done. It all has to do with IMFs. Similar IMFs will dissolve together.

Solubility is the amount of solute that will dissolve in a given amount of solvent.

The rules for this are different for solids and gases.

Temperature

- Higher temperatures make gases less soluble in liquid
- Most solids are more soluble in liquid

Pressure

- Higher pressure above a solution will increase the solubility of a gas. This relationship is known as Henry's Law

- have no effect on solid solubility

Solubility is a physical property and is the amount of a substance that can be dissolved in a liquid.

On a solubility graph, if the amount is on the line, it is called saturated.

Saturation of a solution is sort of like saturation of a sponge. There are only so many holes in a sponge to hold a liquid, and when it is full there is simply no more room.

A saturated solution is when no more solute can dissolve.

Unsaturated solution is when more solute could dissolve. Any point below the line is unsaturated.

Supersaturated is when more solute is dissolved than normally could be. Any disturbance or seed crystal will cause the excess to precipitate out. Supersaturated means you are above the line and all solute is dissolved. It's more likely that you actually have a saturated solution with some undissoved solute present.

10.2 Units of Concentration

A concentrated solution has a relatively high amount of solute.

A dilute solution has a relatively low amount of solute.

What concentration units are most useful in chemistry?

- Molarity = moles solute/liters solution
- Molality - moles solute/kg solvent
- % Mass = grams part/grams total × 100
- Mole Fraction = moles part/moles total

Molarity

- The molarity of a solution is a measure of how many moles of the solute are present for each liter of solution.
- The liters of solution is not necessarily how much water was added. The final volume is usually more than the amount of pure water added since the solute adds to the volume.
- M = moles solute/Liters solution

Higher molarity values mean more concentrated.

Exercise How many grams of HCl would be necessary to create 2.00 L of a 4.0 M solution?

Molality is the concentration calculated by dividing the number of moles of solute by kilograms of solvent use to dissolve. This is useful for colligative properties - something we'll get to later.

m = mol solute/kg solvent

Exercise A solution contains 5.3 grams of carbon dioxide dissolved in 450. grams of water. What is the molality of the solution?

The molarity or molality of a solution does not tell you whether it is strong or weak. It tells you whether it is concentrated or dilute.

Mole fraction for one component of a solution is moles of component/total moles of all components. Mol fraction has no units.

Exercise You mix 30. grams of lithium chloride into 100.0 grams of water. What is the mole fraction of lithium chloride? % mass is grams of component/total mass of mixture × 100.

Exercise 60.0 grams of dextrose are added to 200. grams of water. What is the % mass of dextrose? What is the % of water?

You can't always find the solution concentration you need. It's common to have to dilute a more concentrated solution to create what you want. The dilution formula is

$$M_1V_1 = M_2V_2$$

where M is molarity and V is volume.

10.3 Colligative Properties

Colligative properties of solutions are properties that depend on the amount of solute particles.

Colligative properties only occur when a nonvolatile solute is added. Nonvolatile means it doesn't evaporate easily, so the solutes are usually solid.

4 important properties

- Freezing point of solution is lowered
- Boiling point of solution is raised
- Vapor pressure of solution is lowered
- Osmotic pressure is raised

Freezing point can be lowered by adding a nonvolatile solute to water, and the "depression", or specific number by which the freezing point is lowered, can be calculated.

The formula for this is

$$\Delta T = K_f m i$$

Where, ΔT is the change in freezing point, K is the molal freezing point constant, m is molality, and i is the number of particles from one "formula" of the solute.

Boiling point is increased by adding a nonvolatile solution, and the "elevation", or specific number by which the BP is elevated, can be calculated.

$$\Delta T = K_b m i$$

where K_b is the molal boiling point constant, and the other variables are defined above.

Vapor pressure lowering - the greater the number of solute particles in a solution, the lower the vapor pressure of the liquid solvent. The resulting vapor pressure is equal to the vapor pressure of the pure solvent times the mol fraction of the solvent.

$$P = P_{solvent} \cdot X_{solvent}$$

Osmotic pressure is the pressure required to stop osmosis from happening. Osmosis is the transfer from a dilute solution to a more concentrated one - osmosis is trying to balance out concentrations. Osmotic pressure increases with increased solute in a solution.

$$P_{os} V = nRT$$

When we previously mentioned that not all ionic compounds dissolve, there is a catch to this. They do dissolve a little, but most of the ionic compound stays in solid form, a bit will dissolve into aqueous ions. An "equilibrium" is reached between the solid and aqueous phases.

In order to determine how much actually dissociates, we have to be able to write what is called an "equilibrium expression".

Note:

- square brackets indicate concentration
- Ion charges are included inside the brackets
- Coefficients become exponents
- Solids and liquids have a concentration of 1

Exercise Write the K_{sp} expression for Lead(II) fluoride.

We can use the K_{sp} expression to determine how much of the substance will actually dissolve.

Math problems with equilibrium will almost always involve "RICE" tables -

- Reaction

- Initial Concentration
- Change in concentration
- Equilibrium concentration

Exercise What is the solubility of a saturated lead(II) chloride solution? K_{sp} for lead(II) chloride is 1.17×10^{-5}.

Chapter Problems

1. Why are gases less soluble at high temperatures?
2. What is a suspension and how is it different from a colloid? Explain the difference in definition and physical properties.
3. How many grams os solute are contained in a solution of 142 mL of 1.40 M potassium sulfate.
4. Calculate the mole fraction of the salt in a solution of 15.7 g of sodium chloride in 100.0 g of water.
5. How many milliliters of 0.400 M HBr solution can be made from 50.0 mL of 8.00 M HBr solution?
6. You have a container of powdered copper(II) sulfate ($CuSO_4$) and all standard lab equipment. For a lab, you need 1.00 L of 2.00 M solution. Create a numbered list describing the steps necessary to create the solution. Show all necessary calculations needed.
7. Explain on a particle basis why the vapor pressure of a solution is lower than a pure solvent.
8. A rock salt (NaCl), ice, and water mixture if used to cool milk and cream to make homemade ice cream. How many grams of rock salt must be added to water to lower the freezing point by $10.0°$C?
9. Write the correct formula for Lead(II) nitrate and determine if it is a strong electrolyte or not.
10. Write the correct formula for sodium permanganate and determine if it is a strong electrolyte or not.
11. Calculate the K_{sp} of the salt MgF_2, whose solubility is 2.7×10^{-3} mol/L.
12. Calculate the K_{sp} of $AgC_2H_3O_2$ given that the solubility is 1.02 g/100 mL.

11 Acids and Bases

The base is under assault! NaCl/NaOH

11.1 Acids & Bases

Acids and bases are two types of chemicals that have special properties.

They react with each other in a certain way - but they are not exact opposites of each other.

Acids:

- Acids are substance that dissociate in water to form hydrogen ions, which combine with water to form hydronium ions.
- The presence of these ions are what give a substance acidic properties.
- The H^+ can come directly from the acid molecule or from the acid's interaction with water. Either way, the chemical yields H^+.

Physical and Chemical Properties of Acids:

- Taste sour
- Produce hydrogen gas when they react with metals
- React with carbonates to form carbon dioxide and water.
- Turns litmus paper red.
- Electrolytes.

Strong acids dissociate nearly 100% and weak acids dissociate very little.

Polyprotic acids are acids that have more than one ionizable hydrogen. The hydrogens are dissociated one at a time.

Bases:

- Bases are substances that react in water to form hydroxide ions.
- The presence of hydroxide is what gives bases their basic properties.
- The OH^- can come from the basic molecule or from an interaction with water.

  Properties:

  - Taste bitter
  - Feel slippery
  - Turns litmus paper blue
  - Electrolytes in water solution
  - Can damage tissue if strong.

Acids and bases react with each other to produce neutral products.

Acid + Base $\rightarrow$ Water + Salt

Arrhenius Definition:

- Arrhenius acids contain hydrogen ions.
- Arrhenius bases contain hydroxide ions.

Bronsted-Lowry Acids & Bases

- Bronsted-Lowry acids are any substances that donate hydrogen ions (which is a proton, so called "proton donors".)
- Bronsted-Lowry bases are any substances that can accept hydrogen ions (from an acid or water - "proton acceptors")
- Lewis acids are electron pair acceptors.
- Lewis bases are electron pair donors.

Recall that when an acid dissolves in water, it donates a H^+ ion to a water molecule, forming hydronium.

The reverse reaction is also an acid/base reaction. In the reverse reaction, we call them conjugates.

We've seen water act like an acid and a bse.

This is called amphoteric. Several substances can act as acid or base, depending on what it is paired with.

The pH scale is used to indicate how acidic or basic a solution is.

- pH = 7 = neutral
- pH $<$ 7 = acidic
- pH $>$ 7 = basic

Because H^+ ions make something acidic, the more H^+ ions present, the more acidic the solution is.

The molar concentrations of H^+ are usually small, written in scientific notation, and hard to compare.

The pH scale, 0-14, is easier to interpret.

$pH = -\log[H^+]$

A change of one pH unit represents a tenfold change in H^+ concentration.

The pOH is a similar scale that mirrors the pH and $[H^+]$ relationship in the pH scale.

$pOH = -\log[OH^-]$

pH + pOH = 14 for any given solution.

If you know the pH or pOH, you can figure out the ion concentrations by working backwards.

$[H^+] = 10^{-pH}$ and $[OH^-] = 10^{-pOH}$

Substances that change color as pH changes are acid base indicators.

11.2 Titrations

The purpose of titrations are to determine the concentration of a solution.

To determine the concentration, if the unknown is the base, you titrate it with an acid of known concentration.

You titrate until the unknown is neutralized as indicated by a change in pH which will change the color of an indicator.

Moles H^+ will equal moles OH^-.

Neutralization is a reaction of an acid and a base to produce a salt and water.

- Water is produced from a union of a H^+ from the acid and an OH^- from the base.
- Salt: ionic compound formed from the positive ion of an aqueous base and the negative ion of an aqueous acid. The salt is not always NaCl.
- Equal amounts of H^+ and OH^- will neutralize completely.
- Salts produced from neutralization may or may not be neutral in solution.

The titration formula:

$$V_a M_a(\#\mathrm{H}^+) = V_b M_b(\#\mathrm{OH}^-)$$

Where V is volume, M is molarity, and $\#H^+$ or $\#OH^-$ is the # of hydrogen or hydroxide ions in the chemical formula of the acid or base.

Titration: lab procedure in which a carefully measured solution of known concentration is slowly added to a known volume of a second solution to experimentally determine its concentration.

Equivalence point: point of neutralization where $[H^+] = [OH^-]$

Endpoint: point where the color change of the indicator occurs.

Performing a titration:

- Buret: a measuring instrument with very small volume increments used to perform titrations.
- Before beginning a titration, rinse the buret with 5 mL of distilled water. Discard the water.
- Then rinse the buret with $\tilde{5}$ mL of the known concentration solution. Discard the rinse solution.
- THEN fill the buret with the known concentration solution.
- Record your initial volume and begin the titration.
- The buret showhs how much solution has been emptied.
- Be careful not to overshoot the end point.

11.3 Molar Mass through Titrations

Titrations are a method to determine many things:

- The molarity of an acid or base
- The molar mass of an acid or base
- The K_a or K_b value of an acid or base.

If you know bot hte mass and number of moles of any substance, you can determine its molar mass by dividing the grams by moles.

11.4 Acid-Base Equilibrium: Ka & Kb

The general reaction for a weak acid is

$$HA + H_2O \leftrightarrow H_3O^+ + A^-$$

Often we leave out the water part and just show:

$$HA \leftrightarrow H^+ + A^-$$

The general equilibrium constant expression for weak acids is shown below:

$$K_a = \frac{[H^+][A^-]}{[HA]} < 1$$

The general equation for weak base reactions is

$$B + H_2O \leftrightarrow HB^+ + OH^-$$

The general equilibrium constant expression for weak bases is shown below:

$$K_b = \frac{[HB^+][OH^-]}{[B]} < 1$$

Note that the concentrations that you substitute in are at equilibrium, meaning after the acid/base has dissociated to its maximum extent. Notice that water is not included in the K expression even if it is in the ionization/dissociation equation.

The K value can help us determine how much of the acid or base will ionize, called "x". We use RICE tables to organize our calculations and information.

- Reaction
- Initial Amount
- Change
- Equilibrium

There are two basic types of acid/base equilibria problems:

- You need to calculate K_a or K_b (you have x)
- You need to calculate x (you have K_a or K_b)

Exercise The weak acid hydrogen fluoride can be used to etch glass. In a 0.25 M solution of HF, the fluoride ion concentration is found to be 0.012 M. Find the K_a value for HF.

Chapter Problems

1. Calculate the acid, base, conjugate acid, and conjugate base for

$$HCl + NH_3 \rightarrow Cl + NH_4^+$$

2. Write a neutralization reaction for the reaction between sodium hydroxide and hydrobromic acid and predict whether the salt produced is acidic, basic, or neutral.
3. If the pH of a solution is 2.37, what is the concentration of hydrogen ions?
4. What is the pOH of a solution with $[OH^-] = 2.3 \times 10^{-6}$M?
5. Can you titrate a solution of unknown concentration with another solution of unknown concentration and get a meaningful answer? Explain your answer in a few sentences.
6. Explain the difference between the end point and the equivalence point of a titration.

12 Equilibrium

Equilibrium Constant. Que? The nature of the equilibrium state:

- Reactions are reversible.
- dynamic - $\rightleftarrows$ indicates that the reaction is proceeding in both the forward and reverse directions.
- Equilibrium does not mean nothing is happening.
- Once equilibrium is established, the rate of reaction in each direction is equal.
- This keeps the concentration of reactants and products constant.

$$\text{A} + \text{B} \rightarrow \text{C} + \text{D}$$

- The nature and properties of the equilibrium state are the same, no matter what the direction of approach.
- Initially, since there is only A and B present, they react very quickly.
- As A and B react, there are less and less of unreacted A and B molecules and so the rate of this reaction slows down.
- As the concentrations of C and D build up, they start to react to form A and B.
- And eventually the rate at which A and B react equals the rate at which C and D react. At this point equilibrium is established.

Note the concentrations of A, B, C, and D are not necessarily equal at equilibrium, but the concentrations are constant at equilibrium.

The equilibrium constant is used to determine equilibrium conditions. It is always temperature dependent.

For the general reaction

$$\text{aA} + \text{bB} \rightleftarrows \text{cC} + \text{dD}$$

The equilibrium constant expression is

$$\text{K} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The product concentrations appear in the numerator and the reactant concentration in the denominator. Each concentration is raised to the power of its stoichiometric coefficient in the balanced equation.

indicate concentration

- K_c is for concentration
- K_p is for partial pressure

In equilibrium constant expressions:

- Pure solids do not appear in the expression
- Pure liquids do not appear in the expression
- Water as a liquid or gas does not appear in the expression

If you know the concentration of things, just plug them into the K expression and solve.

Changing coefficients:

- When the stoichiometric coefficients of a balanced equation are multiplied by some factor, the K value is raised to the power of the multiplication factor.
- For the reverse reaction, take the reciprocal of K

- When adding reactions to produce another, multiply the respective K's to determine the K of the final reaction.

K_p only applies when all reactants and products are gases. Pressure must be units of atmospheres.

K_c and K_p are not interchangeable!

$$K_p = K_c(RT)^n$$

where n is the change in the number of moles of gas from the reactant to product side of the arrow.

Use a RICE table if you are not given all the concentrations.

External factors affecting equilibria:

Le Chatelier's Principle - If you disrupt a system in equilibrium, shifts in direction occur to reestablish equilibrium positions.

Temperature:

- Exothermic reactions - heat is a product
- Endothermic reactions - heat is a reactant
- The value of K is a constant only at constant T
- If you change T, the value of K changes

Changing reagent and product amounts:

- Adding reagent - shift tries to consume additional material by favoring forward reaction
- Removing a reagent - shift tries to replace material by favoring reverse reaction
- Adding product - shift tries to consume the product by favoring reverse reaction
- Removing product - shift tries to replace the product by favoring forward reaction
- Has no effect on the value of K

Pressure:

- An increase in pressure favors the side with the least # of gas moles
- A decrease in pressure favors the side with the most # of gas moles
- Has no effect on solids and liquids
- Has no effect on the value of K

Catalysts:

- No effect on the value of K
- But the reaction gets to equilibrium faster!

Weak acids only dissociate partly. There is an equilibrium constant associated with weak acid dissociation, K_a. It works just like K_c.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} < 1$$

The reaction quotient, Q.

For a general reaction

$$aA + bB \leftrightarrow cC + dD$$

the reaction quotient is

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Q has the appearance of K but the concentrations are not necessarily at equilibrium.

If $K > Q$, the system is not at equilibrium.

- Products are too small and the reactants are too big.
- The system will move towards equilibrium by making products and consuming reactants.

If K = Q, the system is at equilibrium.

If $K < Q$, the system is not at equilibrium.

- Reactants are too small and the products are too big.
- The system will move towards equilibrium by making reactants and consuming products.

Chapter Problems

1. K_{eq} for the reaction 2A + B $\rightleftarrows$ 2C is 8.0 Find the concentration of C when the concentration of A is 5.00×10^{-4} M and the concentration of B is 2.50×10^{-4} M.
2. For the reaction

   $$4NH_3\ (g) + 3O_2\ (g) \leftrightharpoons 2N_2\ (g) + 6H_2O\ (l)$$

   How will the concentration of each chemical be affected byd ecreasing the volume of the container?
3. Bromine and chlorine react to produce bromine monochloride according to the equation below. K_c = 36.0 under the conditions of the experiment.

   $$Br_2\ (g) + Cl_2\ (g) \leftrightarrow 2BrCl\ (g)$$

   If 0.180 moles of bromine gas and 0.180 moles of chlorine gas are introduced into a 3.0 Liter flask and allowed to come to equilibrium, what is the equilibrium concentration of the bromine monochloride? How many moles of bromine monochloride are produced?
4. When 2.0 mol of carbon disulfide and 4.0 mol of chlorine are placed in a 1.0 Liter flask, the following equilibrium system results. At equilibrium, the flask is found to contain 0.30 mol of carbon tetrachloride. What quantities of the other components are present in this equilibrium mixture? What is the equilibrium constant at this temperature?

   $$CS_2\ (g) + 3Cl_2\ (g) \leftrightarrows S_2Cl_2\ (g) + CCl_4\ (g)$$

5. Suppose you dissolved benzoic acid, C_6H_5COOH, in water to make a 0.15 M solution. K_a for benzoic acid $= 6.3 \times 10^{-5}$ at $25°$ C. Solve for the concentration of benzoic acid, the concentration of hydronium ion, the concentration of benzoate anion, and the pH.
6. Calculate the hydronium ion concentration and the pH of a 0.50 M solution of HNO_2. For HNO_2, K_a $= 4.5 \times 10^{-4}$.
7. At $60.2°$ C, the equilibrium constant for the reaction, N_2O_4 (g) $\leftrightarrow$ $2NO_2$ (g), is 8.75×10^2. At this temperature, a vessel contains dinitrogen tetroxide at a concentration of 1.72×10^{-2} M at equilibrium. What concentration of nitrogen dioxide does it contain?
8. If you wished to maximize the products in the following reaction, would you run the reaction at a high pressure or at a low pressure?

   $$CaCO_3\ (s) \leftrightharpoons CaO\ (s) + CO_2\ (g)$$

9. At $500°$ C, the equilibrium constant for the following reaction is 0.080. What is the value of Q if $[NH_3]$ = 0.059 M, $[N_2]$ = 0.59 M and $[H_2]$ = 0.42 M? How will this reaction proceed?

   $$N_2\ (g) + 3H_2\ (g) \leftrightarrows 2NH_3\ (g)$$

10. When 54.0 grams of ammonia is placed in an evacuated 3.50 liter rigid container and heated to 550. K, 12.75% of the gas has dissociated. What is the K_p value?
11. Write the equilibrium expression for the following reaction: Dinitrogen pentoxide gas decomposes into nitrogen dioxide gas and oxygen gas.
12. For the chemical equilibrium A + 2B $\leftrightharpoons$ 2C, the value of the equilibrium constant, K, is 10. What is the value of the equilibrium constant for the following reaction: 4C $\leftrightharpoons$ 2A + 4B?

13 Thermochemistry

If I were a Rapper, I'd be MC Delta T

13.1 Enthalpy, Enthalpy of Reactions, Spontaneity

Thermochemistry is the study of the energy occurring during chemical reactions and phase changes.

System vs. Surroundings - the part of the universe we are studying vs. everything else.

Endothermic is adding heat to the system and exothermic is removing heat from the system.

Enthalpy means the energy in a system at constant pressure. H is the symbol used for enthalpy. We can never fully measure the enthalpy of a system, but we can measure changes in enthalpy.

$$\Delta H = \text{change in enthalpy} = H_{products} - H_{reactants}$$

Exothermic

- System loses energy
- Energy listed as a product
- ΔH is negative

Endothermic

- System gains energy
- Energy listed as a reactant
- ΔH is positive

A reaction pathway shows the change in energy during a chemical reaction.

In an energy diagram, the energy needed is called the activation energy.

A thermochemical reaction includes a chemical reaction and the corresponding ΔH.

Combustion reactions are very common and release a lot of energy.

The energy change of combustion reactions is usually given as the "Standard Enthalpy of Combustion", or ΔH°_{comb}

The value is in the unit "kJ/mol".

- A spontaneous process is one that occurs without any outside intervention.
- A nonspontaneous process is one that occurs due to some intervention.

In general, exothermic reactions are spontaneous and endothermic reactions are nonspontaneous. There are some exceptions, and we will why later.

13.2 Hess's Law

Energy change in an overall chemical reaction is equal to the sum of the energy changes in the individual reactions comprising it.

The law is based on the principle of conservation of energy and the path independence of energy changes.

Hess's law can be used to predict energy changes that are not easily measured.

Hess's Law - energy change in an overall chemical reaction is equal to the sum of the energy changes in the individual reactions comprising it.

How to Solve Hess's Law Problems:

1. First decide how to rearrange equations so reactions and products are on appropriate sides of the arrows.
2. If equations had to be reversed, reverse the sign of ΔH.
3. If equations had be multipled to get a correct coefficient, multiply the ΔH by this coefficient.
4. Check to ensure that everything cancels out to give you the exact equation you want.
5. It is often helpful to being your work backwards from the answer you want!

13.3 Big Mama Equation

ΔH°_{f} is the enthalpy of formation.

It is the production of one mole of compound from its elements in their standard states.

It is zero for elements in their standard states.

Hess's law can be summarized into a similar equation:

$$\Delta H^{\circ}_{\text{rxn}} = \sum n\Delta H^{\circ}_{f(products)} - \sum n\Delta H^{\circ}_{f(reactants)}$$

You must still multiply ΔH by the coefficient, n. The reactants are subtracted because they are equations that had to be reversed using Hess's Law.

Notes:

- Sometimes all values are not found in the table of thermodynamic data. For most substances it is impossible to go into a lab and directly synthesize a compound from its free elements.
- The heat of formation for these substances must be found by working backwards from its heat of combustion.

13.4 Reaction Spontaneity, Energy & Heat Transfer

Entropy, S, of a system is the randomness or disorder of that system. The universe wants more entropy.

Reaction which increase entropy are more likely to occur than ones in which the entropy decreases.

Reactions that decrease entropy are less likely to occur than ones in which the entropy increases.

Entropy is measured in J/K.

- Entropy changes associated with changes in state can be predicted.
- The dissolving of a gas in a solvent always results in a decrease in entropy.
- The dissolving of a solid (or liquid) in a solvent always results in a increase in entropy.
- The entropy of a system increases when the number of gaseous product moles is greater than the number of gaseous reactant moles.
- An increase in the temperature of a substance is always accompanied by an increase in the random motion of its particles.
  - Recall that the kinetic energy of molecules increases with temperature.
  - Increased kinetic energy means faster movements, meaning more disorder.

We can calculate entropy changes using the Big Mama Equation!

We've said that these changes are favorable/spontaneous:

- Exothermic
- Increases in entropy

Free Energy

- Gibbs free energy, G, is energy that is available to do work. It is a combined function of entropy and enthalpy.

$$\Delta G = \Delta H - T\Delta S$$

- When ΔG for a reaction is negative, the reaction is spontaneous
- When ΔG for a reaction is positive, the reaction is non-spontaneous

Another way to say a reaction is "spontaneous" is by saying it is "thermodynamically favorable". A non-spontaneous reaction would be "thermodynamically unfavorable". Big Mama can also be used for Gibbs free energy!

Types of Energy

- Potential - stored energy
- Kinetic - energy of motion
- Chemical - form of potential energy related to the structural arrangement of atoms or molecules
- Thermal - form of kinetic energy resulting from the motion of particles and is transferred as heat

The Law of Conservation of Energy states that the total amount of energy in a closed system remains constant. Energy cannot be created nor destroyed. In a closed system, energy can only change form.

Heat:

- Energy flows from something warm to something cool
- A hotter substance gives energy to a cooler one
- When heat is transferred, there is a change in the energy within the substance
- There are three types of heat transfer - conduction, convection, and radiation

Conduction:

- The direct transfer of heat through contact
- Occurs in solids, liquids, and gases. When 2 objects at different T's are in contact with each other, KE is exchanged.

Convection:

- Transfer of heat by the motion of fluids
- Occurs in liquids and gases

Radiation:

- Radiation is heat that is transferred through electromagnetic waves.
- Radiation is the only form of heat transfer that does not require matter, and can move through space.

13.5 Specific Heat

Different substances have different capacities for storing energy.

Specific heat is the amount of heat needed to raise the temperature of 1 g of a substance by 1° C.

- Conductors are materials that transfer heat easily and quickly. Metals are the best conductors of heat because they have a low specific heat.
- Insulators do not conduct heat easily. They have a high specific heat.

Water has a uniquely high specific heat compared to other substances.

There are moderate climates around lakes & oceans.

- The body of water absorbs heat from air on hot days and release it back into the air at night or on cold days.

Water is sprayed on citrus fruit to protect it from freezing. As the water freezes, it releases heat, which warms the fruit.

Specific heat can be calculated.

$$q = mc\Delta T$$

where q is the heat energy, m is the mass of the substance, c is the specific heat capacity, and ΔT is the change in temperature.

Specific heat, c, is a physical property. The higher a substance's specific heat, the more heat it takes to raise its temperature, so the longer it takes to heat it up or cool it off.

Energy is the capacity of doing work or supplying heat.

Heat is a form of energy that transfers from one object to another because of a temperature difference between them. Heat always flows from a warmer object to a cooler object.

Energy units: calorie [cal] or joule [J]

A calorie is a quantity of heat needed to raise the temperature of 1 g of water 1°C.

- System - the part of the universe on which you focus your attention.
- Surroundings - everything else in the universe

The Law of Conservation of Energy states that in any chemical or physical process, energy is neither created nor destroyed.

To measure energy changes

- We use a device called a calorimeter
- The reaction in the cup either absorbs or releases energy causing the water temperature to rise or fall
- The temperature change of the surrounding water tells you whether energy is absorbed or released.

Chapter Problems

1. A 4.50 g nugget of pure gold absorbed 276 J of heat. What was the final temperature of the gold if the initial temperature was $25.0°$ C?
2. Draw a potential energy diagram to represent the following reaction. A set of reactants begin with 120 kJ of energy. The activation energy is 60 kJ. The ΔH of the reaction is 30 kJ.
3. Pure liquid acetic acid is made from the reaction between methanol and carbon monoxide, as seen below. If you produce 1.00 L of acetic acid (which has a mass of 1044 grams), how much energy is evolved?

$$CH_3OH\ (l) + CO\ (g) \rightarrow CH_3COOH\ (l) \qquad \Delta H = -355.9\ kJ$$

4. Calculate ΔH for the reaction: $4NH_3\ (g) + 5O_2\ (g) \rightarrow 4NO\ (g) + 6H_2O\ (g)$, from the following data.

$$N_2\ (g) + O_2\ (g) \rightarrow 2NO\ (g) \qquad \Delta H = -180.5\ kJ$$
$$N_2\ (g) + 3H_2\ (g) \rightarrow 2NH_3\ (g) \qquad \Delta H = -91.8\ kJ$$
$$2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (g) \qquad \Delta H = -483.6\ kJ$$

5. Use a standard enthalpies of formation table to determine the change of enthalpy for the following reaction

$$2NO\ (g) + O_2\ (g) \rightarrow 2NO_2\ (g)$$

6. Use the table of standard free energies for the following reaction

$$2SO_2\ (g) + O_2\ (g) \rightarrow 2SO_3\ (g)$$

7. Solve for ΔG of the reaction

$$2H_2O_2\ (l) \rightarrow 2H_2O\ (l) + O_2\ (g)$$

Is the reaction spontaneous or not? Assume standard temperature.

8. Does the following reaction shown an increase or decrease in entropy?

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

9. The conditions in which ΔG is always negative when ΔH is ______and ΔS is ______.

14 Nuclear Chemistry

Radioactive Cat. Has 18 Half-Lives.

Types of Reactions:

Chemical Reactions:

- involve valence electrons
- elements are rearranged to form new compounds
- Law of Conservation of Mass obeyed
- relatively little energy can be absorbed or released

Nuclear Reactions:

- changes occur in the nucleus; involve protons, neutrons, and electrons
- Transmutation occurs (one element turns into another)
- Law of Conservation of Mass is not always obeyed; mass can be changed into energy
- lots of energy can be produced

Roentgen discovered x-rays in 1895 as a result of exposing photographic plates to the rays emitted from a cathode ray tube. He found that x-rays could travel through some substances.

Becquerel in around 1900 studied phosphorescence (glowing) in uranium ore when he discovered that the ore also exposed photographic plates.

The Curies further experimented with uranium ore. They isolated uranium as the substance with the unique properties. They were the first to call uranium "radioactive" - meaning a substance that gave off rays of energy from the nucleus.

Radioisotopes are isotopes that have unstable nuclei; they emit radiation to become more stable.

The radiation emitted can be one of several types. The three most common types of radiation are:

- alpha, α
- beta, β
- gamma, γ

Alpha decay occurs when an atom loses an alpha particle (2 protons + 2 neutrons), aka a helium nucleus. The loss of two protons decreases the atomic number by 2 and decreases the mass number by 4. Since the number of protons changes, a new element is formed. Alpha radiation isn't powerful (compared to other types of radioactive decay) and can be stopped by paper alone.

Beta decay occurs when an atom loses an electron from the nucleus. It is the result of a neutron converting to a proton and electron and the high energy electron is ejected. The new proton increases the atomic number by 1, which results in a new element. Note that the mass number is unchanged.

Gamma radiation occurs when high energy photons (electromagnetic energy) are ejected from the nucleus. Usually accompanies other nuclear emissions and represents the energy lost when the nucleus becomes more stable. Does not change the atomic number or mass and it is difficult to stop; it requires several cm of lead or several feet of concrete to block.

Relative Strengths of Radiation:

- Alpha particles can be stopped by paper
- Beta particles can be stopped by a few mm of aluminum

- Gamma particles can be stopped by a few cm of lead or several feet of concrete

Nuclear Mass Defect - The difference in the calculated mass of an atom (based on C-12) and the measured mass. Measured mass if always smaller than expected. Lost mass has been converted to energy.

Nuclear Binding Energy - the energy that comes from the lost mass. Holds the nucleus together. Lots of energy produced, according to Einstein's famous equation, $E = mc^2$

Strong Nuclear Force - a force that keeps the protons from repelling each other in the nucleus. Acts over a short distance and overcomes electrostatic repulsion (like charges repel). Strong nuclear force struggles in very large elements.

Stability of Isotopes:

- Nuclei that are not stable tend to undergo radioactive decay, a process that emits radiation and particles from the nucleus
- Spontaneous decay is the result of instability in the nucleus of an atom.

What determines nuclear stability?

- Neutron to proton ratio. If this ratio is close to 1:1 in a low mass (small) element, the isotope is likely to be stable.
- Neutron to proton ratio. For larger elements, the stable ratio is 1.5: 1-1.5 neutrons to protons

In a graph of protons vs neutrons, the area of the graph within which all stable nuclei are found is called the band of stability.

- Nuclei with a high neutron to proton ratio will undergo beta decay.
- Nuclei with low ratio will undergo positron emission or electron capture.
- Nuclei with high atomic masses tend to undergo alpha decay.

Electron Capture:

- The nucleus captures an electron from the emission cloud
- The electron combines with a proton to convert into a neutron
- The mass number remains the same and the atomic number decreases since a proton was converted

Positron Emission:

- The radioactive decay process in which a positron is emitted from the nucleus.
- A positron is a particle with the same mass as an electron but opposite in charge.
- The number of protons decreases bu the mass number remains the same.

Artifical Nuclear Charges Neutron Bombardment

- A neutron enters the nucleus at a high speed and is captured; usually creates a radioactive isotope
- Creates synthetic isotopes useful in medicine and research
- Increases the mass number by one, but the atomic number remains the same.

A radioactive decay series is a series of nuclear reactions that begins with an unstable nucleus and ends with the formation of a stable nucleus.

- Nuclei that are not stable tend to undergo radioactive decay, a process that emits radiation (gamma radiation) and particles from the nucleus
- Spontaneous decay is the result of instability in the nucleus of an atom

Most elements have some radioactive isotopes (radioisotopes) that occur naturally, but the percentage of radioactive isotopes is so low that we do not consider that element radioactive.

Other elements, such as radon, have a high percentage of radioisotopes, so we consider that element, in general, to be radioactive.

Radioactive isotopes can be created from stable ones by neutron bombardment, which causes instability in the nucleus by disturbing the 1:1 neutron:proton ratio.

All elements heavier than bismuth are naturally radioactive due to their instability. They tend to spontaneously undergo alpha decay to bring them closer to a stable nucleus.

All elements heavier than uranium, atomic number 92, are artifically created and are always radioactive. These are the transuranium elements. They are usually formed by bombarding a heavy atom with neutrons or alpha particles. Scientists have created elements up to about 118 protons.

Half life is the amount of time it takes for one half of a radioactive element to change from the parent element to the daughter element (from radioactive element to product element).

Use for formulas

$$\mathrm{m}_t = \mathrm{m}_0(1/2)^n$$
$$\mathrm{n} = \frac{t}{T}$$

where m_t is the final mass, m_0 is the original mass, n is the number of half lives, t is the elapsed time and T is the time for one half life.

You can also use the following formula

$$\ln(\mathrm{m}_t) = -(0.693/\mathrm{T})\mathrm{t} + \ln(\mathrm{m}_0)$$

- Fossils are dated by knowing the ratio of parent element to daughter element
- Two things must be known - length of half life and ratio of parent:daughter
- The carbon-14 isotope is used for wood, bones, tissue, anything that used to be alive
- All living things contain carbon, and part of the carbon is C-14
- After something dies, it stops exchanging carbon with the atmosphere and the C-14 begins to decay
- Because we know the half-life of C-14 to be 5730 years, we can measure the amount of C-14 left and calculate how long the organism has been dead.

Nuclear Fission:

- Fission occurs when a neutron hits a large, unstable nucleus and causes it to break into two smaller nuclei.
- Extra neutrons are emitted from this reaction, and those neutrons go on to hit other nearby atoms, causing them to go through fission.
- A chain reaction occurs if there are enough atoms present.

Critical mass is the amount of a substance that must be present in order for a chain reaction to occur.

If not enough mass is present, the neutrons will escape the sample without hitting other atoms, and there will be no chain reaction. This is called subcritical mass.

In addition to neutrons being released, larget amounts of energy are also released. Nuclear fission is highly exothermic. In nuclear power plants, the energy and chain reaction is controlled and slow. In an atom bomb, the fission is uncontrolled and very fast.

Parts of a Nuclear Reactor in a Power Plant

- Fuel rods - usually uranium-235 or plutonium, a fissionable isotope
- Control rods - surround the fuel rods and absorb some of the neutrons to control the chain rxn
- Graphite Moderator - slows the neutrons so they can enter the fuel atoms at the most effective rate
- Water heated to steam - same as fossil fuel plant. The exothermic reaction heats the water to steam, which turns a turbine
- Concrete Containment Building - dome shaped building around the nuclear reactor to keep radioactivity inside

Nuclear Fusion

- Fusion is the process by which small nuclei fuse together to form a heavier nucleus
- This is the main process that goes on in our sun and other stars. It is responsible for the light and energy given of - highly exothermic for small atoms
- The primary reaction is two hydrogen isotopes building helium

Fusion produces lots of energy, and its products are generally not radioactive. It is appealing as an energy source. However, to get fusion started, lots of energy is needed to get two repelling hydrogen nuclei to move together. A thermonuclear reactor is required. Fusion isn't a feasible energy source for now.

Uses of Radioactivity

- Radiation therapy for cancer - radiation targets cancerous cells
- Radiotracers for medical diagnosis and research
- Nuclear Power
- Food preservation - it doesn't make your food radioactive, it kills bacteria and fungus
- Smoke Detectors

Properties of Radioactive Elements

- Radioactive elements will expose photographic film
- Ionize other atoms, causing them to become charged
- Can cause some other substances to fluoresce, or glow in the dark
- Prolonged exposure can cause damage to cells

2 types of radiation damage:

Somatic

- affects nonreproductive body tissue
- Only affects organism during its lifetime
- Causes burns, cancer

Genetic

- Damages sex cells
- Affects chromosomes and therefore offspring
- Causes mutations, lifelong disease, and deformities

Chapter Problems

1. Explain the relationship between an atom's neutron-to-proton ratio and its stability.
2. Write a balanced nuclear equation for the alpha decay of americium-241.
3. Technetium-104 has a half-life of 18.0 minutes. How much of a 165.0 sample remains after 90.0 minutes?
4. A sample initially contains 150.0 mg of radon-222. After 11.4 days, the sample contains 18.7 mg of radon-222. Calculate the half-life.
5. Discuss how the amount of fissionable material present affects the likelihood of a chain reaction.
6. Describe the penetration power of alpha, beta, and gamma radiation.

15 Organic Chemistry

What do you call an acid with an attitude? A-mean-oh acid.

Organic chemistry: the chemistry of carbon compounds

- Contain carbon
- Have covalent bonds - low melting points, low boiling points, are soluble in nonpolar solvents, burn in air
- Can form large molecules (polymers)

Hydrocarbons

- Compounds that contain only carbon and hydrogen

Complete structural formulas show all the atoms with the bonds between each of the atoms represented by lines.

Condensed structural formulas are a shorthand way and list all atoms in order and tells how they are bound together.

Molecular formula - atoms w/ subscripts making up molecule.

Skeleton structure/formula - leaves out H atoms, only C skeleton and connecting bonds shown and bonds are represented as lines

Bond-line Drawing - Only C-C bonds are shown. Each vertex represents a carbon atom. It is understood that hydrogen atoms are attached as needed to complete the bonding.

Alkanes

- Alkanes are hydrocarbon chains that have only single bonds between the carbons

Branched chain alkanes

- Sometimes carbons are not bonded in a straight chain. Sometimes there are branches.
- The longest branch is called the parent chain. The side branches are called the substituent groups.

Name of Alkyl groups

- The group names have the same prefix as their corresponding parent chains, but the -ane suffix is replaced with -yl

Rules for Naming Organic Structures

- Count the carbons in the longest chain. This parent chain provides the base name of the structure.
- Number each carbon in the parent chain, starting with the end closest to a substituent group.
- Name each alykl group substituent before the name of the parent chain. Include the alykl name the number of the carbon it is attached to in the parent chain.
- If the alykl group occurs more than once, use a prefix before its name to indicate how many times it appears.
- Whenever different alykl groups are attached to the same parent chain, name them in alphabetical order
- Write the entire name using hyphens to separate numbers from words and commas to separate numbers. No spaces.

Cyclic Alkanes

- Cyclic hydrocarbons are organic compounds that exist as carbon rings.

- The prefix cyclo- is used
- Alkyl groups added to front of name of ring

Naming branched cycloalkanes

- The ring is considered the parent chain
- Number the carbons beginning with the one attached to a substituent group that gives the lowest possible sum of numbers in the name

Multiple Carbon-Carbon Bonds

- Alkanes all have single bonds between carbons. These are called saturated hydrocarbons.
- Some hydrocarbons contain double or triple bonds. These are referred to as unsaturated hydrocarbons.

Alkenes

- Alkenes contain at least one double bond between carbon atoms.

Naming alkenes

- Alkenes are named like alkanes, but their names have -ene at the end.
- The double bond will always be part of the parent chain.
- You also must specify the location of the double bond with a number. This is the number of the carbon atom where the double bond starts.

Naming cycloalkenes

- Named like cycloalkanes, but carbons #1 & 2 must be attached to the double bond.

Naming branched alkenes

- The longest carbon chain must contain the double bond. Always start numbering closest to the double bond.
- If there is more than one double bond, use a prefix before the -ene to indicate how many.

Alkene Geometric Isomers

- All atoms are bonded in the same order but are arranged differently in space
- We'll discuss isomers further later
- cis- the functional groups are on the same side of the molecule
- trans- the functional groups are on opposite sides of the molecule

Alkynes

- Alkynes are hydrocarbons that contain at least one triple bond. They are named in the same way as alkenes, but with suffix -yne.
- Can have 1 triple bond
- Parent c hain must contain both C atoms of triple bond

Remember: each C forms a covalent bond with 4 other atoms.

Classes of hydrocarbons:

- Aliphatic - does not contain benzene ring
- Aromatic - contains benzene ring
- Aliphatic -
  - Alkanes - simplest class of organic compounds contain only C and H and have only single bonds
  - Alkenes have a C-C double bond
  - Alkynes have a C-C triple bond

Saturation:

- Saturated: hydrocarbon has maximum #H'2
- Unsaturated: hydrocarbon has less than maximum #H's (can be cyclic or have double or triple bonds)

Substituents: atoms that take the place of H

Functional Groups

- Small structural units within molecules where most chemical reactions occur
- R (radical) represents any hydrocarbon attached to functional groups
- Since double & triple bonds are chemically reactive, they are considered functional groups

Benzene

- Compounds containing benzene rings are called aromatic compounds.
- There are not 3 single bonds and 3 double bonds in a benzene ring. Instead, the electron pairs are delocalized, which means they are shared among all 6 C's in the ring.
- If there are other groups present on the benzene ring the compound is said to be a substituted benzene.
- When benzene is a substituent on a carbon chain, it is called phenyl.

Alcohols

- Carbon compounds containing a hydroxyl group, -OH
- -ol added on to name of compound
- Classified by how many substituents are attached to the C of -OH group

Halocarbons

- Carbon compounds containing halogens (Cl, F, Br, I)
- Alkyl halide - halogen attached to carbon of aliphatic chain
- Aryl halide - halogen attached to aromatic hydrocarbon
- Occur by substitution reaction (hydrogen replaced by halogen)

Ethers

- 2 C's single bonded to oxygen atom
- General formula R_1-O-R_2
- Simple ethers can be named by namimg alykl groups alphabetically followed by word "ether".
- Another way is name smaller hydrocarbon prefix, add -oxy, and join it to alkane name of the larger hydrocarbon group.

Ketones

- In middle of compound
- Change final -e of alkane to -one.
- Indicate number before name to indicate position of ketone group
- Less reactive than aldehydes, so popular as solvents

Aldehydes

- at end of compound
- Change final -e of alkane to -al
- Many have characteristic odors/flavors

Amines

- Contain N-C in aliphatic chains or aromatic rings

- Primary amine general formula R-NH_2
- Called amino group (found in amino acids)
- Suffix -amine
- Amines are stinky!

Amides

- -OH of carboxylic acid is replaced by N bonded to other atoms
- Write name with same number of carbon atoms, replacing final -e with -amide

Carboxylic Acids

- Organic acid with carboxyl group (-CO_2H or -COOH)
- Change -ane of alkane to -anoic acid
- Weak acids (ionize slightly to give carboxylate ion and proton)

Esters

- Carboxyl group where H of hydroxyl group replaced by alkyl group (-OR)
- General formula $R_1CO_2R_2$
- Formed by reaction of an alcohol and a carboxylic acid
- Substituent suffix -oate

Aromatic Hydrocarbons

- Substituents attached to benzene ring are usually named as derivatives of benzene
- Benzene is the parent compound if there is no continuous HC longer than 6 C's joined with it

Disubstituted Aromatic Hydrocarbons

- o-/m-/p- not used with 2 different substituents
- Name substituents in alphabetical order

Isomers

- Compounds with same molecular formula but different structures
- Structural isomers - same molecular formula but different connections between atoms
- Funtional isomerism - substances have same molecular formula but different functional groups
- Positional isomerism - occurs when same functional groups are in different positions on same C chain
- Stereoisomers - same molecular formula, same connections between atoms, but different arrangements of atoms in 3-D space
- Geometric isomers - result from different arrangements of groups around double bond

Properties and uses of Alkanes

- Nonpolar
- Low boiling point/melting point
- Immiscible in water but soluble in nonpolar solvents
- Low reactivity due to nonpolarity
- Relatively strong C-C and C-H bonds limit their use

Properties and Uses of Alkenes

- Nonpolar/low solubility in water
- Low melting/boiling points

- More reactive than alkanes
- Several occur naturally in living organisms
- Ethene used in plastics
- Scents of lemons, limes and pine trees

Properties and Uses of Alkynes

- Useful starting material in many synthesis reactions due to reactivity
- Physical and chemical properties similar to alkenes.

Chapter Problems

1. Draw a structural formula for 2,2-dimethylpropane
2. Draw the condensed structural formula for 2,2-dimethyl-3-hexyne
3. Draw a structure for 1,2,-difluoro-3-iodocyclohexane
4. Draw a structure for hexyl methanoate
5. Draw the organic molecule butyl pentyl ether
6. Draw the molecule 1,3-cyclopentadiene
7. Draw the compound 2-bromophenol
8. Which types of carbon-carbon bonds can rotate and which ones cannot rotate?