

AP Chemistry Course Outline - This outline follows the AP outline as found on College Board website and provides details of content from first year chemistry. This year additional depth and concepts will be added.

I. Structure of Matter (20%)

A. Atomic Theory

1. Evidence for Atomic Theory

- a. Democritus- Greek philosopher proposes idea of smallest piece of matter- coined term atom
- b. **Dalton**- solid sphere model of atom
 - i. compiled experimental data available at early 1800s to propose Atomic Theory
 - ii. proposes all matter made of small invisible indivisible atoms
 - iii. atoms neither created nor destroyed from Lavoisier evidence of conservation of mass
 - iv. atoms combine in simple whole number ratios from the constant ratios of combination from Proust
- c. **Thomson**- plum pudding model
 - i. cathode ray in evacuated tube- bends towards pos plate and away from neg plate
 - ii. proposed behavior consistent with beam of neg chg particles
 - iii. discovers electron
 - iv. Millikan oil drop exp
 - determines charge/mass ratio- indicates e- must be much smaller than atom
 - v. Thomson model of atom - plum pudding model
 - neg e- embedded in positive matrix
- d. **Rutherford**- nuclear model
 - i. Goldstein studies positive rays of modified cathode ray tube
 - positive ions emitted – H⁺ ion (H nuclei) identified as proton
 - positive ion emitted – He²⁺ ion (He nuclei) identified as alpha particle α
 - ii. Gold Foil experiment
 - alpha particles shot at thin sheet of gold foil
 - unexpectedly some of the alpha particles were deflected backwards
 - most go straight through
 - iii. could only be explained if
 - most of atom's mass and positive protons concentrated in small nucleus
 - nucleus must be very small compared to size of atom since most go through
 - calculated diameter of nucleus- mostly empty space
 - iv. Rutherford atom
 - e- moving around small but massive nucleus
 - since chg of p = to oppo of chg on e-, and atom neutral #p = #e-
 - Mosely x-rays proof of atomic number concept for diff atoms
 - Chawick discovers neutron
 - e. Planck, Einstein and Bohr (details later)
 - f. de Broglie, Heisenberg and Shrödinger (details later)

2. Atomic Masses determination by chemical and physical means

- a. Weighted atomic mass averages
 - i. average considers the relative abundance of different isotopes
 - ii. ex gallium
Ga-69 mass of 68.9257 x 60.27% abundance/100 = 41.54
Ga-71 mass of 70.9249 x 39.73% abundance/100 = 28.18
The weighted average is found as the sum Sum 69.72

b. Mass spectrometer

- i. ionization chamber- ions of diff isotopes deflection depends upon mass
- ii. can be directed towards strike detector measures ion current
- iii. resulting plot- mass spectrum shows mass number compared to relative intensity
- iv. indicates relative abundance of diff isotopes

3. Atomic Number and mass number: isotopes

- a. Atomic number = # protons
- b. Mass number = # protons + # neutrons for an isotope
- c. isotopes atoms with same atomic #but diff mass #

- d. Carbon-12 and carbon-14
- Both have 6 protons
 - C-12 with mass # of 12 has only 6 neutrons
 - C-14 with mass # of 14 has 8 neutrons [radioactive isotope- unstable nucleus]
- e. U-238 another radioactive isotope #92 so 92 protons and 146 neutrons
4. Electron Energy Levels: atomic spectra, quantum numbers and atomic orbitals
- a. electromagnetic waves
- travel at the speed of light (c) 2.9979×10^8 m/s
 - wavelength x frequency = speed of light ($\lambda f = c$)
 - ↑ wavelength ↓ frequency
 - ↑ frequency ↑ energy
 - spectrum from long waves (low energy & low frequency) to short waves
radio waves --- infrared -- ROYGBIV -- ultraviolet --- x-rays --- gamma rays
- b. spectroscope separates incoming wavelengths to spectrum based upon wavelengths
- c. wave model vs particle model
- interference patterns and diffraction only explained by wave model
 - Max Planck analyzed radiation emitted by hot glowing objects
 - intensity of light varies in manner not explained by wave model
 - proposed radiant energy emitted not in continuous characteristic of waves but rather in tiny bundles of or packets called quanta
 - Albert Einstein uses Planck's proposition to explain photoelectric effect
 - that light consists of particle-like quanta called photons
 - e- dislodged from the electrode surface if photons have sufficient or threshold energy
 - ex. Bright red light did not dislodge e but dim blue does
light is not continuous- a single photon must have sufficient energy
red photon does not have enough energy
 - Energy carried by a single photon proportional to energy
 $E = h f$ Energy = Planck's constant x frequency
- d. **Bohr and the spectrum of hydrogen**
- Balmer and Rydberg independently
 - work out mathematical equation to describe the spectral lines
 - emission spectrum of hydrogen only certain wavelengths observed
 - only certain energy states possible for atoms and emitted photons only have energies corresponding to the difference between these states
 - no model to explain only mathematical relationship
 - Bohr
 - proposes that e- move around nucleus in only certain allowable orbits
 - increasing energy as increase distance from nucleus assigned integral #n = 1, 2, 3, ...
 - derives an equation using some assumptions and classic physics to propose model
 $E_n = -2.179 \times 10^{-18} \text{ J/n}^2$
 - integer n is called a quantum number
 - ground state e- in most strongly bonded state closest to nucleus
 - excited state e- in higher energy level
 - photon of energy given off or absorbed each time e- moves from one level to another
 $E_{\text{photon}} = E_f - E_i = \Delta E$
If e- moves up to higher level energy absorbed calculated ΔE is POS
If e- falls down to lower level energy emitted calculated ΔE is NEG
 - Limitations of Bohr
 - path of e- has proven to more elusive- not in well-defined orbits
 - not consistent with atoms with multiple e- or later discoveries
- e. emergence of quantum model
- de Broglie- matter waves
 - $\lambda = h/mv = \text{Planck's constant/mass x speed}$
 - with large mass wavelength is very very small

- ii. son of JJ Thomson demonstrates wave behavior of e-
 - Max Born proposes e- waves related to probability of finding e- at particular point
- iii. Heisenberg – uncertainty principle
 - impossible to make simultaneous and exact measurements of both position & momentum of e-
 - inexactness - predict only the most probable location of e and its behavior

f. wave mechanics (quantum mechanics) Schrödinger

- i. recognizes as Bohr did that only certain allowable states
- ii. each state described by a wave function- only a calculation of probability
- iii. e- density map corresponds to orbital region of space

g. e- orbitals and quantum numbers

- i. principal quantum number (**n**)
 - designates the energy level related to distance from nucleus
 - shell all orbitals with same principal quantum number
- ii. angular momentum quantum number (also azimuthal) (**ℓ**)
 - determines shape of orbital
 - values possible from 0 to n-1 for that shell
 - ℓ = 0 is the spherical shape of s orbital
 - ℓ = 1 is the double lobe shape of p orbital
 - ℓ = 2 is the four lobe shape of the d orbital
- iii. magnetic quantum number (**m_ℓ**)
 - determines the orientation of orbital
 - possible values from -ℓ... 0... +ℓ
 - for ℓ = 0 s-orbital only possible value is 0 so only one s-orbital per shell
 - for ℓ = 1 p-orbital only possible values is -1, 0, +1 so only three p-orbitals per shell
 - for ℓ = 2 d-orbital only possible values is -2, -1, 0, +1, +2 so only five d-orbitals per shell
- iv. magnetic spin number (**m_s**)
 - line spectrum changes in magnetic field
 - e- behaves as a tiny magnet- moving charges generate magnetic fields
 - only two possible ways- spin states
 - equal but opposite spins possible values either +½ or -½

v. orbitals (n, ℓ values)

1s (1, 0)			
2s (1, 0) 2p (2,1)			
3s (1, 0) 3p (3,1)	3d (3,2)		
4s (1, 0) 4p (3,1)	4d (4,2)	4f (4, 3)	

- vi. draw e- config - show box for each orbital -
 - Pauli exclusion principle- only two e- can occupy an orbital
 - Aufbau fill in order of lowest sum of first two quantum #
 - Hund's rule a single e- in each orbital of subshell before doubling

5. Periodic Relationships including atomic radii, ionization energies, electron affinities, oxidation states

a. Atomic Radii- size of atom

- i. increase as move down a column since
 - another energy level is added - shielding effect of another layer of neg e- between
 - increasing n – increasing distance from nucleus
- ii. decrease as move across a row since
 - electrons are in same level about the same distance from nucleus
 - but another protons pulls in so greater electrical force pulling in to make smaller

b. Ionization Energy- energy required to remove electron

- i. decrease as move down a column since
 - last electron farther from nucleus- since another level out
 - so less electrical force pulling in- takes less energy to remove
- ii. decrease as move across a row since
 - electrons are in same level about the same distance from nucleus
 - but another protons pulls in so greater electrical force pulls in to make smaller

c. Electron Affinity (E released when e⁻ added)

follows same trend as electronegativity (pull of atom on electrons)

- i. decrease as move down a column since
 - another energy level is added
 - last electron farther away so less electrical force
- ii. increase as move across a row since
 - electrons are in same level (same distance) but more protons pull in

B. Chemical Bonding

1. Binding Forces

a. Types: ionic, covalent, metallic, hydrogen bonding, van der Waals (including London dispersion forces)

i. IONIC BONDS

a. difference in electronegativity of more than 1.9

significance difference between the two of their pull on the electrons

b. electrostatic attraction between oppositely charged ions

electrically force depends

- directly on charge $F = k(q_+q_-)/d^2$
- inversely (squared) on distance

lattice energy- energy released when ions in gaseous state for ionic crystal

- increases with increasing charge
- decreases with increasing size ions – ions cannot get as close

solids at room temp due to ionic bond

- solid does not conduct electricity since charged ions are NOT free to move
- molten conducts electricity since charged ions are free to move

c. Dissolve in solution

ions dissociate (split up) and form ion-dipole attractions with water

solution conducts electricity since the charged ions are free to move

ion-dipole attraction-

- charged ions attracted to oppositely charged ends of water molecule

degree of solubility depends on strength of ionic bonds

- saturated solution forms when equilibrium reached- maximum dissolved
- K_{sp} calculated indicates degree of solubility
- see solubility rules

ii. COVALENT BONDS

a. sharing of pair(s) of electrons between two atoms

b. sigma bond- single bond – shared pair of e⁻ located primarily directly between two nuclei

minimize repulsions- neg e⁻ between two positive nuclei reduces repulsion

maximize attractions

- neg e⁻ in closest location to both nuclei
- increases attraction of e⁻ to both nuclei

c. pi bonds- second and third bond forms between some pairs of atoms

d. difference in electronegativity (EN) determines polarity of bond

EN diff less than 0.5 - nonpolar covalent bond forms

- equal or almost equal sharing of electrons
- e⁻ spend almost the same time around each atom
- leads to no oppositely charged ends
- induced dipole-induced dipole attractions form between nonpolar molecules

EN diff between 0.5 – 1.9

- unequal sharing of electrons
- e⁻ pair spends more time with one atom than the other (greater e⁻ density)
- leads to neg and pos chg ends
- dipole-dipole attractions between polar molecules

e. bond length- equilibrium distance between the centers of two bonded atoms

f. bond dissociation energy- energy required to break bond

- bond breaking is endothermic (E_{in})

- bond forming is exothermic (E_{out})
- Smaller atoms tend to form stronger and shorter bonds
- Multiple bonds tend to form stronger and shorter bonds

iii. **METALLIC BONDS**

- loose free moving sea of electrons shared between all the atoms
- allow metals to conduct electricity since e^- have charge and can move

iv. **INTERMOLECULAR FORCES (IMFs)** - van der Waals forces

- dipole-dipole attractions**
 - attractions between oppos chg poles of polar molecules
 - increases with polarity of molecules
 - much weaker than ionic bond- since only partial chgs rather than full chg
- hydrogen bond**- strongest IMF
 - hydrogen atom bonded to small highly electronegative atom (N, O, F)
 - very polar bond results
- induced dipole-induced dipole attractions** (London dispersion forces)
 - even nonpolar atom or molecule has polarity that varies from instant to instant
 - fluctuations lead to induced dipoles- only temporary
 - larger atoms or molecules are more easily polarized
 - higher BP and MP for larger atoms or molecules if similar attractions

b. Relationships to states, structure and properties of matter

c. Polarity of Bonds, electronegativities

2. Molecular Models

a. **Lewis structures**

- count # valence electrons (use column on periodic table)
- draw skeleton structure- generally first atom listed is central atom
- distribute the valence e^- to give each atom stable octet (4 pr of e^-)
- insert double and triple bonds if not enough e^- to give every atom stable octet
- expanded octet- atoms with available d orbitals in row 3 or lower
- reduced octet- always H w/ only 1 pr. typically Be w/ only 2 pr and B w/ only 3 pr
- formal charge- compare # e^- in neutral to # assigned in Lewis structure
 - Increased stability with reduced formal charges
 - Resonance structures can be drawn for the same group of atoms
- sigma bond- a single bond forms between two atoms
 - e^- density between two nuclei
 - minimizes repulsion with neg e^- between pos nuclei
 - maximizes attraction with neg e^- smallest distance to each pos nuclei
- pi bonds- double and triple bonds
 - formed by the overlap of p-orbitals above and below the plane or front and back
 - weaker than sigma bond

b. Valence Bond: hybridization of orbitals, resonance, sigma bonds, pi bonds

- count # bonding pairs (count single, double or triple as single BP) & # nonbonding prs around atom
- sum of BP and NBP = # electron pairs (EP)
- Each EP requires an orbital and repels other EP
- EP will push each other as far apart as possible
- atomic orbitals are not suitable for bonding-
 - e^- density between two nuclei stronger bond
 - think p-orbital EP spends $\frac{1}{2}$ time on side oppos the location between the two atoms
 - atomic orbitals hybridize (combine) to form hybrid orbitals for bonding

c. **VSEPR- Valence Shell Electron Pair Repulsions**

- EP repulsions push EP as far apart as possible
- Use count of EP to determine the shape

3. ex. w/ only Two EP push apart 180° for linear arrangement of orbitals- 2 hybrid orbitals required
4. ex. w/ only Three EP push apart 120° for trigonal planar arrangement
5. # EP determines EP geometry (where are the EP) and orbital hybridization
Molecular geometry- where are the atoms – use # BP and NBP
6. shape and orbital hybridization

2 EP	linear	2 sp hybrid orbitals	180°
3EP	trigonal planar	3 sp^2 hybrid orbitals	120°
2 BP	bent		
4 EP	tetrahedral	4 sp^3 hybrid orbitals	109°
3 BP	trigonal pyramidal		
2 BP	bent		
5EP	trigonal bipyramidal	5 sp^3d hybrid orbitals	90° and 120°
4 BP	seesaw		
3 BP	T-shape		
2 BP	linear		
6 EP	octahedral	6 sp^3d^2 hybrid orbitals	90°
5 BP	square pyramidal		
4 BP	square planar		

3. Geometry of molecules and ions, structural isomerism of simple organic molecules and coordination complexes; dipole moments of molecules; relations of properties to structure
4. Nuclear Chemistry: nuclear reactions, half-lives, and radioactivity; chemical applications
 - a. Becquerel- accidentally discovered exposure of photographic plates with a uranium salt
 - b. Marie Curie discovers new radioactive element Polonium and Radium
 - c.
 - α alpha particle helium nuclei --- least penetrating
 - β high spin electron
 - γ gamma ray- no mass at all ---- most penetrating
 - d. nuclear reaction- nucleus changes
 - e. radioactive decay - unstable nucleus releases radioactive particles
 - α decay- nucleus emits helium nuclei mass # decreases by 4 atomic # decreases by 2

$${}_{92}^{238}\text{U} \rightarrow {}_2^4\text{He} + {}_{90}^{234}\text{Th}$$
 - β decay a neutron converted to proton releasing an e^- atomic # increases by 1 no Δ in mass #

$${}_{90}^{234}\text{Th} \rightarrow {}_{-1}^0e + {}_{91}^{234}\text{Pa}$$
 - e- capture inner e^- pulled in proton converted to neutron atomic # decreases by 1
 - f. half-life - time required for $\frac{1}{2}$ the sample to become stable by decay
 5730 years for carbon-14
 If start with 50 grams, then after one half-life 25 grams remain
 After two half-lives 12.5 grams remain

II. States of Matter (20%)– including particles model for solids, liquids, and gases

A. Gases

1. Pressure caused by collision of particles with unit square (force/area)
 - a. measured with barometer- height of mercury column held up (mmHg)
 - b. also measured in atm (atmospheres 1 atm = 760 mmHg)
2. Law of Ideal Gases
 - a. Equation of state for an ideal gas $PV = nRT$ R= ideal gas constant
 - b. Partial pressures
 - Total pressure equals sum of partial pressures of each gas $P_T = P_1 + P_2 + P_3\dots$
 - mole fraction times total pressure equals the pressure of the individual gas
3. Kinetic Molecular Theory
 - particles small compared to space between them
 - particles in constant random motion
 - particles assumed to exert no force on each other- no attractions or repulsions
 - average KE of particles assumed to be directly proportional to the Kelvin temperature

- a. Interpretation of ideal gas laws on the basis of this theory
- Charles' Law $\uparrow T \rightarrow \uparrow V$
 $\uparrow T \uparrow KE \uparrow P$ speed of p
 \uparrow # collisions push out the sides to expand keeps P constant
 - Boyle's Law $\downarrow V \rightarrow \uparrow P$
 $\downarrow V \rightarrow$ same # p in less space \uparrow collisions $\uparrow P$ since cannot push out sides
 - Avogadro's Law $\uparrow n \rightarrow \uparrow P$
 \uparrow # particles in same volume $\rightarrow \uparrow$ collisions $\uparrow P$
 - Gay-Lussac Law $\uparrow T \rightarrow \uparrow P$
 $\uparrow T \uparrow KE \uparrow P$ speed of p
 \uparrow # collisions cannot push out sides so $\uparrow T \uparrow KE \uparrow P$
- b. Avogadro's hypothesis and the mole concept
 Equal volume of gas at same temperature and pressure contains same # molecules
 Initial mole concept
- c. Dependence of kinetic energy of molecules on temperature
 Temperature – average Kinetic energy of particles
- d. Deviations from ideal gas laws
 real gases exert have at least some attractions between particles
 exert a significant impact at very low temps
 real gases are not volumeless spheres
 exert a significant impact at very high pressures-
 particles push very close- cannot ignore radii compared to space between

B. Liquids and Solids

1. Liquids and solids from kinetic molecular viewpoint
2. Phase diagrams of one component systems
3. Changes of state, including critical points and triple points
4. Structure of solids; lattice energy

C. Solutions

1. Types of solutions and factors affecting solubility
2. Methods of expressing concentrations
 - a. molarity = moles solute / Liter of solution
 - b. molality = moles of solute / kg of solvent
 - c. percent by mass = mass of solute / mass of solution x 100
 - d. mole fraction = moles of solute / total # molecules
3. Raoult's Law and colligative properties
4. Nonideal behavior (qualitative aspects)

III. Reactions (35-40%)

A. Reaction Types

1. Acid-base reactions; concepts of Arrhenius, Bronsted-Lowry and Lewis, coordination complexes, amphoterism

a. Arrhenius model - Double Replacement Reaction

- i. H⁺ ion from the acid
- ii. OH⁻ ion from base
- iii. H⁺ from acid combines with OH⁻ from base combine to form water molecules
- iv. ex. $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O(l)}$
- v. ex. $\text{H}_2\text{SO}_4(\text{aq}) + \text{Ba(OH)}_2(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{H}_2\text{O(l)}$

b. Brosted-Lowry proton (H⁺) transfer

- i. must have H to donate protons
 must have lone or nonbonding pair of electrons to accept H⁺ by forming covalent bond with it
- ii. $\text{NH}_3(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
 BL base BL acid conjugate BL acid conjugate BL base
 Accepts proton donates proton can donate H⁺ can accept H⁺
- iii. $\text{H}_2\text{O(l)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$

	BL base		BL acid		conjugate BL acid		conjugate BL base
iv.	stronger the acid the weaker the conjugate base						
v.	$\text{H}_3\text{O}^+(\text{aq})$		strongest acid that exists in water.		HCl		strong acid completely dissociates in water
	$\text{H}_2\text{O}(\text{l})$	+	$\text{HCl}(\text{aq})$	\rightarrow	$\text{H}_3\text{O}^+(\text{aq})$	+	$\text{Cl}^-(\text{aq})$
	BL base		BL acid		conjugate BL acid		conjugate BL base
vi.	$\text{H}_2\text{O}(\text{l})$	+	$\text{HCN}(\text{aq})$	\rightleftharpoons	$\text{H}_3\text{O}^+(\text{aq})$	+	$\text{CN}^-(\text{aq})$
	BL base		BL acid		conjugate BL acid		conjugate BL base

c. Lewis model- generally not covered in first year

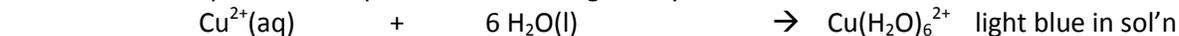
i. Lewis base donates e-pair

ii. Lewis acid accepts e- pair



has empty orbital w/ reduced octet has lone pair of e-

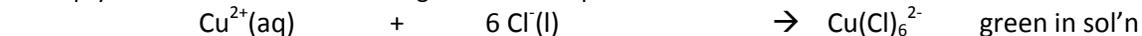
iv. coordination complexes – complex ions with charges may remain in solution



As empty orbitals available for bonding has lone pair of e- to donate to form a bond



As empty orbitals available for bonding has lone pair of e- to donate to form a bond

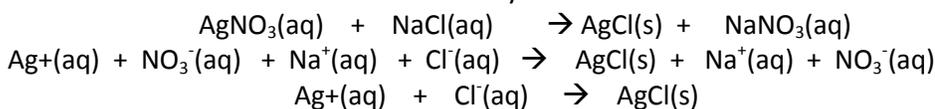


As empty orbitals available for bonding has lone pair of e- to donate to form a bond

2. Precipitation Reactions - Double Replacement reactions that form insoluble solid- see solubility rules

a. strong ionic bonds form between pair of ions and form ppt- saturated solution

b. ex. Remember IA ions and nitrate ions always soluble



Solubility rules that apply to water solution:

- (1) All alkali metal (lithium, sodium, potassium, rubidium, and cesium) and ammonium compounds are soluble.
- (2) All acetate, perchlorate, chlorate, and nitrate compounds are soluble.
- (3) Silver, lead, and mercury(I) compounds are insoluble.
- (4) Chlorides, bromides, and iodides are soluble
- (5) Carbonates, hydroxides, oxides, phosphates, silicates, and sulfides are insoluble.
- (6) Sulfates are soluble except for calcium and barium.

3. Oxidation-Reduction reactions

a. oxidation number- assigned to keep track of e- and identify the species oxidized and reduced

- element in uncombined state oxid # = 0
- monoatomic ions oxid # = charge on the ion
- oxygen in compounds oxid # = -2 except in peroxide the oxid # = -1
- hydrogen in compounds oxid # = +1 except in metal hydrides the oxid # = -1
- sum of oxid # in neutral compound = 0
- sum of oxid # in polyatomic ion = charge on ion

b. role of electron in redox reactions LEO the lion goes GER

- e- transfer in redox reactions
- species oxidized loses e-
- species reduced gains e-

c. electrochemistry: electrolytic & galvanic cells, Faraday's Laws, std half-cell potentials; Nernst equation; prediction of direction of redox rxns

- spontaneous redox reaction may occur in solution with direct transfer of e- between
If separated into two half-cells
the e- transferred by traveling through a wire generating electrical voltage
oxidation occurs at anode – e- travels through wire – cathode where reduction occurs

B. Stoichiometry

1. Ionic and molecular species present in chemical systems: net ionic equations
2. Balancing equations including those for redox reactions
3. Mass and volume relations with emphasis on the mole concept
including empirical formulas and limiting reactants

C. Equilibrium

1. Concept of dynamic equilibrium, physical and chemical; Le Chatelier's principle, equilibrium constants

a. at equilibrium

- rate of forward reaction is equal to the rate of reverse reaction
- conditions remain constant at equilibrium
- constant conditions allow the calculation of equilibrium constant
- $K_{eq} = \frac{[\text{reactant}]}{[\text{products}]}$ raised to exponents corresponding to coefficients in equation

b. vapor pressure equilibrium $l \rightleftharpoons g$

- in closed container
- molecules of liquid escape and move into gas
- at same rate as molecules in gas move back into liquid
- VP measure of collisions of the molecules of the liquid escaped into gas
- As temp increases – more molecules escape liquid – increases VP
- when VP of liquid equals outside air pressure liquid will boil

c. ionic compounds in saturated solution $(s) \rightleftharpoons (aq)$

- $\text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq)$ K_{eq} or $K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$
- at equilibrium ions dissociate at rate equal to rate of crystallization
- may be disturbed

by adding an acid H^+ react with CO_3^{2-} producing CO_2 that leaves as gas

- $\downarrow [\text{CO}_3^{2-}]$ leads to less \downarrow collisions between Ca^{2+} and CO_3^{2-} ions thus rate of reverse rxn
Rate of forward reaction temporarily faster than rate of reverse- shift to RIGHT
- Le Chatelier predicts shift- removing product system responds to replace product

d. $Q =$ reaction quotient using same form ($[\text{P}]/[\text{R}]$) as equilibrium constant

- can be calculated to predict equilibrium shift
- if $Q = K_{eq}$ than system is at equilibrium
- If $Q > K_{eq}$ than excess product shift to left
- If $Q < K_{eq}$ than excess reactant shift to right

2. Quantitative treatment (limited coverage in first year)

a. Equilibrium constants for gaseous reactions: K_p and K_c

b. Equilibrium constants for aqueous solutions

i. constants for acids and bases; pK and pH

ii. solubility product constants and their application to precipitation and dissolution of slightly soluble compounds

iii. common ion effect; buffers and hydrolysis

D. Kinetics

1. Concept of rate of reaction

a. rate of reaction depends upon the # effective collisions/second

b. effective collisions- occur with enough KE (activation energy) at right orientation to break bonds

c. factors affecting rate of reaction

- \uparrow concentration \uparrow # p in same volume $\rightarrow \uparrow$ # collisions
- \uparrow surface area (more p exposed on surface) \uparrow # collisions
- \uparrow temp \uparrow average KE \uparrow # collisions
- agitation moves particles around \uparrow collisions of reacting particles

2. Use of experimental data and graphical analysis to determine reactant order, rate constants, and reaction rate laws (not covered in first year)

3. Effect of temperature change on rates- above

4. Energy of activation; role of catalysts

a. activation energy- energy required for an effective collision the uphill part of energy hill diagram

b. catalyst lowers the activation energy required

c. a greater fraction of the collisions are effective therefore faster rate

5. The relationship between rate-determining step and reaction mechanism (not covered in first year)

E. Thermodynamics

1. State functions- value of the change is independent of pathway-

same value regardless of whether occurs in one step or 4 small steps

2. First Law: change in enthalpy, heat of formation; heat of reaction; Hess's Law, heat of vaporization and fusion; calorimetry

- a. ΔH_f heat of formation change in enthalpy associated with formation from elements in standard state
 - b. $\Delta H = \sum \Delta H_f \text{ products} - \sum \Delta H_f \text{ reactants}$
 - c. endothermic – energy transferred out of system to surroundings
 - positive ΔH - increase in enthalpy
 - products with greater enthalpy are less stable than reactants
 - NOT favored based upon enthalpy
 - d. exothermic– energy transferred into system from surroundings
 - negative ΔH - decrease in enthalpy
 - products with less enthalpy are more stable than reactants
 - favored based upon enthalpy
3. Second Law: entropy; free energy of formation, free energy of reaction; dependence of free energy on enthalpy and entropy changes
- a. entropy – disorder
 - $\Delta S = \sum S \text{ products} - \sum S \text{ reactants}$
 - increase in entropy is favored- tendency to become more disordered (pos ΔS)
 - decrease in entropy is not favored
 - entropy is about probability and randomness
 - b. free energy $\Delta G = \Delta H - T\Delta S$ T = temperature must be in Kelvins
 - neg ΔG reaction is spontaneous – will occur – but no info on rate
4. Relationship of change in free energy to equilibrium constants and electrode potentials (not covered in first year)

IV. Descriptive Chemistry (10-15%)

- A. Knowledge of specific facts of chem. essential for understanding
- B. Chemical reactivity
- C. Relationships on the periodic table
- D. Intro to Organic chem.

V. Labwork - Major Labs from First Year Chem (10-15%)

- A. Determining the Heat of Fusion of Ice – Ice was melted in a calorimeter to calculate the Joules absorbed by the melting ice.
- B. Determining the K_{sp} of Sodium Chloride – Saturated solutions of sodium chloride and other salts were prepared. Different volumes were heated to dryness to remove the water and find the concentration of the solution.
- C. Comparing the Reactivity of Metals in Acid – Metals (magnesium, copper, zinc, iron, aluminum) were added to hydrochloric acid to compare their reactivity and construct an activity series in single replacement reactions.
- D. Comparing the Solubility of Ionic Compounds – Combinations of ionic solutions were mixed and the resulting mixes lead to the models of double replacement reactions and differential solubility.
- E. Determining the Empirical Formula of Magnesium Chloride – Hydrochloric acid was added to magnesium ribbon until half of it had reacted. The resulting magnesium chloride was isolated and massed to calculate the mole ratio of magnesium to chlorine.
- F. Determining the Empirical Formula of a Hydrate – Blue copper sulfate hydrate was heated driving off the waters of hydration to determine the empirical formula of the hydrate
- G. Stoichiometry of a single replacement reaction – A copper wire is allowed to react with silver nitrate. The silver product is isolated and percent yield is determined.
- H. Emission Spectrum – The emission spectra of six elements are observed. The behavior of electrons are used to explain the observed colors.
- I. Periodicity – Reactions of alkali metal and alkali earth metals are observed. Patterns on the periodic table are constructed to explain the differences in reactivity.
- J. Titration of acids and bases – The volume of acids and bases mixed to reach neutrality are used to calculate the moles or molarity of one of the solutions.
- K. Collection of Hydrogen Gas from the reaction of magnesium and hydrochloric acid –A small ribbon of magnesium is placed within a copper cage in a long gas collecting tube. The hydrogen gas collects at the top of the tube and displaces the solution. The volume of the gas and ideal gas law are used to calculate the moles of the gas.

VI. Chem Calculations

- A. percentage composition
- B. Empirical and molecular formulas from experimental data
- C. molar masses from gas density, freezing point and boiling point measurements
- D. Gas laws, including ideal gas law, Dalton's Law and Graham's Law
- E. Stoichiometric relations using the concept of mole; titration calculations
- F. Mole fractions; molar and molal solutions
- G. Equilibrium constants and their applications, including their use for simultaneous equilibria
- H. Standard electrode potentials and their use; Nernst equation
- I. Thermodynamic and thermochemical calculations
- J. Kinetics calculations

AP Chemistry: Curricular Requirements from the College Board

- The course provides instruction in each of the following five content areas outlined in the Course Description:
 - Structure of Matter (Atomic theory and atomic structure, Chemical bonding)
 - States of Matter (Gases, Liquids and solids, Solutions)
 - Reactions (Reaction types, Stoichiometry, Equilibrium, Kinetics, Thermodynamics)
 - Descriptive Chemistry (Relationships in the periodic table)
 - Laboratory (Physical manipulations; Processes and procedures; Observations and data manipulation; Communication, group collaboration, and the laboratory report)
- The course emphasizes chemical calculations and the mathematical formulation of principles.
- The course includes a laboratory component comparable to college-level chemistry laboratories. A minimum of one double-period per week or its equivalent is spent engaged in laboratory work. A hands-on laboratory component is required. Each student should complete a lab notebook or portfolio of lab reports. Note: Online course providers utilizing virtual labs (simulations rather than hands-on) should submit their laboratory materials for the audit. If these lab materials are determined to develop the skills and learning objectives of hands-on labs, then courses which use these labs may receive authorization to use the "AP" designation. (For information on the requirements for an AP Chemistry laboratory program, the Guide for the Recommended Laboratory Program is included in the Course Description.)*
- This course is designed to be the equivalent of the general chemistry course usually taken during the first college year. For some students, this course enables them to undertake, as freshmen, second-year work in the chemistry sequence at their institution or to register for courses in other fields where general chemistry is a prerequisite. For other students, the AP Chemistry course fulfills the laboratory science requirement and frees time for other courses.
- AP Chemistry should meet the objectives of a good general chemistry course. Students should attain a depth of understanding of fundamentals and a reasonable competence in dealing with chemical problems. The course should contribute to the development of the students' abilities to think clearly and to express their ideas, orally and in writing, with clarity and logic. The college course in general chemistry differs qualitatively from the usual first secondary school course in chemistry with respect to the kind of textbook used, the topics covered, the emphasis on chemical calculations and the mathematical formulation of principles, and the kind of laboratory work done by students. Quantitative differences appear in the number of topics treated, the time spent on the course by students, and the nature and the variety of experiments done in the laboratory. Secondary schools that wish to offer an AP Chemistry course must be prepared to provide a laboratory experience equivalent to that of a typical college course.
- To develop the requisite intellectual and laboratory skills, AP Chemistry students need adequate classroom and laboratory time. It is expected that a minimum of 290 minutes per week will be allotted for an AP Chemistry course. Of that time, a minimum of 90 minutes per week, preferably in one session, should be spent in the lab. (Time devoted to class and laboratory demonstrations should not be counted as part of the laboratory period.) In addition, students will probably need to spend at least five hours a week studying outside of class.
- The AP Chemistry course is designed to be taken after the completion of a first course in high school chemistry. It is strongly recommended that credit in a first-year high school chemistry course be a prerequisite for enrollment in an AP Chemistry class. In addition, the recommended mathematics prerequisite for an AP Chemistry class is the successful completion of a second-year algebra course. The advanced work in chemistry should not displace any other part of the student's science curriculum. It is highly desirable that a student have a course in secondary school physics and a four-year college preparatory program in mathematics.