Science

* After revisions and tests over time, hypotheses becomes theories, that can eventually become laws, after revisions and tests
  + Scientific laws are definite definitions of a phenomenon
  + Hypothesis is an educated guess
  + Revisions are done by many scientists in a field
  + Everything’s that’s not a law is subjected to change
* The scientific method is universal
* Science is collecting evidence and connecting them to discover the workings of the world
* Science is objective
* To experiment, form a hypothesis, design an experiment, collect data, create claim, connect data to claim, and repeat.
* A model can represent a concept
  + Black box is an unknown phenomenon

Measurements and Analysis

* SI
  + Base units
    - Length - meter m
    - Mass - kilogram kg
    - Time - seconds s
    - Electrical current - ampere A
    - Temperature - Kelvin K
      * 1 K = 273.15 oC
    - Amount of substance - Mole M
      * 6.022045 x 1023 particles
        + There are Avogadro’s number of atoms in 12g of carbon 12

6.022045 x 1023 amu weights 1 g

* + - * Moles to mass
        + Multiply moles by conversion factor (gram mass formula aka )
      * Mass to moles
        + Multiply mass in grams by conversion factor (reciprocal of GMF)
    - Luminous intensity - Candela cd
  + Prefixes
    - Tera (T) 1012
    - Giga (G) 109
    - Mega (M) 106
    - Kilo (k) 103
    - Deci (d) 10-1
    - Centi (c) 10-2
    - Milli (m) 10-3
    - Micro (µ) 10-6
    - Nano (n) 10-9
    - Pico (p) 10-12
* Density
  + 1 cm3 = 1 mL volume
* Precision
  + Ability to repeat task
  + Use standard deviation
* Accuracy
  + Ability to get close to what’s accurate
  + Use percent error
* Scientific notation
  + add/subtract
    - Make the exponent terms have same exponent, and add/subtract the non exponent term
  + Multiply/divide
    - For exponent term, add exponent for multiply, subtract for divide
* Uncertainty
  + Measurement is between 2 marks on measurer
  + Report all certain digits, +1 uncertain
    - +/- can indicates error margin
  + Percent error
  + Significant figures
    - Keeps numbers accurate by indicating to which place value accuracy is guaranteed up to
    - Significant
      * Non zeros
      * Zeros between non zeros
      * Trailing zeros if there’s a decimal
    - Not significant
      * Leading zeros
      * Trailing zeros when there’s no decimal
    - adding/subtracting
      * Round answer at place where the number with fewest significant figures after decimal ends the number
      * Propagation of error
        + When rounding 5, only round up when previous digit is odd
    - multiplying/dividing
      * Round answer to have the amount of figures as in the number with least significant figures

Everyone matters

Matter, Atoms, Elements, Compounds

Matter

* Has mass (quantity of matter) and takes volume (space)
* Properties of matter
  + Physical
    - Observable without altering composition or identity of matter
  + Chemical
    - Observable when chemical change occurs (alters identity)
  + Extensive
    - Depends on amount present
  + Intensive
    - Independent on amount present
* Changes
  + Physical
    - Doesn’t change composition
  + Chemical
    - Changes into different substance
* Seperation
  + Filtration
    - Separate solid residue from liquid filtrate, using filter
  + Distillation
    - Separates liquids
    - Substance with lowest boiling point is boiled into gas, rising to a flask, where it condenses into liquid
  + Chromatography
    - Separates substances based on solubility, ability to travel on surface, polarity
    - A solvent carries the compounds of a mixture across a surface
  + Crystallization
    - Liquid in aqueous solution evaporates, leaving behind solid
  + Decantation
    - When 2 substances don’t mix, let one substance settle to bottom, then pour out top substance
* States
  + Commonly liquid, solid, gas
  + Solid: definite shape, volume
  + Liquid: definite volume, indefinite shape
  + Gas: indefinite shape and volume
  + Energy change required to change states
* Mixture
  + Homogeneous
    - solution
    - Uniform composition
  + Heterogeneous
    - Varying composition
* Substances
  + Pure elements and compounds

Atomic Theory and Structure

* Atoms make up matter
* Atomic theories/history
  + 500 BC Democritus knows matter is make up of particles atomos
  + 100-1600 AD Alchemists found extensive chemical data
  + 1800 Joseph Priestley, Antoine Laurent Lavoisier found combustion is a reaction between matter and oxygen
    - Disproved, as can chemically combust without oxygen
    - Oxygen is needed for combustion if carbon is present
  + Law of conservation of mass
    - Mass can’t be created or destroyed
      * Disproven
  + Law of definite proportions
    - The elemental composure of different samples of the same compound are in proportion
  + Law of multiple proportions
    - The proportion of the mass of an element combined with a fixed mass of another element is a small whole number
  + Dalton’s postulate
    - Matter is made of tiny, spherical indivisible atoms
      * Disproved, as atoms can be broken down into electrons, protons, and neutrons
    - The same type of atoms are identical
      * Disproven, as the amount of neutrons can vary by isotope, and allotropes exist
    - Different types of atoms aren’t identical
    - Atoms combine in whole number ratios
    - Atoms can’t be changed into another atom, produce new atoms or destroyed in a reaction
      * Disproven, as different types of atoms can be made in nuclear reactions
  + JJ Thomson
    - Charge to mass ratio = -1.76 \* 108 coulomb/g
    - In a tube, he had a cathode (-) and anode (+) side that created rays of fluorescence, that were repelled by the negative pole of a magnet, and when blocked, shown to originate from the cathode end, proving there are negatively charged electrons
    - Plum pudding
      * A positively charged sphere with electrons embedded in it
  + 1990 atom model
    - A positively charged sphere with electrons imbedded
  + Ernest Rutherford: nucleus - 1911
    - Shot alpha particles (+) at gold foil while detecting screen was around it. Most alpha particles traveled straight, proving a lot of space around nucleus. Other alpha particles were deflected, proving presence of nucleus, positively charged.
  + RA Millikan - 1909
    - Oil drop
      * X-ray runs through telescope to negatively charge oil droplets, and oil gets attracted to the positively charged plate above it. Knowing density of oil and the force of gravity, he concluded that electrons have charge of -1.6 \* 10-19 coulomb
        + Proved mass of electron to be 9.11 \* 10-31 kg
  + Niels Bohr - 1912
    - Electrons orbit in discrete energy levels called orbits, model referred to as the planetary model
    - Gain energy, move out shell. Lose energy, move in shell
  + Henri Becquerel: radioactivity
    - Spontaneous emission of radiation
    - Ernest Rutherford
      * Alpha particles (+)
      * Beta particles (-)
      * Gamma rays (0)
  + James Chadwick 1932: neutron
  + Quantum mechanics
    - Electrons in the electron cloud, moving in wave like pattern. Likely to find electron in orbital
* Composed of proton (+), neutron (0), electron (-)
  + Mass of proton (1.6726 \* 10-27 kg) is 2000x mass of electron (9.1096\*10-31
  + Neutrons (1.675 \* 10-27 kg) are slightly heavier than protons
  + Protons and neutrons are about 1 amu
  + Quantum mechanics make electrons lose energy in short bursts
* Atoms have 0 charge
  + Otherwise, it’s an ion
* Atomic structure
  + Bohr model - planetary model
    - Nucleus made of neutrons and protons, with electrons circling in shells around nucleus
    - First shell/orbit/energy level is n=1, 2nd n=2 and so on
    - Outermost shell with electrons carry valence electrons
    - Electrons move move outwards a shell when gaining energy, inwards a shell when losing
    - Each shell can only hold 2n2 electrons (n is shell number)
    - An atom is stable when there are 8 valence electrons (octet rule)
      * Exception: 2 valence electrons if there’s only one shell (duet rule)
      * When unstable, they’re tend to bond
    - When an electron gains energy, it can get excited and move to a higher energy level (excited state)
    - An excited electron can release energy to return to its original energy level (ground state)
    - Energy used to change states: ΔE = Eoriginal energy orbit-Enew energy orbit=Ephoton=hv, where E=energy, v=velocity and h=Planck’s constant
    - Limitation: this model only accounts for hydrogen
    - Electron configurations: electrons in shell 1-shell 2- shell 3… ex: 2-8
    - Bonding pairs: electrons used for covalent bonds
    - Lone pairs: electrons not used for covalent bonds
  + Quantum mechanics
    - Energy: λ=, where λ=wavelength, m=mass
    - Heisenberg uncertainty: measuring something will disturb it
    - Electrons don’t orbit in circles, they exist in orbitals (location where it’s probable to find an electron)
    - Energy levels are in numbers
    - Each energy level has sublevels
      * Each sublevel can carry 2 electrons that spin in opposite directions (Pauli exception)
      * Sublevel s
        + Spherical
        + Exists in energy level 1 and above
        + One per level
      * Sublevel p
        + Dumbbell shaped
        + Exists in level 2 and above
        + 3 per level

Px exists along x axis

Py exists along y axis

Pz exists along z axis

* + - * Sublevel d
        + “X” shaped
        + Exists in level 3 and above
        + 5 per level

Dxy on xy plane

Dxz on xz plane

Dyz on yz plane

Dx2-y2 cut by xy plane

Dz2 is like pz

* + - * Sublevel f
        + Intricate shapes
        + Exists in level 4 and above
    - Electron configuration: energy level # followed by sub energy level followed by # of electrons in it, ex: 1s21p2 means 2 in 1s, 2 in 1p
    - Noble gas notation
      * Put in bracket symbol of the noble configuration already in atom, then outside bracket, put notation for the valence orbitals
    - Orbital diagram
      * Box with orbital name underneath
      * Arrow inside box pointing up or down to show spin
    - Aufbau principle
      * Electrons occupy lowest energy orbitals available
      * Order: make a right triangle, first row: write down all sublevels possible in energy level 1, repeat for the other levels
      * Fill sublevels diagonally, start from first point on the triangle’s hypotenuse, go on to next point once a diagonal is filled
    - Hund’s rule
      * Electrons of the same spin occupy orbitals before opposite spin does
* Ionic charge
  + Subtract amount of protons - amount of electron (ex: 6-8=2-) number then sign
  + Positive ion: cation
  + Negative ion: anion
  + Electrostatic forces
    - Ions of opposite charges attract, else repel.
* Atomic emission spectra
  + A discontinuous spectrum that can be used to identify an atom
  + When an excited atom returns to ground state, a unique amount of energy is released when electron goes from one discrete energy level to another. Energy needed is dependent on which atom it is.

Elements and Compounds

* The same type of atoms makes elements
  + Elements can’t be broken into other elements
* Different elements make compounds
* Organic compounds
  + Has carbon
* Inorganic compounds
  + Everything else
* Ionic compound
  + Ionic bonded
  + Naming
    - Start with cation’s name
      * Some metal cations have multiple possible amount of valence electrons. Use roman numerals to indicate amount
    - Then, anion’s name, with ending of -ide
    - Hard
    - Brittle
    - High melting point
    - Isn’t electrolyte (ability to conduct electricity) when solid
      * When aqueous, dissociates into mobile ions, making it electrolyte
    - Crystals in regular shapes when solid, and cleaves on planes
* Network solids
  + Covalently bonded
  + Large macromolecule held by covalent bonds
  + Hard
  + Brittle
  + High melting point
  + Not electrolyte
  + Solid crystals in regular shapes cleaving along planes
* Molecular compound
  + Covalent bonded
  + Naming
    - Start with atom with less electronegativity
    - Then the atom with more electronegativity, followed by -ide
    - When there’s multiple of a type of atom, use Latin prefix it to indicate it (mono, di, tri, tetra, penta, hexa, hepta, octo, nona, deca)
      * Don’t use mono for first element
  + Basic molecules are monomers
  + Multiple molecules make polymers
  + Discrete molecules held by intermolecular forces
  + Soft
  + Low melting point
  + Nonelectrolyte
    - Unless polar and aqueous
  + Like dissolves like (polar-polar, non-non)
* Diatomic molecule
  + Has 2 of the same atoms
  + H, N, O, F, Cl, Br, I naturally found in 2
    - N, O, F, Cl, Br, I forms a “7” on the periodic table
* Polyatomic ions
  + Charged stable particles with different types of atoms bonded covalently
  + Can form ionic compounds
  + Parenthesis around it in formula
  + Usually has oxygen
    - Most common form ends in -ate
    - One less oxygen then common ends in -ite
    - 2 less than common starts with hypo- and ends in -ite
    - 1 more oxygen than common starts with per- and ends with -ate
  + Acids ending in -ic ends in -ate when polyatomic ion by losing H
  + Acids ending in -ous ends in -ite when polyatomic ion by losing H
* Formulas
  + Empirical
    - Number of elements present are written as reduced ratios, noted by subscripts
  + Molecular
    - Exact number of elements present noted by subscript
  + Condensed structural formula
    - Left to right order to indicate order of atoms
  + Coefficients tell amount of molecule there is, usable for ratios
  + In formulas, charge of first element is subscript of second, charge of second is subscript of first
* Net ion
  + Aqueous compounds are written as their ions, ions identical on both sides are canceled out, net reaction are remaining items.
* Lewis structure
  + Shows the bonding
  + Central atom is least electronegative atom
  + Put other atoms around central
  + Surround dots around atom to indicate amount of valence electrons
    - Add 2 on top first, then add clockwise
  + Connect dots of different atoms at indicate bonding: dots are now shared
  + Make sure each atom are stable unless BH3, SF6, Co(NH3O)6
  + If it’s an ion, put a bracket around it, and in superscript, note its charge
  + Resonance structures
    - Multiple Lewis structures can be drawn
    - Real form is between the possible structures
  + Formal charges
    - Of each atom: valence electrons - bonds - lone electrons (dots). sign then number
* Geometry VSEPR - Valence Shell Electron Pair Repulsion
  + Electrons repels each other, giving the shape of the molecule
  + Linear
    - 180 degree repulsion
    - 1 or 2 bonding pairs
  + Trigonal planar
    - 120 degree repulsion
    - 3 bonding pairs
  + Tetrahedral
    - 109.5 degree repulsion
    - 4 bonding pairs
  + Trigonal pyramidal
    - 107 degree respulsion
    - 3 bonding pairs and 1 lone pair
  + Bent
    - 104.5 degree repulsion
    - 2 bonding pairs and 2 lone pairs

The Periodic Table

* First letter of element symbol is capitalized
* Atomic number is the amount of proton an atom has, determines type of atom
* Mass number is amount of proton and neutrons in atom
  + In atomic mass units
    - Carbon 12 has a mass of exactly 12 amu
      * 1 amu = 1/12 of mass of carbon 12
* Atomic mass is weighed average of the mass numbers of naturally occurring isotopes based on abundance
* Isotopes has same atomic number, but different mass number due to different neutron count
* Molar mass is mass of a mole of the elements (gmf)
* Electron configuration from left to right, indicates electron count from innermost shell to outermost shell
* Organization
  + Ordered by increasing atomic number
  + Columns are groups
    - Groups share the same amounts valence electrons, except for transition metals
    - Similar properties
    - Group 1: alkali metals
      * H is a gas
        + Reactive
        + Colorless
        + odorless
      * Readily bond, has 1 valence electron
      * Very reactive
      * Almost never found in pure form
      * Silver and shiny
      * Low density
    - Group 2: alkaline earth metals
      * has 2 valence electrons
      * Bit less reactive than group 1
      * Silver
      * Denser than group 1
    - Group 3-12: transition metals
      * Range of reactivities and properties
      * Good conductors
      * Shiny, colorful
      * Higher densities and melting points
      * Has few possible valence electron count, indicated with roman numerals when writing compound/molecule name
    - Group 13 has 3 valence electrons
    - Group 14 has 4 valence electrons
      * Easily bond by giving or receiving
      * All but carbon is dangerous in life, as carbon, like the rest of group 14, can easily bond, but only carbon carries out life processes. If another group 14 atom bonded in place of carbon, carbon can’t do its job.
      * Group 15 has 5 valence electrons
      * Group 16 has 6 valence electrons
    - Group 17: halogen (X)
      * Very reactive, has 7 valence electrons
      * Forms salts (ionic compound) with metals
    - Group 18: noble gases
      * Doesn’t bond with other elements, as they are stable
      * Colorless, odorless
      * Low concentration on Earth
  + Lanthanides and actinides are rare
    - Also transition metals
    - Lanthanides are shiny and reactive
    - Actinides are radioactive
    - Number 95 and 103 can only be manufactured in lab
  + Rows are periods
    - Periods share the same amount of energy levels
  + Metal: easily becomes cations
    - Most elements are metals
  + Metalloids: easily becomes cation or anion
    - By “staircase”
  + Non-metals: easily becomes anions
  + Br, Hg are liquid at room temperature
  + H, N, O, F, He, Ne, Ar, Kr, Xe and Rn are gas at room temperature
  + Rest are solid at room temperature
* Trends
  + Mendeleev’s periodic law
    - When elements are ordered by atomic number, chemical properties repeat periodically
    - Used to predict properties of undiscovered elements
  + As atomic number increases across a period, more protons are added, increasing the nuclear charge, while electrons are being added to the same energy level, so the valence electrons are more strongly attracted to the nucleus, making the atomic radius smaller, first ionization energy and electronegativity to increase. (noble has no electronegativity)
  + As atomic number increases down a group, more protons are added, increasing the nuclear charge, while electrons are being added to higher energy levels farther from the nucleus, allowing for more inner electrons to shield the valence electrons from the nucleus’ pull, leading to larger atomic radius and smaller first ionization energy and electronegativity.
  + Atomic radius: half the distance between adjacent nucleus
  + Electronegativity: atom’s ability to attract electrons for bonding
  + First ionization energy: energy required to remove first valence electron

Forces

* Intramolecular
  + Between parts of molecule
  + Bonds
    - Ionic
      * For ionic compounds
      * Electronegativity difference between atoms is >1.7
      * Usually between metal (usually cation) and nonmetal (usually anion)
      * Atom with less valence electrons donate electrons to atom with more, until all atoms are stable (satisfy octet/duet rule whichever applicable)
      * Receiver: anion, doner: cation
      * Cation (+) gets attracted to anion (-)
      * In formulas, charge of anion is subscript of cation, charge of cation is subscript of anion
      * Named cation followed by anion with -ide ending.
        + Some metals have multiple possible charges, so their charge is indicated with roman numerals
    - Covalent
      * For molecules
      * Usually between nonmetals
      * Polar bear
        + Electronegativity .4<x<1.7
        + Polar means one atom has more pull on shared electrons than other
        + More electronegative atom is slightly negative
      * Nonpolar
        + Electronegativity <.4
      * Electrons are shared
      * Named first by the atom with lower electronegativity, followed by the atom with higher electronegativity, with -ide ending.
        + If there’s multiple of an atom, use greek prefixes to indicate (mono, di, tri…)
      * In formulas, charge of first element is subscript of second, charge of second is subscript of first
    - Metallic
      * Between metals
      * sea of mobile electrons
    - Bond dipole
      * Points from positive to negative to show net pull on electron
      * Negative has higher electronegativity
    - Molecule dipole
      * Net vector of bond dipoles
      * If molecule has symmetrical distribution of charges, (no net vectors) molecule is nonpolar. Can also check shape
  + Forming bonds is exothermic and increases stability
  + Breaking bonds is endothermic and decreases stability
* Intermolecular
  + Between molecules
  + Van der Waals/London dispersion force
    - * all
      * Electrons of a molecule repels electrons of another, shifting electron cloud density, causing momentary dipoles. Opposite charges attract
  + Dipole-dipole
    - Attraction between opposite charges of molecules with dipoles
    - Hydrogen bond
      * Strong dipole-dipole between H and highly electronegative N, O or F.
  + Molecule-ion
    - Ion attracted to part of polar molecule that has opposite charge
  + Interactions are stronger when more electrons are involved
  + Strong to weak: molecule ion, hydrogen, dipole-dipole, LDF
  + Strength of IMF determines energy required to change physical states (stronger needs more, increasing melting, boiling points)

Stoichiometry

* Moles to mass
  + Multiply moles by conversion factor (gram mass formula aka )
* Mass to moles
  + Multiply mass in grams by conversion factor (reciprocal of GMF)
* Molar mass is in g/mol, atomic mass is in amu, they’re numerically equal
* To find mass of a compound, take mass of an element and multiply by number of occurence in compound, and add up the masses
* Stoichiometry
  + Determines relationship in amount of reactants and products
  + Use balanced equations
  + Convert to masses to mole
  + Find limiting reagent
    - If < , then reagent A is limiting. Else, B is limiting
  + Moles of limiting reagent present \*
* Leftover reagents
  + Do stoichiometry, but use excess reagent as substance in question
  + Subtract result from amount of excess reagent given at start
* Percent yield
* Dimensional analysis
  + Keep units when adding/subtracting/multiplying/dividing (basically treat them like variables)
  + Only add/subtract when units are same
  + When dividing, units can be simplified out (vertically + cross)
  + When multiplying, units can turn exponential when multiplied to same unit
* Percent composition
  + Take mass of whole compound, then mass of element in question (all atoms of that element).
  + Take percentage out of whole compound 100
  + Percents can be used to relate back to empirical formula to ID compounds

Solutions

* Homogeneous
* Solute dissolved in solvent
* Solvent > solute
* Phase change to continuous state typically only occurs to solute
* Like polarity dissolves like polarity
* Solvation
  + Solvent clusters around solute
  + Intermolecular attraction between solute particles are weaker than intermolecular attraction between particles of solute and solvent
  + Intermolecular attractions form between solute and solvent particles, that disassociate intermolecular attractions between solute particles
    - In ionic solutes, ions dissociate too
  + If solvated particles attract again, crystal forms
* Solubility
  + Amount of solute that can be dissolved
  + Factors
    - agitation: stirring, shaking increases rate of solubility
    - Surface area: increases will increase rate of solubility
    - Temperature: increasing temperature increases rate of solubility for solid solute, while decreasing for gaseous solute
    - Pressure: no impact unless gaseous solute, where increase in pressure increase solubility
    - Polarity: Like dissolves like (polar-polar, nonpolar-nonpolar)
  + Rate of solvation > rate of crystallization
* Aqueous
  + Water as solvent
  + Water is universal solvent
    - Can attract to both cations and anions
    - O of water is slightly negative (attract +), H of water is slightly positive (attract -)
  + Hydration
    - Solvation in water
    - Dissociates ionic bonds
      * O bonds to positive ions
      * H bonds to negative ions
* Saturation
  + Concentration
    - M(molarity) = mol of solute / L of solution
    - Ppm(parts per million): percent mass \* 1000000
      * Percent mass = mass of solute / mass of solution
  + Subscripts indicate moles of that ion present
  + Comparisons can only be made at same temperature and pressure
  + Unsaturated: less solvated particles than saturated
  + Saturated: maxed out; any more solutes added will crystalize
  + Supersaturated: more solvated particles than saturated
    - Since solvated particles capacity increases as temperature increases, add solute at higher temperature, then bring temperature down so that amount of solvated particles are higher than the capacity at new temperature
* Dilution
  + Decrease saturation: M(molarity)initial \* V(volume: L)initial = Mfinal \* Vfinal
* Electrolytes
  + Substance that conducts electricity when dissolved
    - Solute is ionically bonded (can dissociate into mobile ions)
    - Weak
      * Not all compounds are dissociated
    - Strong
      * All compounds are dissociated
  + Non electrolytes is opposite
    - Solute is covalently bonded
* Colligative properties
  + Based on amounts
  + Freezing point depression
    - Increasing amount of solute particles lowers freezing point
  + Boiling point elevation
    - Increasing amount of solute particles raises boiling point
* Dipole moment
  + When + and - centers don’t coincide
  + µ = e \* d, in debye units

Reactions

* Reactants on left side ----> products on right
  + 2 arrows running both direction means reversible
* Coefficients indicates amount of moles for mole ratio
* State of matter in () next to molecule formula as subscript
  + (s), (g), (l), (aq) -dissolved in water, is homogeneous
* Requires gaining/losing energy to change states
* Both sides balanced with coefficients. amount of an element on reagent side = amount on product side. law of conservation: no matter made or destroyed
* Displacement
  + Precipitation
    - metathesis/double replacement
      * A+B-(aq) + C+D-(aq) -> A+D-(aq or s) + C+B-(aq or s)
    - A solid precipitates from the liquid
  + Single replacement
    - A+B- + C+ or - -> A+C- + B- or C+B- + A+
* Synthesis
  + Substances combine to create a homogeneous substance
* Decomposition
  + Substances break down to simpler substances
* Collision theory
  + In order for colliding particles to react, they need to collide with sufficient energy and proper orientation
  + Reaction rate is based on rate of effective collisions
  + Reaction mechanism: series of steps involved in a reaction
  + Increasing temperature increases kinetic energy, which increase likelihood of effective collisions, which increases reaction rate
  + Increasing surface area exposes more particles for reaction to occur with, increasing reaction rate
  + Increasing concentration provides more particles to react with, increasing reaction rate
  + Increasing pressure in a system with gas reagents increases reaction rate as the particles would be closer together
  + The presence of a catalyst increases reaction rate by lowering the activation energy

Kinetics and Equilibrium

* Gas
  + Barometers (Evangelista Torricelli)
    - A tube in a cup. Tube is vacuuming liquid up, while gravity is pushing liquid down.
    - Pressure = density of liquid \* gravity’s force \* height of liquid in tube
  + Manometers
    - Gas in bulb connected to S shaped tube with liquid in it. Liquid would be closer to bulb if bulb’s pressure is less than gravity’s, otherwise, other way around.
    - Bulb pressure = atmospheric pressure - liquid height difference (side closer to bulb - other side)
  + 1 atm = 760 mmHg = 760 torr = 101325 Pa = 101.3 kPa = 14.7 psi = 1.013 bar
    - 1 Pa = 1 Newton/m2
  + Kinetic molecular theory (ideally…)
    - Individual particles travel in a straight line, in random motion
    - Particles have elastic collisions with each other and wall of containers
    - No forces of attraction exist
    - Compared to the volume of container, volume of the individual particles are negligible
  + Laws
    - Variables
      * P = Pressure (atm)
      * V = volume (L)
      * T = Temp (K to prevent 0s and negs conundrum)
      * n = moles
      * To find relationship between 2 variables, make the other variables constant, write 2 equations, one for before, other for after. Substitute constant together to form 1 equation
    - R=
    - 1 = initial
    - 2 = final
    - Z = compression factor
      * A ratio between volume per 1 mole and molar volume of ideal gas
    - Boyle’s Law
      * P1V1=P2V2
      * As volume decreases, pressure increases
      * Inversely proportional: one goes up, other goes down
      * Curve downwards graph
    - Charles’ Law
      * As heat increases, volume increases
      * Directly proportional: one goes up, other goes up
      * Straight, positive slope
    - Avogadro's Law
      * As mole increase, volume increases, directly proportional
      * 1 mole of ideal gas takes up 22.4 L
    - Lussac’s Law
      * As temperature increases, pressure increases, directly proportional
    - Combined gas law
      * Based on Lussac, Boyle and Charles’ laws
    - Ideal gas law
      * PV=nRT
      * Gas ideal when particles move in random motion and don’t interact. They are never perfect in real life
        + The higher the temperature and the lower the pressure, the more closer the gas is to perfection
    - Van der Waals
      * (P+(​))(V−nb)=nRT
      * Gases aren’t ideal in reality, this takes into account that real molecules only travel within a finite space in a container
      * a and b are constants depending on gas
      * takes into account of decreased pressure from molecule attraction
      * nb takes into account that only part of container’s volume have gas molecules
    - Standard Temperature and Pressure (STP)
      * 273 K and 1 atm
  + Total pressure = sum partial pressures, or the sum of pressures of each molecule
  + Vapor pressure
    - Upwards pressure exerted by gas above liquid
    - Pressure increases as temperature increases, or as IMF weaken
    - Liquids slowly evaporate due to small bits of particles escaping surface of liquid, which causes vapor pressure
    - Once vapor pressure reach/pass atmospheric pressure, vaporization occurs. Temp required to boil @STP is standard boilpt
* Thermodynamics
  + Chemical reactions require a change in energy between system and surrounding
  + Heat energy (q), measured in ~~juuls~~ joules (J), 1000 J in kJ
    - Can’t be negative
    - Heat moves from warm to cool
  + Temperature is average of kinetic (motion) energy of particles in sample
    - Measured in K or oC (1 oC = 273 K)
    - Calorie: heat needed to raise 1oC of 1g of water
  + Heat is total kinetic energy of particles in sample
    - Heat can either raise temperature or change a matter’s state
  + heating/cooling curves (phase changes)
    - X axis: heat added/removed. Y axis: temperature
    - Positive slopes indicate increase in temperature, kinetic energy
    - Negative slopes indicate decrease in temperature, kinetic energy
    - Flat lines indicate no change in temperature, kinetic energy
      * Phase change is occuring
      * Increase in heat over a flat line in increase in potential energy
      * Decrease in heat over a flat line is decrease in potential (stored) energy
  + Reactions
    - Exothermic
      * Releases energy as product
      * Has negative enthalpy value (ΔH)
      * Forms bonds
      * Required for solidification (ΔHsolidification) and condensation (ΔHcondensation), deposition (g to s) phase changes
      * Is spontaneous (happens on its own), increases stability
    - Endothermic
      * Absorbs energy as reactant, stored in chemical bonds
      * Has positive enthalpy value
      * Break bonds
      * Required for melting (ΔHfusion) and evaporation (ΔHvaporization), sublimation (s to g) phase changes
    - Specific heat (Csp)
      * Heat needed to raise 1 g of substance by 1oC, in J/(g\*oC)
    - Heat involved in thermal process: calorimetry
      * q = g \* Csp \* ΔoC
      * When there is a state change, use formula once for each state. Adjust ΔoC based on the range possible in the state. Adjust Cspbased on the corresponding state. Then, find energy used in phase change by multiplying mass by enthalpy of that state change. Add up all values
  + Enthalpy
    - ΔH is change in energy going from a state to another, in kJ/mol
    - ΔHrxn = Hproducts - Hreactants
    - Reversing reaction makes sign of ΔH opposite
    - If a reaction is multiplied by a factor, ΔH is multiplied too
    - Hess’ law
      * ΔH is sum of energy in reaction, which will be equal no matter the order of steps in the reaction
    - Standard enthalpy of formation (ΔHof)
      * ΔH to change 1 mole of standard state compound(s) to another state/compound
        + Standard state: condition of substance at 25oC
      * Direct formations
        + Elements in compound are the same elements that were in reaction to form compound or broke apart compound.
    - Standard enthalpy of reaction (ΔHorxn)
      * Overall ΔH in reaction
      * ΔHorxn = Σ ΔHof for products - Σ ΔHof for reactants
      * Indirect reactions
        + Rewrite reaction as balanced (can balance with fractions) reactions that are direct formations, with states indicated
        + For each reaction, write their ΔHof
        + Add up ΔHof
        + Add up reactions and reduce like compounds on both side to get net reaction, with target as product
  + Entropy (ΔS)
    - Disorder of a system
    - Systems in nature tend to favor low energy and high entropy
    - S -> l -> g states in order of increasing entropy
  + Potential energy diagrams
    - X axis: reaction coordinate. Y axis: potential energy
    - From 0 to plateau at start is PE of reagents
    - From 0 to plateau at end is PE of products
    - From 0 to maximum point is PE of activated complex
    - From plateau at start to maximum point is forward activation energy
    - From plateau at end to maximum point is reverse activation energy
    - From plateau at start to plateau at end is ΔH
* Equilibrium
  + In reversible reactions in closed systems, achieved when rate of forward reaction = rate of reverse reaction
  + Visually there are no changes
  + Dynamic phase equilibrium: reversible phase changes, achieved at flat line parts of heating/cooling curve
  + Dynamic solution equilibrium: in a saturated solution. For every solute particle added, a particle precipitates
  + Chemical equilibria: reversible chemical reactions
  + Le Chatelier's principle
    - In a system at equilibrium, if stress is added, the system will shift in a way that relieves the stress, resulting in a new equilibrium
    - If heat is added to the system, the system will favor the endothermic reaction
    - If the concentration of a reagent is increased, the system will favor the forward reaction
    - In a system with gaseous species, when pressure is increased, the system will favor the direction that produces less moles of gas

Acids and Bases

* Acids
  + Compound with H+ as cation
  + Sour
  + Form H2 gas on reaction with some metals
  + Strong acids means more H+ ions dissociate than with weak acid
  + Acids with C are typically weak
  + electrolyte
  + Arrhenius: has H+ as cation
  + Bronsted: donates H+
    - Becomes conjugate base after donation
  + Monoprotic
    - Has a proton that can be lost (Arrhenius acids) or donated (Bronsted acids)
    - Diprotic when it has 2, triprotic for 3
  + Nomenclature
    - Without a polyatomic anion
      * Hydro+anion+ic ending
    - With a polyatomic anion
      * If anion ends in ite, replace with out
      * If anion ends in ate, replace with ic
  + pH
    - Power of hydrogen
    - pH = -log[H+]
      * Log is x in 10x=number.
    - [H+] = 10-pH
    - pH = 14 - pOH
    - [ ] means concentration (molarity) of
    - pH < 7 is acidic, 7 is neutral and pH > 7 is basic
    - From 1 and onwards, each increase is 10x more basic/lower [H+]
    - Indicators
      * pH Below the range shows first color
      * hH In the range shows mix of both colors
      * pH Above the range shows the second color
* Bases
  + Compound with OH- as anion
  + Bitter
  + Slippery
  + Electrolytes
  + Strong bases have cation from groups 1 or 2
  + Arrhenius: donates OH-
  + Bronsted: receives H+
    - Becomes conjugate acid after acceptance
  + pOH
    - Power of hydroxide
    - pH = -log[OH-]
    - [OH-] = 10-pOH
    - pOH = 14 - pH
* Amphoteric: sometimes a base sometimes an acid
* Water
  + slowly self ionizes to H+ and OH-, reversible
  + Has pH and pOH of 7, as amount of H+and OH- are equal
  + 2H2O ⇋ OH- + H3O+
  + Adding acid to water releases hydrogen ions, which reacts with water to form hydronium, which reacts with existing hydroxide to form water
  + Adding base to water releases hydroxide, which reacts with existing hydronium to form more water
* Weak acids and bases have reversible reactions
* Neutralization reaction
  + When neutralized, forms a salt in water
  + Is a double replacement reaction
  + Everything but hydrogen and hydroxide are spectator ions
  + Neutralized when moles of acid = moles of base
    - M(molarity)acid \* V(volume in L)acid = Mbase \* Vbase
    - If acid isn’t monoprotic, multiply MAVA term by amount of H+ it has
    - Likewise for base with multiple hydroxides, multiply MBVB term by amount of OH-
* Gas forming reaction
  + Single replacement reaction between a metal more active than hydrogen and an acid, forming hydrogen gas and a salt
* titration
  + Buret contains acid/base dripping into beaker containing indicator and base/acid. Amount in buret and beaker before starting is record. Drip content from buret into beaker until indicator changes color. Record amount in buret and beaker after.
  + Curves
    - X axis: volume of acid/base added. Y axis: pH
    - As base is added to an acid, pH increases a bit, then stays relatively constant until enough base (equivalence point) is added to neutralize the acid. The, the pH sharply increases, passing through the threshold of the indicator. It then remains relatively constant again.
    - Backwards for adding acid to a base

REDOX and Electrochemistry

* Redox/oxidation-reduction
  + Not a double replacement reaction
  + Both oxidation and reduction takes place, mass and charge conserved
  + Is spontaneous (doesn’t require added energy) when species being oxidized is more active than species being reduced for metal, other way around for nonmetals
  + Oxidation
    - atom/ion loses electron, increasing oxidation number
    - Reducing agent causes oxidation
    - Less electronegative = easier to oxidize
  + Reduction
    - atom/ion gains electron, decreasing oxidation number
    - Oxidizing agent causes reduction
  + Disproportional is when an oxidizing or reducing agent is only some of the same type of ion
  + Oxidation number
    - Indicates charge once reaction occurs
    - Elements alone have oxidation number of 0, unlike ions
    - Hydrogen has oxidation number of +1, unless when bonded to a group 1 or 2 metal
    - Halogens have oxidation number of -1
    - Group 1 has oxidation number of +1
    - Group 2 has oxidation number of +2
    - Fluorine has oxidation number of -1
    - Oxygen has oxidation number of -2 except in H2O2 and OF2
    - Sum of oxidation numbers in compound equals charge
  + When balanced
    - Separate into half reactions: a oxidation half and a reduction half
    - Balance ions except for O and H
    - Add H2O on other side to balance O
    - Add H+ on other side to balance H
    - Add electrons (e-) to balance charge
    - Multiply both half reactions to make both have same amount of electrons
    - Simplify (find difference and substitute into greatest number, eliminate lowest one) ions that are same on one side of a half reaction as the other side of the other one.
    - Combine half reactions
    - If product is base instead of acid: add OH- on both side to neutralize H+ present. OH- and H+ on same side becomes H2O
* Batteries
  + Voltaic
    - Spontaneous reaction converting chemical energy to electrical energy
    - 2 half cells, each containing an aqueous solution and an electrode
    - A salt bridge to allow ions to flow to prevent charge buildup
    - A wire connecting the electrodes
    - Wire passes through a voltmeter to measure electricity
    - Electrode of more active metal oxidizes, is anode. Electrode of less active metal reduces, is cathode
    - Anode (positive terminal, site of oxidation) oxidizes, causing positive ions to be released into the solution. Anode shrinks. These positive ions flow through salt bridge to the cathode half cell. Electrons released from oxidation flows through wire to cathode, and generates electricity on its way there. Negative ions flow through salt bridge to anode half cell.
    - Cathode (negative terminal, site of reduction) electrode receives electrons, making it reduce. Cathode ions in solution receive the electrons and become atoms of the cathode electrode. Cathode grows.
  + Electrolytic
    - Nonspontaneous reactions
    - Uses a battery to oxidize a less active metal and reduce a more active metal
    - A cell with aqueous solution and 2 electrodes
    - Anode is less active metal, cathode is more active metal
    - Electrodes are connected by a wire
    - Cathode connected to battery’s negative terminal
    - Anode connected to battery’s positive terminal
    - Battery’s positive terminal causes anode to oxidize. Anode releases its electrons to the battery through the wire, and releases positive ions into the solution
    - Battery’s negative terminal causes cathode to reduce. Battery sends electrons to cathode through the wire. Positive ions from solution become atoms of the cathode. (can electroplate)
    - Electroplating: to coat cathode with anode using an electrolytic cell

Organic Chemistry

* Organic compounds: Compounds with carbon and hydrogen
* Hydrocarbons
  + hydrogen and carbon only
  + Aliphatic
    - Straight chain
    - Homologous series: As the chain increase by one carbon, it increases by x hydrogens
    - Saturated: has all the hydrogens possible given the carbon backbone
      * Alkanes
        + Single bond between carbons
        + Suffix: -ane
        + General formula: CnH2n+2
    - Unsaturated: doesn’t have all the hydrogen possible given the carbon backbone
      * Alkenes
        + Has a double bond between carbons
        + Suffix: -ene
        + General formula: CnH2n
      * Alkynes
        + Has a triple bond between carbons
        + Suffix: -yne
        + General formula: CnH2n-2
  + Aromatic
    - Closed chain
    - Exhibit resonance, or double bonds that move throughout the molecule
* Substituted hydrocarbons
  + Contains elements other than carbon and hydrogen
  + General formula
    - R is residue, which is everything except the functional group
    - Functional group is an atom or group of atoms with their bonds “substituted” into a hydrocarbon
* Allotropes: different configurations of a network structure
* Condensed structural formula
  + For each carbon of backbone, it is CHnumber of H attached to the C
  + Next carbon goes immediately after previous
  + Elements in () mean it branches out of the previous C
  + Isomers: molecules with the same molecular formula, but different condensed structural formula
* Hydrocarbon nomenclature for Regents
  + If is alkene or alkyne, start with carbon location number of the double/triple bond. Counting can start in any direction of the chain, but the number used is the lowest number
  + Then a dash
  + Then the correct prefix based on the number of carbons in the backbone
  + Then the correct suffix based on the presence of double/triple bonds
* Line structures
  + Line segments representing bonds are drawn, so that endpoints of segments represent C, unless another atom is written there
  + H is assumed to be present to make all atoms satisfy the octet rule
* Nomenclature
  + Prefix/suffix
    - -ane: single bonds between carbon
    - -ene: double bond between carbon
    - -yne: triple bond between carbon
    - Meth-: one carbon
    - Eth-: 2 carbon
    - Prop: 3 carbon
    - But-: 4 carbon
    - Pent-: 5 carbon
    - Hex-: 6 carbon
    - Hept-: 7 carbon
    - Oct-: 8 carbon
    - Non-: 9 carbon
    - Dec-: 10 carbon
    - N-: normal form
    - Cyclo-: formed in polygon shape
    - Branch
      * Methyl-: one branch of carbon
      * Ethyl-: one branch of carbon with another branch of carbon out of it
      * Isopropyl-: one branch of carbon that has 2 branches of carbon coming out of it
      * Tert-butyl-: one branch of carbon that has 3 branches of carbon coming out of it
      * Vinyl-: a double bonded branch of carbon
      * Phenyl-: benzene (ring of 6 carbons)
  + Linear alkanes
    - Carbons in a linear chain
  + Cycloalkanes
    - Carbons in polygon shape
  + Hydrocarbons
    - Compounds are named by their longest linear carbon chain(chain is drawn without lifting up writing utensil), or cycloalkane if present
    - Based on longest chain/cycloalkane, carbons are numbered in the direction that makes the numbers lowest where there is a branch
    - Numbers (separated by comma, in numerical order) for carbon location in chain is written, followed by a hyphen and type of branch. Type of branch is prefixed by multiples based on amount present (di-, tri-, tetra-, etc)
    - Branches are placed in alphabetical order, separated by a hyphen
  + Substituted hydrocarbons
    - Halide
      * Functional groups: F (fluoro-), Cl (chloro-), Br (bromo-) and I (iodo-)
      * Replaces an H in a hydrocarbon
      * Name: location of functional group on carbon backbone followed by a dash followed by prefix of functional group followed by the name of it as a hydrocarbon
    - Alcohol
      * Location of functional group on carbon backbone followed by a dash followed by the name of its hydrocarbon, with ending replaced with -ol
    - Ether
      * Hydrocarbon left of the functional group with ending replaced with -yl then space and hydrocarbon right of the functional group ending replaced with -yl, followed by ether
    - Ketone
      * Functional group: carbonyl
      * Name: location of functional group followed by a dash followed by name as hydrocarbon, with ending replaced with -one
    - Aldehyde
      * Functional group: carbonyl with an H attached
      * Name as hydrocarbon but with -al ending instead
    - Organic acid
      * Functional group: carbonyl with OH attached
      * Named as hydrocarbon, but with ending replaced with -oic, followed by a space then acid
    - Ester
      * Has a smell
      * Functional group: carbonyl with O attached
      * Named as hydrocarbon left of functional group, then a space followed by the whole substituted hydrocarbon with ending replaced with -oate
    - Amine
      * Start with location of the functional group on the carbon backbone, followed by a dash, then its named as a hydrocarbon, but with the suffix replaced to -amine
    - Amide
      * Start with name as a hydrocarbon, but replace its ending with -amide
* Reactions
  + Combustion
    - Reactants: oxygen and carbon
    - Product: water, carbon dioxide, light and heat energy
  + Substitution
    - Reactants: a saturated hydrocarbon and a diatomic molecule
    - Product: a saturated organic compound and a molecule
  + Addition
    - Reactants: an unsaturated hydrocarbon and a diatomic molecule
    - Product: a saturated organic compound
  + Esterification
    - Reactants: an acid and an alcohol
    - Product: an ester
  + Fermentation
    - Reactants: sugar and catalyst
    - Product: ethanol and carbon dioxide
  + Saponification
    - Reactants: fat and a base
    - Product: soap and glycerol
  + Polymerization
    - Addition
      * Unsaturated hydrocarbon as monomer forming a polymer
    - Condensation
      * Water is a product causing monomers to join in polymer
  + Cracking
    - Heating a hydrocarbon to break it into smaller hydrocarbons

Nuclear Chemistry

* Notation: top left is mass number, bottom left is the number of protons, right is a symbol
* Emission particles
  + Alpha (𝜶 or He)
    - Charge of 2, mass of 4
    - Can be stopped by paper
  + Beta (ß- or e-)
    - Charge of -1, mass of 0
    - Can be stopped by aluminium
  + Positron (ß+)
    - Charge of 1, mass of 0
    - Can be stopped by aluminum
  + Gamma (γ)
    - Emitted by all unstable nuclides
    - No charge no mass
    - Can be stopped by a concrete wall
* Stability
  + Having more than 83 protons make the nucleus unstable
  + Everything else depends on proton to neutron ratio, which stability requirements differs by proton number
* Transmutation
  + Sum of mass numbers are equal on both sides, sum of charge on both sides
  + Natural decay
    - An unstable nuclide decays into an emission particle and a more stable nuclide
    - Check the reactant nuclide’s mode of decay, write the corresponding notation on the product side, and add the correct nuclide on product side to make balance both top and bottom numbers
  + Artificial decay
    - Releases a lot of energy, especially fusion reaction
    - Mass defect: mass is converted into energy via E = mc2
    - Fission: Launch a neutron at a nucleus, causing it to split into 2 nuclides and release neutrons
    - Fusion: launch 2 nuclides at each other to make them combine into one nucleus
* Half life
  + Half life is the amount of time it takes for half of the sample to transmutate
  + a(0.5)t/h=b, where a is initial sample mass, t is time, h is half life and b is mass of sample after t time passes
* Uses
  + Can be used for dating, medical treatment
  + Can create waste that is biologically harmful

Add grah grah boom