UNIT 1: MATTER AND MEASUREMENTS IN CHEMISTRY

Sig Figs

- They measure the degree of certainty; the more sig figs, the more precise the measurement is.

Determining Sig Figs

- Exact numbers (numbers that you have not measured, that have no uncertainty) have unlimited sig figs
- Measured numbers: Nonzero numbers and middle 0's always count, leading 0's never count, trailing 0's only count if there is a decimal place.
- When you round sig figs, you can write them in scientific notation to get the correct number of sig figs.

Sig Figs in Mathematical Operations

- Add/Subtract - do the calculation, final answer should have the same number of decimal places as the number with the least decimal places
- Multiply/Divide - do the calculation, final answer should be rounded to the amount of sig figs as the number with the least sig figs
- When measuring always add 1 uncertain digit

Matter

- Elements contain 1 type of atom, found on the periodic table
- Compounds contain 2 or more different types of atoms, chemically bonded together
- Molecule is 2 or more atoms chemically bonded together but they can be the same
- Mixtures are a blend of 2 or more substance
- Heterogenous mixtures are not uniform, ex chocolate chip cookies
- homogeneous mixtures are uniform, ex salt water
- Phase is any part of a mixture with uniform composition, ex choc chip cookies have 2 phases, choc & cookie
- Alloy is a homogenous mixture of metals
- Mixtures can be separated by physical means

Metric System/Conversions

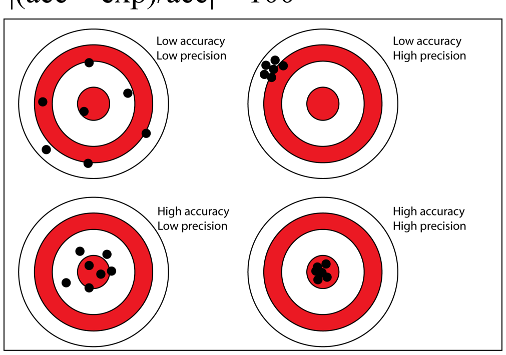
- Metric system, gram, meter, second, kelvin/celsius, liter
- Kilo, hecto, deca, base, deci, centi, milli

Dimensional Analysis

- 43 miles   (5280 ft)
- ---------- = 227040 ft
- (1 mile)
- Mile and mile cancel out, 43 * 5280 = 227040 and feet is the only unit left

Formulas

- D = m/v mass-grams, volume-mL or cm^3, density-g/cm^3, density of water is 1.00g/mL
- Percent error formula - |(acc * exp)/acc| * 100
- Precision vs accuracy



UNIT 2: THE ATOM

Scientists

- Democritus: atoms cannot be cut
- Dalton: solid sphere model
- Thomson: discovered the electron, cathode ray experiment, plum pudding model
- Millikan: discovered that electron has negative charge with oil drop experiment
- Rutherford: discovered the nucleus, gold foil experiment, nuclear model
- Bohr: discovered energy levels, planetary model
- Plank, Heisenberg, DeBroglie, Schrodinger: electron cloud model, small positively charged nucleus

Atomic Structure

- Proton: in nucleus, +1, relative mass 1, p+
- Neutron: in nucleus, 0, relative mass 1, n
- Electron: outside nucleus, -1, relative mass 0, e-
- Atomic number(Z) is the number of protons
- Number of protons = number of electrons
- Mass number(A) is the number of protons + neutrons
- C-12 - carbon, mass of 12 or 12/6 C - carbon, mass of 12, atomic number 6

Isotopes

- Isotopes are atoms of an element that have a different number of neutrons
- An ion is an atom that has a net electrical charge
- Cation: positive ion that lost an electron; it lost a negative charge so it's more positive
- Anion: negative ion that gained an electron; it gained a negative charge so its more negative

Moles/Mole Conversions

- 6.023*10^23/1 mole
- xg/1 mole; count up the masses of all the atoms

Nuclear Decay/Radiation/Radioactivity

- Alpha radiation; mass goes down by 4, atomic number by 2 (4, 2)$\boldsymbol{\alpha}$
- Beta radiation: mass unchanges, atomic number up by 1 (0, -1)β
- Gamma radiation: mass unchanges, atomic number unchanged (0, 0)γ

Formulas

- Average atomic mass = (Mass of isotope 1 * percent of isotope 1) + (Mass of isotope 2 * percent of isotope 2)/100 amu
- Half life formula: N(t) = N(initial)(½)^(t/half life)

UNIT 3: THE PERIODIC TABLE AND ELECTRON CONFIGURATIONS

Periodic Table Basics/Structure

- Each level can hold 2, 8, 8, 18, 18, 32 according to bohr's model
- 118 known elements, 1 - 98 are naturally occurring
- Period - horizontal row, 7 periods; Group/Family - vertical column, 18
- At room temperature: H, N, O, F, Cl, He, Ne, Ar, Kr, Xe, Rn are gasses and Hg, Br are liquids
- Metals - left of the staircase, lustrous, malleable, ductile, efficient conductors of heat and electricity
- Non metals - right of staircase + H, Metalloids - on the staircase - B, Si, Ge, As, Sb, Te, Po, At
- Diatomic molecules: Br2, I2, N2, Cl2, H2, O2, F2; Triatomic and others O3, P4, S8

Periodic Table History

- John Newlands arranged elements according to chemical properties and in order of increasing atomic mass. He also came up with the Law of Octaves; all elements in a given row have similar physical and chemical properties, this repeats every 8 elements
- Dimitri Mendeleev ordered by increasing atomic mass and similar chem properties in groups, he realized that elements do not fit in order of increasing atomic mass
- Henry Moseley arranged elements by increasing the atomic number. Discovered Periodic Law; when elements are arranged by atomic number, elements with similar properties appears at regular intervals

Periodic Table Properties & Trends

- Nuclear Charge: it equals the number of protons equal to atomic number
- Energy Levels: elements in period 1 have 1 energy level, period 3 has 3 energy levels, ect
- Electron Shielding: Across a period; nuclear charge/atomic number/# protons increase and energy level/levels of electrons stays the same so there is a stronger coulombic attraction. Down a group; nuclear charge increases but energy level/number of electrons in the way also increases so there is a weaker coulombic attraction and electron shielding increases
- Atomic Radius: decreases across a period, increases down a group
- Ionization Energy: the energy required to get rid of an e-; increase across a period, decreases down a group
- Electronegativity: the energy required to attract an e-; increase across a period, decreases down a group

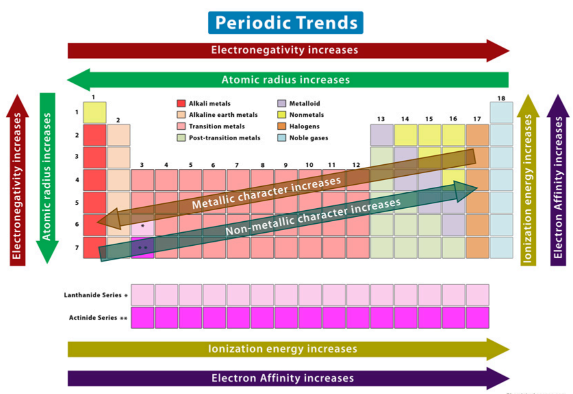
Bohr Model, Valence Electrons, Lewis Dot, Electron Configurations

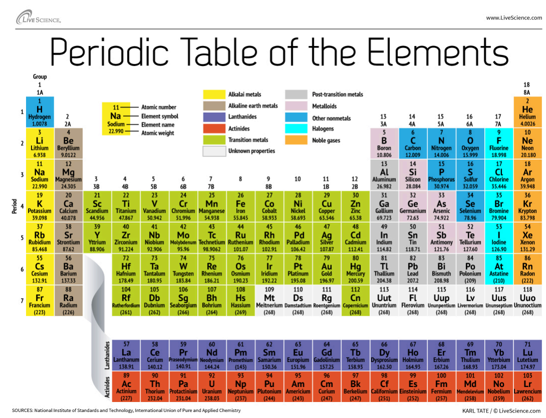
- Lewis dot diagrams - put the same number of dots as VALENCE electrons, remember He is outlier cuz it can only hold 2
- Bohr model is used to show
- Energy levels numbered 1 to 7 (columns)
- There are 4 sublevels: s-2, p-6, d-10, f-14
- Each sublevel has orbitals, s has 1, p has 3, d has 5, f has 7
- Each orbital can hold 2e- and they must have opposite spin
- Aufbau principle - e- occupy orbitals of lower energy before filling higher energy level orbitals
- Pauli Exception Principle - an orbital can only hold 2 e-, they must have opposite spins
- Hund's Rule - in a set of orbitals, the e- will fill so that each orbital gets 1 electron before any of them get a 2nd, to create as many parallel spins as possible
- Standard electron configuration - write every single sublevel out
- Shorthand electron configuration - write the noble gas before the element in brackets [], then write the rest of the configuration
- Nth shell can hold 2n^2 electrons

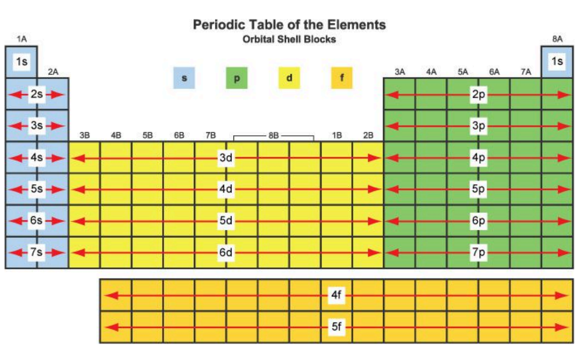
- Charges to get a full valence shell are based on number of electrons, could go up or down, losing/gaining but metals tend to lose electrons and become positive cations while nonmetals like to gain electrons and become negative anions

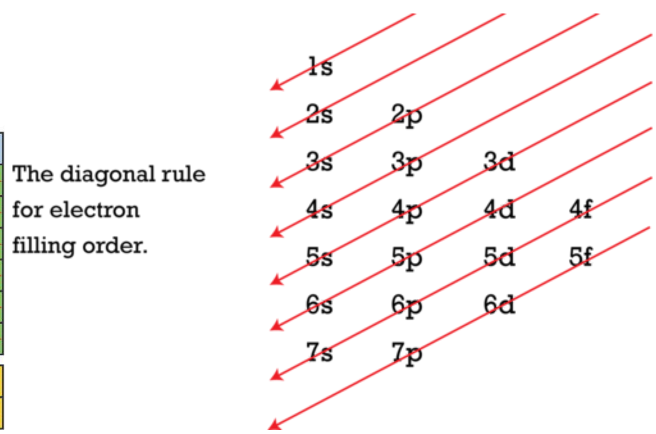
Waves and Energy0
- E = energy (J)
- c = speed of light (m/s)
- h = Planck's constant (Js)
- Upside down y = wavelength (m)
- v = frequency (Hz)
- Conversion of nanometers to meters 1 * 10^9

Valence Electrons: Ground State VS Excited State
- Electron absorbs an amount an energy that allows it to reach the next energy level, that amount of energy it called a photon
- When it goes to the next energy level, that's excited state, but it must come back to it's ground state (original energy shell) and while doing this is loses the energy (a photon)









UNIT 4: NOMENCLATURE, CHEMICAL FORMULA WRITING, AND BONDING

Ions and Binary Ionic Compounds

- Ions are atoms that either lost or gained an electron; cations - lost an electron, metal; anion - gained an electron, nonmetal
- Ti name monatomic ions, cation - it's regular name followed by "ion"; anion with the ending changes to ide; Na is sodium ion, Cl is chloride
- Ionic compounds are formed with a cation and an anion, when writing the cation comes first normally then the anion with the ending changed to ide

Ionic Compounds w/ Polyatomic Molecules

- Per..ate, ..ate, ..ite, hypo..ite

Ionic Compounds w/ Transition Metals

- Transition metals form cations but the charge is determined by a roman numeral
- Tin(II) and lead(II or IV)also need roman numerals
- Silver, zinc, and a few others do not
- Roman numerals are I, II, III, IV, V, VI, VII, VIII, IX, X

Lattice Energy

- Lattice energy is negative and exothermic(gives off heat)
- The greater the coulombic force, the stronger the lattice
- The smaller the radius, the stronger the lattice

Covalent Molecules

- Covalent bonds are between two non
- metals, they share electrons, no ions are gained
- Covalent bonds use prefixes to say the number of atoms there are

Metallic Bonding

- Metals do not form ionic bonds but they do form lattices
- Metallic bonding has the sea of electrons where all the metal atoms contribute there valence electrons to form a sea of electrons
- Because the valence electrons are free to more around, they are called delocalized electrons

Conductivity

- Conductivity increases as concentration increases for electrolytes

characteristics	**Ionic**	**Covalent**
Other names	salts	molecules
Origin of bonding	Electron transfer	electron sharing
State of matter at room temperature	solid	Solid, liquid, gas
hardness	Hard but brittle	Sofer but flexible
Electrical conductivity(electrolyte)	Good when melted/dissolved in water	poor(non-electrolytes)
Melting and boiling points	higher	lower
Forces between particles	Very strong	Fairly strong

UNIT 5: COVALENT BONDING, MOLECULAR GEOMETRY, AND ENERGY

Electronegativity & Covalent Bonds

- Nonpolar covalent bond: en difference of 0 - 0.4; Polar covalent bond: en difference of 0.5 - 1.6
- From 1.7 - 2.0 it is ionic if there is a metal otherwise it is polar covalent
- Non polar bonds have equal electron sharing while polar bonds have non equal sharing; ionic bonds don't share at all

Lewis Dot Structures

- Count the total number of valence electrons, add for negatively charged ions and subtract for positively charged
- The central atom is the atom with the lowest electronegativity but hydrogen is always a terminal atom
- Place single bonds and then place lone pairs to fill the octet rule
- If there are more put them in lone pairs on the central atom; this leads to an expanded octet for the central atom which only atoms in the 3rd period(horizontal) and lower can have
- Only C, N, O, P, S can form multiple bonds; F, Cl do not form multiple bonds
- If there is an odd number of electrons give the central atom 7 instead of 8
- Only central atoms from third pd or lower can have an expanded octet

Formal Charges & Resonance Structures

- Formal charge = # valence electrons - (# non bonding electrons + ½ # bonding electrons)
- The charges of each atom in the molecule should all add up to the net charge of the molecule
- The one that is the best is the one with the smallest charge on each individual atom
- If two resonance structures have the same charges the highest charge should go on the most electronegative element
- Find the formal charge of each, whichever has individual atoms with the lowest charge is the best

Polarity

- A molecule is polar if it has different terminal atoms or lone electron pairs around the central atom
- For checking polar bonds look at the electronegativity, the ranges will tell you if its polar or non polar

Hybridization & Sigma Phi Bonds

- Number of areas around the center atom; S -> 1, p > 3, d>5
- Single bonds are 1 sigma; Double bonds are 1 sigma and 1 phi; Triple bonds are 1 sigma and 2 phi

IMF's

- Intermolecular forces: between DIFFERENT molecules vs intramolecular forces: SAME molecule
- Ions but not polar: ionic bonding
- Ions and polar: ion-dipole
- Polar and H bonded to N, O, F: hydrogen bonding
- Polar and no H bonded to N, O, F: dipole dipole
- Not 23q]: london dispersion

Vapor Pressure and Boiling Points

- Temperature is a measure of the amount of kinetic energy that a substance has
- HEAT moves from HIGH HEAT to LOW HEAT
- When you are absorbing heat: endothermic; when you are releasing heat: exothermic

Energy/Phase Changes

- Endothermic: melting, vaporization, sublimation(solid to gas)
- Exothermic: freezing, condensation, deposition(gas to solid)
- Temperature: F = 9/5(C) + 32; K = C + 273
- Potential energy changes with phase changes; Kinetic energy changes with temperature changes

UNIT 6: CHEMICAL REACTIONS

Balancing Chemical Equations

- Count them up, same on left as on right
- If you are pairing 2 elements together, change the subscript so it makes sense

Types of Reactions; Predicting Products

- Synthesis: A + B -> AB
- Decomposition: AB -> A + B
- Single Replacement: A + BC -> AC + B; check activity series, if its above it, it can replace it
- Double Replacement: AB + CD -> AD + CB
- Acid-Base/Neutralization: HA + BOH -> H2O + BA
- Combustion: CH + O -> H2O + CO2

Oxidation States/Numbers

- Oxidation number of an uncombined element is 0
- Number of an ion is its charge
- Fluorine, -1; oxygen, -2 unless combined with F(+2) or in a peroxide(-1)
- In compounds, the elements in group 1 have +1, 2 are +2, ect
- Sum of all the oxidation numbers in a neutral compound is 0
- Sum of all the oxidation numbers in a polyatomic ion is the charge of the ion

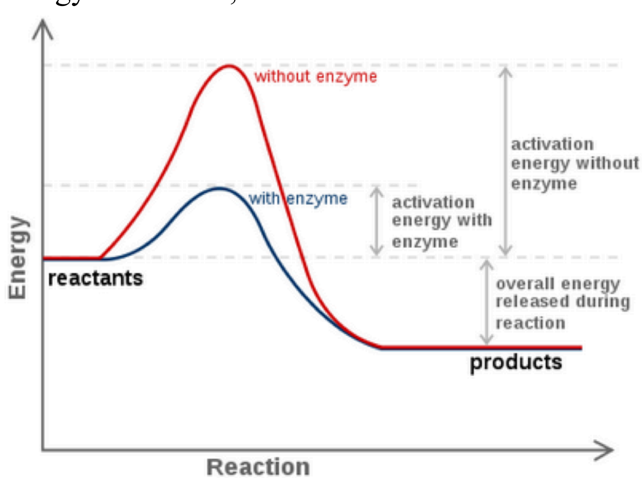
Net Ionic Equations

- Write out entire BALANCED reaction, put in the states (aq, s, l)
- Split up the aq ones, include the charges
- Cancel out the spectator ions

REDOX Reactions

- Write out oxidation numbers, if it increases its being oxidized, if it decreases its being reduced
- LEO(lose electrons, oxidation) the lion say GER(gain electrons, reduction)

Reaction Rates + Collision Theory

- Change in molarity / time
- \+ is the appearance of a product
- \- is the disappearance of a reactant
- Factors that affect: concentration, surface area, temperature, pressure
- Delta H is the heat of the reaction: its positive when energy is absorbed, endothermic; its negative when energy is released, exothermic
- 

UNIT 7: MOLAR RELATIONSHIPS

Mole + Mole Conversions

- 1 mole = 1 atom
- 1 mole/6.02 * 10^23 particles
- 1 mole/molar mass in grams(PT)
- 1 mole/22.4 Liters for a gas at standard temperature and pressure

Percent Composition

- Add up the masses of all the elements to get the molar mass
- Divide mass of one element by molar mass and multiply by 100 to get the percent composition

Empirical Formulas

- Change % to grams and convert into moles by dividing by the molar mass
- Divide all by the smallest mole numbered element
- Convert answers to whole numbers and write them as subscripts

Molecular Formulas

- Find the molar mass of the empirical formula
- Divide the given molar mass by the empirical molar mass
- Multiply all the elements by this number (multiply the subscripts)

Stoichiometry

- Use balances equation to form mole to mole ratios
- Grams A to Moles A to Moles B to Grams B (can substitute grams with particles and volume)

Limiting and Excess Reactants + Theoretical Yield

- A + B -> C
- Calculate a in grams down to c in grams and b in grams down to c in grams
- Whichever has less grams is the limiting reactant, the other is the excess reactant
- Limiting is the theoretical yield

Percent Yield + Law of Definite Proportions

- Percent yield = (actual/theoretical)*100
- No matter how much you have of a substance, if it's the same substance the element proportions will be the same

UNIT 8: SOLUTIONS AND ACID/BASE

[H] means molarity which is equal to concentration

Solutions

- Solute is dissolved in the solvent
- Molarity: M = moles/L
- M1V1 = M2V2; v2 is the whole volume nor the volume added
- Arrhenius + B-L acids you can just separate but for B-L add H2O to see full reaction

Acid-Base

- Arrhenius: Acid produces H+, Base produces OH-
- Bronsted-Lowry: Acid proton donor, Base proton acceptor
- Strong acid/base completely ionizes → but weak only partially → ←
- [H][OH] = 1 * 10^(-14)
- pH = -log[H] pOH = -log[OH]
- [H] = 10^(-pH) [OH] = 10^(-pOH)

Titrations

- [H]MaVa = [OH]MbMb

UNIT 9: KMT AND GAS LAWS

Unit Conversions

- K = C + 273
- C = (F - 32) * (5/9)\
- 1 atm = 101.325 kPa = 101325 Pa = 760 mmHg = 760 torr = 14.69 psi

Kinetic Molecular Theory

- Gas particles collide without losing energy, their movement is independent of other particles
- More pressure equals more kinetic energy

Gas Laws

- Boyle's Law: P1V1 = P2V2; pressure and volume are inversely proportional
- Charles' Law: V1/T1 = V2/T2; volume and temperature are directly proportional
- Gay-Lussac's Law: P1/T1 = P2/T2; pressure and temperature are directly proportional
- Combined Gas Law: (P1V1)/(T1) = (P2V2)/(T2)
- Avogadro's Principle: V1/n1 = V2/n2; equal number of gasses contain equal number of molecules
- Ideal Gas Law: (PV)/(Tn) = k or PV = nRT; R/n = Universal Gas Constant
- Universal Gas Constant: 0.0821 (L*atm)/(mol*K) or 8.314 (L*kPa)/(mol*K)
- Dalton's Law of Partial Pressure: P total = P1 + P2 + P3 …

Heat Calculations

- 1 Cal = 1000 cal = 1 kcal = 4184 J
- Specific heat capacity: energy required to raise temp of 1 g of substance by 1 degree, equal to a calorie
- Q = mcΔT, nrg gained/released = mass * specific heat capacity(J/(g*Cel) or (cal/(g*Celsius)* change in temperature; use when changing temperature
- Q = m(ΔHfus), nrg required = mass * heat of fusion or heat of vaporization; use when changing states
- Ice: heat of fusion = 6.01 kJ/mole or 334 J/g
- Water: heat of vaporization = 40.7 kJ/mole or 2260 J/g
- Fusion is solid to liquid; evaporation is liquid to gas

Phase Change Diagrams