Honors Chemistry Notes

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Contents

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 ($1s^22s^22p^3$)

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^+ - $1s^22s^22p^63s^23p^63d^{10}$
- cupric, Cu^{2+} - $1s^22s^22p^63s^23p^63d^9$

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

4.2 Ionic Nomenclature

4.3 Covalent & Acid Nomenclature

4.4 Mole Problems

4.5 Percent Composition

4.6 Empirical & Molecular Formulas

4.7 Oxidation Numbers

5 Reactions

5.1 Balancing Equations

5.2 Synthesis & Decomposition

5.3 Single Replacement, Double Replacement, & Combustion

5.4 Reaction Rates

5.5 Redox Reactions

5.6 Net Ionic Equations

6 Stoichiometry

6.1 Stoichiometry

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

7 VSEPR/IMFs

7.1 Types of Bonding

7.2 Bonding

8 States of Matter

9 Gas Laws

9.1 Kinetic Molecular Theory, Temperature, and Pressure

9.2 Gas Laws & Density

10 Solutions

10.1 Solutions, Colloids, Suspensions, Electrolytes & Solubility

10.2 Units of Concentration

10.3 Colligative Properties

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