

Unit 1: Atomic Structure and Properties

- Periodic table: pg 64-65
 - Group 1: Alkali Metals
 - Group 2: Alkaline Earth Metals
 - Group 3-12: Transition Metals
 - Group 17: Halogens
 - Group 18: Noble Gases
 - First row below table is Lanthanides
 - Second row beneath table are Actinides
- Avogadro's number: 6.022×10^{23} pg 66
- $PV = nRT$ pg. 67
 - N: number of moles
 - P: pressure (atm)
 - V: volume (L)
 - T: temperature (K)
 - R: gas constant, $0.0821 \text{ L atm mol}^{-1}\text{K}^{-1}$
- 1 mol gas: 22.4 L pg. 67
- Energy of electrons is quantized pg. 69
 - Only exist at specific energy levels, each specific intervals apart
- Coulomb's Law: equation for calculating amount of energy electron has a certain distance away from nucleus pg. 70
 - $E = k(+q)(-q)/r^2$
 - E: energy
 - K: Coulomb's constant
 - +q: magnitude of the positive charge (nucleus)
 - -q: magnitude of the negative charge (electron)
 - r: distance between the charges
 - Further away, the greater the energy of electron, with it being 0 at infinity
- Bohr Model: model that predicts electron orbit around fixed radius, but is not accurate Pg. 70
 - Electrons absorb electronic radiation to jump to higher energy levels and release it to fall to lower ones
- Ionization energy: the energy required for the binding energies of an electron to be broken Pg. 71
 - Caused when atom is exposed to more electromagnetic radiation than the binding energies of electrons
 - Energy is spent breaking electron free, and leftover is converted to kinetic energy for ejected electron
 - Incoming Radiation Energy = Binding Energy + Kinetic Energy
 - Requires either visible or ultraviolet range of electromagnetic radiation
- Photoelectron Spectra: amount of ionization energy for all electrons ejected from nucleus Pg. 71

- Number of electrons binded on y axis to binding energy/ionization energy on x axis
- Subshells exist within each energy level Pg. 72
 - s: holds 2 electrons or 1 orbital
 - first 2 groups + helium
 - p: holds 6 electrons or 3 orbitals
 - groups on right of table
 - d: holds 10 electrons or 5 orbitals
 - transition metals
 - f: holds 14 electrons or 7 orbitals
 - inner transition metals
 - when entering d-block energy level drops by 1
- Aufbau Principle: electrons will fill up orbitals, subshells, and shells in order of increasing energy Pg. 73
- Pauli Exclusion Principle: two electrons in same orbital can't have same spin Pg: 73
- Hund's Rule: empty orbital in subshell will always be occupied first Pg: 73
- Most stable configs have outermost energy level full Pg. 74
 - For s and p blocks that means 8 electrons
- Atoms will form ions based on the easiest way to attain an outer shell of 8 electrons Pg. 74
 - anion: negatively charged
 - cation: positively charged
 - all transition metals form cations
 - zinc always forms +2 ions and silver only forms +1 ions
 - always lose their higher-level s electrons before d electrons
- Periodic Trends Pg. 75-77
 - Formed from 3 basic rules
 - electrons are attracted to protons
 - the closer, the stronger
 - the more protons, the stronger the force
 - Electrons shield electrons in lower energy levels by repelling them
 - completed shells are very stable
 - Metals donate electrons when bonding
 - Nonmetals generally gain electrons when bonding
 - metallic character of elements decrease from left to right
 - elements in borderline between metal and nonmetal are called metalloids
 - Atomic radius is distance from nucleus to valence electrons
 - decreases from left to right
 - increases from top to bottom
 - Cations are smaller than atoms with same protons and anions are opposite
 - Ionization energy increases the more electrons are removed of a shell and greatly increases when going to a higher energy level
 - From left to right ionization energy increases
 - from top to bottom ionization energy decreases

- Electronegativity: the strength of the attraction between electrons outside of atom and nucleus
 - The smaller the atom, the greater the electronegativity
 - The closer to completion of energy level, the greater the electronegativity
 - From left to right electronegativity increases
 - From top to bottom electronegativity decreases
- Electron Affinity: describes energy change from adding electron to gaseous atom
 - for purpose of exam, follows same periodic trends as electronegativity
 - opposite if ionization energy

Unit 2: Molecular and Ionic Compound Structure and Properties

- Ionic Bond: held together by electrostatic attractions between ions in lattice structure pg. 92
 - Melting Point of Ionic Substance is affected by several factors: Pg. 92
 - charge of ions
 - greater charge causes greater bond energy so higher melting point
 - size of ions
 - inversely proportional to bond energy, so bigger ions cause lower melting point
 - poor conductors of electricity because electrons do not move around lattice
- Metallic Bonds: positively charged core of metal is stationary while valence electrons are mobile
 - delocalized electrons cause good conductivity, malleability, and ductility
 - Interstitial alloy: smaller metal atoms fill in space between larger ones
 - Substitutional alloys: metal of similar size replaces other metal
- Molecular Covalent Bonds: two atoms share electron/s as part of their outer shell Pg. 93
 - two or more covalently bound atoms are a molecule
 - molecules have definite composition
 - types of covalent bonds
 - single bonds consist of sigma bonds
 - bond order of one
 - longest bond length
 - least bond energy
 - double bonds consist of one sigma and one pi bond
 - bond order of 2
 - intermediate bond length
 - intermediate bond energy
 - triple bonds consist of one sigma and two pi bonds
 - bond order of three
 - shortest bond length
 - greatest bond energy
 - try to achieve lowest possible potential energy by being perfect distance apart

- Network Covalent Bonds: Atoms held together by lattice of covalent bonds Pg. 95
 - are very hard and have very high melting points
 - poor conductors because electrons are localized
- Ionic bonds are good conductors when in aqueous and liquid form because they are not held tightly together anymore, so the electrons can move freely Pg. 96
- Resonance Forms: different variations of the same molecule that can form Pg. 98
- molecules with d subshells can have more than 8 valence electrons but not more than 12
- Equation for calculating formal charge: valence electrons - assigned electrons
- Vsepr Theory: way to predict shape of molecule pg. 100
 - 2 electron pairs:
 - sp hybridization
 - linear shape
 - 3 electron pairs:
 - sp² hybridization
 - trigonal planar (bond angles are 120 degrees)
 - with 1 lone pair: bent
 - 4 electron pairs:
 - sp³ hybridization
 - tetrahedral (bond angles are 109.5 degrees)
 - with 1 lone pair: trigonal pyramidal
 - with 2 lone pairs: bent
 - 5 electron pairs:
 - trigonal bipyramidal
 - with 1 lone pair: folded square, seesaw, distorted tetrahedron
 - with 2 lone pairs: T-shaped
 - with 3 lone pairs: linear
 - 6 electron pairs:
 - octahedral
 - with 1 lone pair: square pyramidal
 - with 2 lone pairs: square planar

Unit 3: Intermolecular Forces and Properties

- Polar covalent bond: when electrons are unequally shared Pg. 112
 - causes dipoles, or areas of either positive or negative charge
- Nonpolar covalent: when electrons are equally shared Pg. 112
 - happens when identical atoms bond, such as diatomic molecules
- Molecular polarity is determined by molecular geometry Pg. 112
 - if it is symmetrical (made up of same atoms and there are no lone pairs) then it will be nonpolar, except for square planar molecules that are frequently nonpolar
- Lewis diagrams are drawn with least electronegative atom in center Pg. 113
 - causes central atom to have positive dipole and terminal atoms to have negative dipole, except if the terminal atom is hydrogen

- Intermolecular Forces: forces that exist between covalently bonded molecules Pg. 113
 - need to be broken for covalent substances to change phase, but bond stays intact
 - Dipole-Dipole Forces: when the positive end of a polar molecule is attracted to the negative end of another
 - increases with polarity
 - attractions are relatively weak, so melting and boiling points are low
 - because of this, often appear as liquids and gases at room temperature
 - Hydrogen Bonds: Type of dipole-dipole attraction where positively charged hydrogen is to negatively charged molecule that is extremely electronegative such as fluorine, oxygen, or nitrogen
 - much stronger than other dipole-dipole forces, causing higher melting and boiling points
 - London Dispersion Forces: occurs between all molecules and is weak attraction caused by random motions of electrons on atoms with molecules
 - At any instant, a nonpolar molecule may have more electrons on one side than the other, giving the molecule a weak dipole for an instant
 - happens more with molecules that have more electrons because it is caused by random movements of electrons
 - substances with only london dispersion forces will have higher melting and boiling points if they have more electrons, and vice versa, however since these forces are weak they will have extremely low melting and boiling points anyway and likely appear as gases in room temperature
 - melting and boiling points of covalent substances are typically lower than that of ionic ones
 - Vaporization decreases as bond strength increases
- Chromatography: the separation of mixture by passing it in solution through a median where its components move at different rates Pg. 118
 - Paper chromatography: paper is used as medium
 - Retention factor: distance traveled by solute / distance traveled by solvent front
 - Column Chromatography: a column is packed with a stationary substance (elute), then the solution (analyte) is injected into the column where it adheres to the stationary substance, then another solution (eluent) is injected into the column that passes through the stationary phase and attracts the analyte to varying degrees
 - polarity measured by how fast the analyte passes through the elute
- Distillation: Uses different boiling points to separate solutions Pg. 119
- Kinetic Molecular Theory: $KE=1/2mv^2$ Pg. 120
 - m: mass of molecule in kg
 - v: speed of molecules meters/s
 - unit is joules
 - Kinetic Energy of ideal gas is directly proportional to absolute temperature

- Average kinetic energy of gas depends only on absolute temperature, not on identity of gas
- volume of ideal gas particle is insignificant when compared with the volume in which the gas is contained
- forces of attraction don't exist between gas molecules in ideal gas
- gas molecules are in constant motion and don't lose energy
- Maxwell-Boltzmann Diagram: shows range of velocities for molecules of a gas Pg. 121
 - higher the temperature, the greater the range of velocities
- Effusion: rate at which gas will escape container from microscopic holes Pg. 122
 - dependent on speed of gas particles and molar mass? or molecule size?
 - higher speed or lower molar mass will cause greater effusion
- $PV = nRT$ pg. 123
 - N: number of moles
 - P: pressure (atm)
 - V: volume (L)
 - T: temperature (K)
 - R: gas constant, $0.0821 \text{ L atm mol}^{-1}\text{K}^{-1}$
 - $1 \text{ atm} = 760 \text{ mmHg or torr}$
- Combined Gas Law: $P_1V_1/T_1 = P_2V_2/T_2$ Pg. 123
 - can only be used when number of moles is constant
- Temperature, Pressure, Volume Relationships: Pg. 123
 - If volume is constant: As pressure increases, temperature increases, vice versa
 - If temperature is constant: As pressure increases, volume decreases, vice versa (Boyle's Law)
 - If pressure is constant: As temperature increases, volume increases, vice versa (Charles's Law)
- Dalton's Law: The total pressure of a mixture of gases is the sum of all partial pressures of individual gases in mixture Pg. 124
 - $P_{\text{total}} = P_a + P_b + P_c + \dots$
 - Partial pressure of gas is directly proportional to number of moles of gas present in mixture
 - $P_a = P_{\text{total}}(X_a)$
 - $X_a = \text{moles of gas A} / \text{total moles of gas}$
- Deviations from Ideal Behavior: Pg. 124
 - gas deviates more from ideal behavior when temperature becomes lower or pressure becomes higher (when gases become more tightly packed together)
 - When gases become packed too tightly:
 - the volume of gas molecules becomes significant
 - gas molecules attract each other and stick
- Density: $D = m / V$ Pg. 125
 - D: density
 - m: mass of gas, usually in grams
 - V: volume occupied by a gas, usually in liters
 - unit is mass per unit of volume

- Electromagnetic Spectrum: Pg. 126
 - Relationship between change in energy level of electron and electromagnetic radiation absorbed or emitted: $E = h\nu$
 - E : energy change (in Joules)
 - h : Planck's constant (6.626×10^{-34} J s)
 - ν : frequency (in s^{-1})
 - Frequency and wavelength of electromagnetic radiation are inversely proportional with higher frequencies and shorter wavelengths leading to more energy
 - $c = \lambda\nu$
 - c = speed of light (2.998×10^8 ms^{-1})
 - ν = frequency (in s^{-1})
 - λ = wavelength (in m)
 - often given in nanometers, so it must be converted to meters
 - 1000000000 or 1×10^9 nm in a m
- Beer's Law: equation used to calculate Absorbance Pg. 128
 - $A = abc$
 - A : absorbance
 - a : molar absorptivity, a constant that depends on the solution
 - b : path length, the distance the light is traveling through the solution
 - c : concentration of the solution
 - absorbance is directly related to concentration of solution
 - most effective with solutions that change color over course of reaction unless used with spectrophotometer
 - spectrophotometer: measures amount of light at given wavelength absorbed by solution

Unit 4: Chemical Reactions

- Types of Reactions Pg. 154
 - Synthesis Reactions: $A + B \rightarrow AB$
 - Decomposition: $AB + \text{heat} \rightarrow A + B$
 - Acid-Base: $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
 - Redox: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
 - Hydrocarbon Combustion: $\text{C}_4\text{H}_{10} + 13/2\text{O}_2 \rightarrow 4\text{CO}_2 + 5\text{H}_2\text{O}$
 - when covalent substance with carbon, hydrogen, and sometimes oxygen is ignited, it reacts with oxygen in atmosphere to produce CO_2 and H_2O
 - Precipitation: $\text{K}_2\text{CO}_3(\text{aq}) + \text{Mg}(\text{NO}_3)(\text{aq}) \rightarrow 2\text{KNO}_3 + \text{MgCO}_3(\text{s})$
- Important Solubility Rules: Pg. 155
 - Compounds with an alkali metal cation (Na^+ , Li^+ , K^+ , etc..) or ammonium cation (NH_4^+) are always soluble
 - Compound with nitrate (NO_3^-) anion are always soluble
- Reasons for solubility Pg. 155

- substances dissolve in water because the attractions of ions to dipoles of water molecules is stronger than attraction to each other
- Polyatomic Ions to know (as well as their charge): Pg. 156
 - Hydroxide: OH^-
 - Nitrate: NO_3^-
 - Acetate: $\text{C}_2\text{H}_3\text{O}_2^-$
 - Cyanide: CN^-
 - Permanganate: MnO_4^-
 - Carbonate: CO_3^{2-}
 - Sulfate: SO_4^{2-}
 - Dichromate: $\text{Cr}_2\text{O}_7^{2-}$
 - Phosphate: PO_4^{3-}
 - Ammonium: NH_4^+
 - Chromate: CrO_4^{2-}
 - Peroxide: O_2^{2-}
 - Oxalate: $\text{C}_2\text{O}_4^{2-}$
 - Thiosulfate: $\text{S}_2\text{O}_3^{2-}$
- Gravimetric Analysis: Used to determine the identity of an unknown sample by mixing it with another solution and seeing what is precipitated out. Pg. 160
- Rules about oxidation state (electrons gained or lost by a molecule): Pg. 161
 - Neutral unbonded atoms have oxidation state of 0
 - oxidation state = charge of ion
 - oxygen has oxidation state of -2 except for specific cases like H_2O_2 where it has charge of -1
 - when bonded to a nonmetal, H has oxidation state of +1, but -1 for a metal
 - In the absence of oxygen, the most electronegative element in a compound will take on an oxidation state equal to its most common charge
 - The combined oxidation states on all elements in a neutral compound must add up to zero. The combined oxidation states on all elements in a polyatomic ion must add up to the charge on that ion
- Things about Redox reactions: Pg. 162
 - oxidation = decrease in oxidation number
 - reduction = increase in oxidation number
 - refers to decrease in electrons which gives positive charge
 - reaction can be broken down into half reactions, one for oxidation and one for reduction
 - Redox Titration: when a known substance is slowly added to a substance of unknown concentration until it stops reacting, usually indicated by a color change, to identify the concentration of the unknown concentration
- Things about Acids and Bases: Pg. 164
 - acids are substances capable of donating protons or H^+ ions while bases are substances capable of accepting them
 - H^+ is the same as a proton because it only has 1 electron, and if you remove it to make it a cation then it has no electrons and is just a proton

- Conjugate pair: a pair of an acid and base
- Amphoteric substance: one that can act as an acid and base
 - H_2O

Unit 5: Kinetics

- Rate Law Using Initial Concentrations: Describes dependence of initial rate of a reaction on concentration of its reactions Pg. 184
 - rate of reaction: the rate of appearance of a product or the rate of disappearance of a reactant
 - Must be determined experimentally
 - Always takes this form: $\text{Rate} = k[\text{A}]^x[\text{B}]^y[\text{C}]^z$
- Rate Law Using Concentration and Time: Pg. 186
 - Zero-Order Rate Law: $\text{Rate} = k$
 - concentration doesn't affect rate
 - First-Order: $\text{Rate} = k[\text{A}]$
 - $\ln[\text{A}]_t = -kt + \ln[\text{A}]_0$
 - Second-Order: $\text{Rate} = k[\text{A}]^2$
 - $1/[\text{A}]_t = kt + 1/[\text{A}]_0$
 - $[\text{A}]_t$ = concentration of reactant A at time t
 - $[\text{A}]_0$ = initial concentration of reactant A
 - k = rate constant
 - t = time elapsed
- Half Life: the time it takes for half of a substance to react Pg. 188
 - $\ln 2/k$
- Collision Theory: chemical reactions occur because reactants are constantly moving around and colliding with one another Pg. 190
 - a reaction occurs when a certain energy threshold, or activation energy (E_a), is reached
 - referred to as effective collisions
 - Increased concentrations of gaseous or aqueous reactants will increase rate of reaction
 - reaction rate will increase for a solid if surface area is increased
 - Stirring increases reaction rate for heterogeneous mixtures
 - Heterogeneous mixture: one in which all parts of a mixture are not identical
 - Homogeneous mixture: one where all portions are identical
 - reaction rate increases with temperature
 - causes more collisions
- Reaction Mechanisms: reactions involve many intermediate steps Pg. 192
 - For reaction $2\text{A} + 2\text{B} \rightarrow \text{C} + \text{D}$ where rate law is $k[\text{A}]^2[\text{B}]$
 - $\text{A} + \text{A} \leftrightarrow \text{X}$ (fast)
 - $\text{X} + \text{B} \leftrightarrow \text{C} + \text{Y}$ (slow)
 - $\text{Y} + \text{B} \leftrightarrow \text{D}$ (fast)

- For above reaction X and Y are intermediates because they are produced and consumed within the reaction
- Elementary steps can have 1 or 2 reactants
 - 1 reactant: unimolecular
 - 2 reactants (even if they are the same): bimolecular
- Slow step is rate determining step
 - used to determine rate law for overall reaction
 - rate law of overall reaction is equivalent to rate law of slow step
 - only way to find rate law of reaction non experimentally
- Catalysts: increase rate of a chemical reaction without being consumed in the process Pg. 194
 - do not appear in balanced equation but are present in elementary steps
 - can appear in rate laws

Unit 6: Thermodynamics

- Difference between heat and temperature: pg 214
 - heat: the energy flow between two substances at different temperatures
 - temperature: the average amount of kinetic energy present in a given substance
- Exothermic vs Endothermic: Pg. 214
 - Exothermic: energy is transferred from reaction to surroundings
 - Endothermic: Energy is transferred from surroundings into reaction
- Enthalpy: a measure of the energy that is released or absorbed by the substance when bonds are broken and formed during a reaction Pg. 214
 - Unit is KJ/mol
 - Energy is released when bonds are formed
 - Energy is absorbed when bonds are broken
 - Reaction is exothermic when more energy is released when the bonds in the products form than is absorbed by the breaking of the bonds of the reactants
 - negative enthalpy change
 - Reaction is endothermic when more energy is required to break the bonds of reactants than is released by the forming of bonds of the products
 - positive enthalpy change
- Enthalpy of Formation (ΔH°): the change in energy that takes place when one mole of a compound (product) is formed from its component pure elements (reactants) under standard state conditions (temperature of 298 K or 25 C) Pg. 218
 - ΔH° of a pure element is always 0
 - ΔH° for a compound is negative when energy is released as the compound is formed from pure elements
 - exothermic
 - product is more stable than constituent elements
 - ΔH° for a compound is positive when energy is absorbed as the compound is formed from pure elements
 - endothermic

- For reaction: $aA + bB \rightleftharpoons cC + dD$
 - $K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$
 - [A], [B], [C], [D] are molar concentrations or partial pressures at equilibrium
 - Products are in the numerator, and reactants are in the denominator
 - Coefficients in the balanced equation become exponents in the equilibrium expression
 - Solids and pure liquids are not included in the equilibrium expression because they can't change their concentration. Only gaseous and aqueous species are included in the expression
 - Units are not given for K_{eq}
- Rules for manipulating equilibrium constant: Pg. 240
 - If you flip a reaction, you take the reciprocal of the equilibrium constant to get the new equilibrium constant
 - if you multiply a reaction by a coefficient, you take the equilibrium constant to that power to get the new constant
 - If you add two reactions together, you multiply the equilibrium constants of those reactions to get the new constant
- Le Chatelier's Principle: When stress is placed on a system in equilibrium, the system will shift in response Pg. 242
 - shifting right = going forwards
 - shifting left - going backwards
 - Concentration: the system will shift towards the side that allows the excess substance to be used up
 - Pressure: when pressure is increased, the system will shift towards the side with fewer gas molecules
 - Temperature: In an exothermic reaction, if the temperature is increased, the reaction will go in the reverse direction, and if it is decreased it will go in the forwards direction (opposite for endothermic)
 - basically it shifts away from the side that has heat
 - Dilution: Adding more water causes the system to shift to the side with more aqueous species
- CHANGES TO CONCENTRATION AND PRESSURE DO NOT AFFECT THE EQUILIBRIUM CONSTANT
 - They are only temporary shifts, and therefore the concentrations of products and reactants will eventually reestablish equilibrium ratio
- Shifts in temperature affects equilibrium constant
- Reaction Quotient Q: the quantitative application of Le Chatelier's Principle Pg. 244
 - determined using Law of Mass Action but can be determined using concentration or pressure values
 - Can be compared with K to predict many things
 - For reaction: $aA + bB \rightleftharpoons cC + dD$
 - $Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$

- If Q is less than K , the reaction shifts right
- If Q is greater than K , the reaction shifts left
- If $Q = K$, the reaction is at equilibrium
- Solubility Product K_{sp} : the measure of the extent of a salt's dissociation in solution Pg. 246
 - For reaction $A_aB_b(s) \leftrightarrow aA^{b+}(aq) + bB^{a-}(aq)$
 - $K_{sp} = [A^{b+}]^a[B^{a-}]^b$
- Solubility of a salt can also be described by molar solubility Pg. 246
 - Molar Solubility: the number of moles of salt that can be dissolved per liter of solution
 - molar solubility will be equal to concentration of any ion that occurs in 1:1 ratio with salt
- Common ion effect: If you add a substance that contains an ion that takes part in a reaction, it will affect equilibrium Pg. 24

Unit 8: Acids and Bases

- pH: a measurement that tells the concentration of $[H^+]$ ions in a solution Pg. 266
 - $pH = -\log [H^+]$
 - $pOH = -\log [OH^-]$
 - $pK_a = -\log K_a$
 - $pK_b = -\log K_b$
 - $pK_w = -\log K_w$
 - A solution is neutral if it has a pH of 7 ($[H^+] = [OH^-]$)
 - A solution is acidic when it has a pH of under 7 ($[H^+] > [OH^-]$)
 - A solution is basic when it has a pH of over 7 ($[H^+] < [OH^-]$)
- Strength: determined by how well an acid or base dissolves in water Pg. 267
 - Strong Acids/Bases: Ones that dissociate completely in water
 - Strong acids: HCl, HBr, HI, HNO₃, HClO₄, H₂SO₄
 - Strong bases: LiOH, NaOH, KOH, Ba(OH)₂, Sr(OH)₂
 - all strong bases end with OH
 - Weak Acids/Bases: Ones that don't dissociate completely in water
 - Conjugate Bases form when an acid donates a proton
 - Conjugate Acids form when an acid accepts a proton
 - K_a and K_b measure the strength of weak acids and bases Pg. 268
 - Acid dissociation constant (K_a):
 - $K_a = \frac{[H^+][A^-]}{[HA]}$
 - $[H^+]$: concentration of hydrogen ions (M)
 - $[A^-]$: concentration of conjugate base ions (M)
 - $[HA]$: concentration of undissociated acid molecules (M)
 - Base dissociation constant (K_b):
 - $K_b = \frac{[OH^-][HB^+]}{[B]}$
 - $[OH^-]$: concentration of hydroxide ions (M)
 - $[HB^+]$: concentration of protonated base ions (M)

- [B]: concentration of unprotonated base molecules (M)
- Ice tables: Pg. 269
- Percent Dissociation: the ratio of the concentration of dissociated hydrogen ion $[H^+]$ to the concentration of the undissociated species $[HA]$ Pg. 271
- Acid/base Structure: Pg. 272
 - not all hydrogens in acid are dissociable
 - almost all dissociable hydrogens are from hydroxyl group, due to high electronegativity difference between oxygen and hydrogen
 - most electrons are congregated around oxygen because of high electronegativity, so hydrogen ion can just 'fall off'
 - dissociable hydrogen ions of an acid are typically written in the beginning of acid and non-dissociable hydrogen ions in the middle
 - Many bases have NH_2 or amine groups
 - called nitrogenous bases???
- pH has effect on solubility because of common ion effect: Pg. 273
 - if a solution has a greater pH, there will be more OH^- and the equilibrium will be shifting away from the OH^-
 - Ex: in $Mg(OH)_2 \rightleftharpoons Mg^{2+} + 2OH^-$, adding $Mg(OH)_2$ will not change much because system is shifted toward $Mg(OH)_2$ meaning the reverse reaction is much stronger than forwards reaction, and $Mg(OH)_2$ will not dissociate
 - If a solution has a lower pH, there are more H^+ ions, which will react with OH^- in the solution, shifting system towards OH^- side and increasing solubility
- Polyprotic Acids: those that can give more than one hydrogen ion in a solution
 - each proton becomes exceedingly more difficult to donate
 - K_a value decreases
- Equilibrium constant of water K_w : $K_w = 1 \times 10^{-14}$ at 25 C Pg. 273
 - $[H^+][OH^-]$ must always equal K_w
 - K_w increases with temperature
- Neutralization Reaction: the process in which an acid donates a proton to a base when they mix Pg. 275
 - 4 different mechanics depending on strength:
 - Strong acid + strong base: both substances dissociate completely
 - only important ions are H^+ and OH^-
 - Strong acid + weak base: strong acid dissociates completely and donates a proton to the weak base, producing the conjugate acid of the weak base
 - Weak acid + strong base: strong base will accept protons from weak acid, producing conjugate base of weak acid and water
 - Weak acid + weak base: acid gives protons to base
- Buffer: a solution with a very stable pH that an acid or base can be added to without the pH being greatly affected Pg. 276
 - formed by placing large amount of weak acid or base along with conjugate base or acid into solution in form of salt
 - will not neutralize each other due to common ion effect

- any hydrogen ions added will be neutralized by base while any hydroxide ions added will be neutralized by acid
- the higher the concentration, the more effective it will be
- Henderson-Hasselbalch Equation: one that is used to calculate pH of buffer Pg. 276
 - $\text{pH} = \text{pK}_a + \log\frac{[\text{A}^-]}{[\text{HA}]}$
 - $[\text{HA}]$: molar concentration of undissociated weak acid (M)
 - $[\text{A}^-]$: molar concentration of conjugate base (M)
 - $\text{pOH} = \text{pK}_b + \log\frac{[\text{HB}^+]}{[\text{B}]}$
 - $[\text{B}]$: molar concentration of weak acid (M)
 - $[\text{HB}^+]$: molar concentration of conjugate acid (M)
 - when concentrations of acid and conjugate base or vice versa is same, $\text{pH} = \text{pK}_a$
 - When picking buffer, pick an acid that has a pK_a close to your desired pH
- Indicators: weak acids that change colors in certain pH ranges Pg. 278
 - HIn = protonated indicator
 - $\text{HIn} \leftrightarrow \text{H}^+ + \text{In}^-$
 - $\text{K}_a = \frac{[\text{H}^+][\text{In}^-]}{[\text{HIn}]}$
 - effective indicators must have protonated (HIn) state with a different color to deprotonated (In^-) state
 - indicator will change colors when HIn and In^- are same concentration, or when $\text{pH} = \text{pK}_a$
 - indicator should be one whose pK_a matches pH at equivalence point of titration
- Titration: Used for neutralization reactions when a base of known concentration is slowly added to an acid or the other way around Pg. 278
 - progress of neutralization reaction is represented by titration curve
 - equivalence point: the point in the titration when exactly enough base has been added to neutralize all the acid in solution
 - pH increases sharply just before as last of acid is neutralized
 - can be found using indicator
 - Half-equivalence point: point in reaction where enough base has been added to convert exactly half of acid into conjugate base Pg. 279
 - the center of the buffer region (area right below second increase in pH that contains equivalence point)
 - $\text{pK}_a = \text{pH}$

Unit 9: Applications of Thermodynamics

- Entropy: a measure of the disorder of a system Pg. 294
 - greater the disorder, greater entropy
 - all substances will have positive entropy because 0 entropy would be at 0 K
 - $\Delta S^\circ = \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}}$
 - solid < liquid < gas
 - aqueous solution > precipitate
 - diluting aqueous solution increases entropy

- Gibbs Free Energy: a measure of whether or not the process will proceed without input of outside energy Pg. 295

- $\Delta G^\circ = \sum G^\circ_{\text{products}} - \sum G^\circ_{\text{reactants}}$
 - if ΔG is negative, reaction is thermodynamically favorable (spontaneous)
 - If ΔG is positive, reaction is thermodynamically unfavorable (nonspontaneous)
- $\Delta G^\circ = T\Delta H^\circ - T\Delta S^\circ$
 - T = absolute temperature (K)
 - Means that systems will move to low energy and high disorder

ΔH	ΔS	$-T\Delta S$	ΔG	Spontaneity
+	-	+	+	Nonspontaneous
-	+	-	-	Spontaneous
-	-	+	+ or -	Low Temp: Spontaneous High Temp: Nonspontaneous
+	+	-	+ or -	Low Temp: Nonspontaneous High Temp: Spontaneous

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- Will always be 0 during a phase change
- $\Delta G^\circ = -RT(\ln K)$
 - R: gas constant, 8.31 J/mol•K
 - T: absolute temperature (K)
 - K: equilibrium constant
- Reduction Potential (E?): The electric potential for a half-reaction
 - always given as reduction half-reactions, but can be reversed and sign flipped to get oxidation potentials
 - The larger the positive reduction potential, the more likely it is to gain electrons??
 - $E(\text{whole reaction}) = E_{\text{oxidation}} + E_{\text{reduction}}$
 - never multiply potentials by coefficients
- Galvanic Cell: A favored redox reaction used to generate a flow of current Pg. 299
 - two half-reactions occur in separate chambers, where the electrons from the oxidation reaction pass through a wire to the chamber of the reduction reaction where they are consumed
 - Current: the flow of electrons from one place to another
 - anode: where oxidation takes place
 - an(ode) ox(idation)
 - cathode: where reduction takes place
 - red(uction) cat(hode)
 - salt bridge is used to maintain electrical neutrality in cell
 - positive cations from salt bridge solution flow into cathode while negative anions from salt bridge solution flow into anode
 - voltage of cell is same as total voltage of redox reaction at standard conditions (all concentrations are 1 mol)
 - equilibrium constants are always greater than 1, with reaction quotients being small, making reaction spontaneous?
 - changes in temperature, pressure, and concentration affect reaction quotient and therefore voltage

- $E_{\text{cell}} = E^{\circ}_{\text{cell}} - (RT/nF)(\ln Q)$
 - won't be required to know, though can be used to mathematically justify change in cell potential
- E is always positive
- Electrolytic Cells: An unfavored redox reaction forced by an outside source of voltage (Pg. 301)
 - mostly occur in aqueous solutions where either ions or water molecules are reduced or oxidized
 - anodes and cathodes are typically just metal bars that conduct current but don't take part in reaction
 - E is always negative
- Electroplating: electrolytic cells are used to produce a metal coating on a solid substrate Pg. 302
 - Steps to figure out electrolysis problem:
 - use equation $I = q/t$ to find charge in coulombs if given current and time
 - I = current (amperes, A)
 - q = charge (coulombs, C)
 - t = time (seconds, s)
 - 2. use equations Mols Electrons = coulombs/96,500 coulombs/mol to convert coulombs to electrons
 - 3. Calculate amount of mols of metal plated using number of mols of electrons.
 - 4. use stoichiometry to calculate number of grams of metal
- Redox reactions will be favored if potential is positive Pg. 303
 - $\Delta G^{\circ} = -nFE^{\circ}$
 - ΔG° : Standard Gibbs free energy change (J/mol)
 - n: number of mols of electrons exchanged in reaction
 - F: Faraday's constant, 96,500 coulombs/mol (1 mol of electrons has charge of 96,500 coulombs)
 - E° : Standard reaction potential (unit is V, which is equivalent to J/C)
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