Honors Chemistry Notes

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1 Nature of Science

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.

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

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

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.

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

$$\text{\& error} = \frac{\text{—accepted-experimental—}}{\text{accepted}} \times 100$$

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? (0.8%)

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? (2%)

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)
2. 0.98 (2)
3. 980 (2)
4. 0.0098 (2)
5. 0.0098000 (5)
6. 98098 (5)

7. 980. (3)
8. 980.0 (4)

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. 400,003 (4.00003 $\times 10^5$)
2. 0.00007 (7 $\times 10^{-5}$)
3. 19000 (1.9 $\times 10^4$)
4. 422000 (4.22 $\times 10^5$)
5. 422000. (4.22000 $\times 10^5$)

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

1. 3.1 $\times 10^4$ (31000)
2. 1.0 $\times 10^{-4}$ (0.00010)
3. 1.00 $\times 10^2$ (100.)
4. 9.9 $\times 10^{-5}$ (0.000099)

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. 2.8 mm $\times$ 4.5039 mm (13 mm^2)
2. 2.097 grams - 0.12 grams (1.98 g)
3. (4.565 $\times$ 2.3)/8.009 (1.3)
4. 2.64 $\times$ 1000. cm + 3.27 $\times$ 100. cm (2970 cm)
5. (15.30 g) $\div$ (6.4 mL) (2.4 g/mL)
6. 18.9 g - 0.84 g (18.1 g)

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

$$\text{Density} = \frac{m}{V}.$$

This can be arranged to solve for mass or 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 - What is the density of a piece of wood that has a mass of 35.99 g and a volume of 45.68 cm^3? (0.7879 g/cm^3)

Exercise - A metal cylinder is placed into a graduated cylinder with 24.0 mL of water. After the cylinder is added, the volume of water rises to 30.4 mL. The density of the cylinder is known to be 8.9 g/mL. What is the mass of the cylinder? (57 g)

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 (23900 m)
2. Convert 4.7 L to mL (4700 mL)
3. Convert 34.98 g to kg (0.03498 kg)
4. Convert 22.8 cm to m (0.228 m)
5. If 1 inch = 2.54 cm, convert 3.00 cm to inches (1.18 in)
6. If 1 gallon = 4.1 L, convert 2.5 gal to L (10. L)

7. If 1 kg = 2.2 pounds, convert 48 pounds to kg (22 kg)

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? (9144 cm)
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? (23900 oz)

2 Atomic Structure and Energy of Electrons

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? (chromium)

- 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.

Answer: atomic number - 44, mass number - 101, protons - 44, neutrons - 57, electrons - 44, hyphen method - Ru-101

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. (40. amu)

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. (Ga-71: 40%, Ga-69: 60%)

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.

Exercise - 1 mole of potassium has a mass of _____ g. (39.10)

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.

Exercise - What is the molar mass of fluorine? (19.00 grams)

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 - How many atoms are in 55.4 grams of lithium? (4.81$\times 10^{24}$ atoms)

Exercise - What is the mass in grams of 3.011×10^{23} atoms of iron? (27.93 g Fe)

Exercise - How many atoms are in 8.43 moles of nickel? (5.07×10^{24} atoms Ni)

Exercise - How many atoms are in 1.00×10^{-10} grams of gold? (3.06×10^{11} atom Au)

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 - Halogens have how many valence electrons? (7)

Exercise - Copper is part of which block? (d)

Exercise - Which group contains the alkaline earth metals? (2)

Exercise - Which block do the lanthanide and actinide series belong to? (f)

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.

Exercise - Write the electron configuration for the nitrogen atom ($1s^2 2s^2 2p^3$)

Exercise - Which noble gas does the nitrogen ion mimic? (neon)

- 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.

Exercise - What is the electron configuration for a copper atom?

- cuprous, Cu^+ - $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$
- cupric, Cu^{2+} - $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9$

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 wavelength of a wave with a frequency of 7600 Hz? (39000 m)

Important conversions:

- 1 m = 1×10^9 nm
- 1 MHz = 1×10^6 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. (1.1 Hz)

Exercise - Calculate the wavelength of a photon with 5.32×10^{-33} J of energy. (3.74×10^{7} m)

3 Periodicity

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? (s)

Exercise - Uranium is in which block? (f)

Exercise - Silicon is in which block? (p)

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? (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 (Be)

Exericse - Which has the highest electronegativity? As, Sn, or S (S)

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

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

4 Bonding and Compounds

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? (Covalent)

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)

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 - Zinc nitride (Zn_3N_2)

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

Exercise - Tin(IV) Sulfate ($Sn(SO_4)_2$)

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

Exercise - $Zn(OH)_2$ (zinc hydroxide)

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 - $NiPO_4$ (nickel(III) phosphate)

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 - Tetraphosphorus decasulfide (P_4S_{10})

Exercise - PCl_5 (phosphorus pentachloride)

- 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 - HCN (hydrocyanic acid)

Exercise - Dichromic Acid ($H_2Cr_2O_7$)

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 (18.016 g)

Exercise - Calculate the molar mass of calcium chloride. (110.98 g)

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 $Ca(NO_3)_2$ (164.10 g)

Exercise - How many atoms of each element are in aluminum carbonate. (Al: 2, C: 3, O: 9)

Exercise - Calculate the molar mass of zinc nitrate. (189.41 g)

Exercise - How many moles are equal to 5.06×10^{23} molecules of Br_2? (0.840 mol Br_2)

Exercise - How many moles are equal to 3.905×10^{23} formula units of calcium hydroxide? (0.6485 mol $Ca(OH)_2$)

Exercise - What is the mass of 0.7880 moles of calcium cyanide? (72.59 g $Ca(CN)_2$)

Exercise - How many nitric acid molecules are in 4.20 g of HNO_3? (4.01×10^{22} mole 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? (51.43%)

Round percent composition answers to 4 sig figs.

Exercise - Calculate the % composition of nitrogen in ammonium nitrate? (35.00 %)

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.

Exercise - Write the formula for magnesium sulfate heptahydrate. ($MgSO_4 \cdot 7H_2O$)

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. (36.09 %)

Exercise - Calculate the % water in magnesium sulfate heptahydrate. (51.17 %)

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? (N_2O_5)

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? (C_6H_6)

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 - I_2 (Iodine: 0)

Exercise - MnO_4^- (Manganese: $+7$, Oxygen: -2)

5 Reactions

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. (solid calcium hydroxide yields solid calcium oxide and liquid water)

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 (2 in front of HCl)

Exercise - Balance $Zn(OH)_2$ + H_3PO_4 $\rightarrow$ $Zn_3(PO_4)_2$ + H_2O (3 in front of $Zn(OH)_2$, 2 in front of H_3PO_4, 6 in front of H_2O)

Exercise - Write the balanced equation for lithium chlorate decomposing into lithium chloride and oxygen gas. ($2LiClO_3 \rightarrow 2LiCl + 3O_2$)

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 - Carbon burns ($C + O_2 \rightarrow CO_2$)

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 - Sodium carbonate decomposes when heated ($Na_2CO_3 \rightarrow Na_2O + CO_2$)

Exercise - Calcium chlorate decomposes when heated ($Ca(ClO_3)_2 \rightarrow CaCl_2 + 3O_2$)

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 - Fluorine + Potassium Bromide ($F_2 + 2KBr \rightarrow 2KF + Br_2$)

Exercise - Copper + Sulfuric Acid ($Cu + H_2SO_4 \rightarrow$ No rxn)

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 - sodium bicarbonate + hydrochloric acid ($NaHCO_3 + HCl \rightarrow NaCl$ (aq) + CO_2 (g) + H_2O (l))

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 $2Na + Cl_2 \rightarrow 2NaCl$ (oxidized: Na, reduced: Cl, reducing agent: Na, oxidizing agent: Cl_2)

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. (Pb^{2+}(aq)+$2Cl^{-}$(aq)$\rightarrow$$PbCl_2$(s))

6 Stoichiometry

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? (6.6 mol CO_2)

Exercise - How many grams of magnesium chloride are produced when 0.500 moles of magnesium reacts with an excess of hydrochloric acid? (47.6 g $MgCl_2$)

Exercise - How many moles of zinc sulfate are produced when 4.55 g of zinc reacts with an excess of sulfuric acid? (0.0696 mol $ZnSO_4$)

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. (1.1 g CO_2)

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? (85.4%)

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? (Carbon, 20.0 g)

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? (0.0090 L $CuSO_4$)

Exercise - How many atoms of oxygen are there in a 3.0 mole sample of $Mg(ClO_3)_2$? (1.1×10^{25} atoms)

7 VSEPR/IMFs

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? (covalent)

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? (5)

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 (nonpolar covalent)

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. (polar)

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? (LDFs)

8 States of 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. (472 kJ)

9 Gas Laws

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: (253 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 (1.05 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? (2.1 atm)

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? (171 K)

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. (752 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? (2.1 atm)

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? (8090 K)

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? (29 g/mol)

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.

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

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. (0.0037 mol O_2)

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.

Exericse - 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? (49.0 m/s)

10 Solutions

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. (solution, colloid)

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? (290 g)

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? (0.27 m)

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? (0.11)

% 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? (23.1% and 76.9%)

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. ($[Pb]^{+2}[F^-]^2$)

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}. $(1.43 \times 10^{-2}$ M)

11 Acids and Bases

11.1 Acids & Bases

11.2 Titrations

11.3 Molar Mass through Titrations

11.4 Acid-Base Equilibrium: Ka & Kb

12 Equilibrium

13 Thermochemistry

13.1 Enthalpy, Enthalpy of Reactions, Spontaneity

13.2 Hess's Law

13.3 Big Mama Equation

13.4 Reaction Spontaneity, Energy & Heat Transfer

13.5 Specific Heat

14 Nuclear Chemistry

15 Organic Chemistry