

AP QUICK REVIEW

STP = 0°C and 1 atm at STP 1 mole = 22.4 L $K = ^\circ C + 273$

$$D = \frac{m}{V} \quad \text{Density equation}$$

$$\begin{aligned} \text{Average Atomic Mass} &= \sum \left(\frac{\% \text{ abundance}}{100} \right) (\text{mass of isotope}) \\ &= \left(\frac{\% \text{ abund. \# 1}}{100} \right) (\text{mass of isotope 1}) + \left(\frac{\% \text{ abund. \# 2}}{100} \right) (\text{mass of isotope 2}) \dots \end{aligned}$$

Solution Stoichiometry:

Determining Molarity:

$$\text{Molarity} = \frac{\text{Moles}}{\text{Volume (in liters)}}$$

Molarity and Volume:

$$M_1V_1 = M_2V_2$$

USE FOR DILUTION PROBLEMS

KEY SOLUBILITY RULES

1. Salts of ammonium (NH_4^+) and Group I are always soluble.
2. All Cl^- , Br^- , I^- are soluble except with Ag^+ , Hg_2^{2+} , and Pb^{2+} which are insoluble.
3. ClO_3^- , NO_3^- , and CH_3COO^- are soluble.
4. SO_4^{2-} are soluble except with; Sr^{2+} , Ba^{2+} , Hg_2^{2+} , Pb^{2+} , which are insoluble $\approx (\text{Ca}^{2+})$

Double displacement. Precipitation, neutralization, gas forming. H_2CO_3 in water = H_2O & CO_2

Single displacement or redox replacement: (metals displace metals and nonmetals displace nonmetals)

Combination or synthesis = two reactants result in a single product

- Metal oxide + water \rightarrow metallic hydroxide (base)
- Nonmetal oxide + water \rightarrow nonbinary acid
- Metal oxide + nonmetal oxide \rightarrow nonbinary salt

Decomposition = one reactant becomes several products

- Metallic hydroxide \rightarrow metal oxide + water
- Nonbinary acid \rightarrow nonmetal oxide + water
- Nonbinary salt \rightarrow metal oxide + nonmetal oxide
- Metallic chlorates \rightarrow metallic chlorides + oxygen
- Electrolysis decompose compound into elements (water in dilute acids or solutions of dilute acids)
- Hydrogen peroxide \rightarrow water + oxygen
- Metallic carbonates \rightarrow metal oxides + carbon dioxide
- Ammonium carbonate \rightarrow ammonia, water and carbon dioxide.

Hydrolysis = compound reacting with water.

- Watch for soluble salts that contain anions of weak acid the anion is a conjugate base and cations of weak bases that are conjugate acids.

Reactions of coordinate compounds and complex

- Complex formation by adding excess source of ligand to transitional metal of highly charged metal ion such as Al^{3+} Al = 4 ligands others 2X ox #
- Breakup of complex by adding an acid \rightarrow metal ion and the species formed when hydrogen from the acid reacts with the ligand

Lewis acid base reactions = formation of coordinate covalent bond

Redox = change in oxidation state = a reaction between an oxidizer and a reducer.

Recognized:

1. Familiarization with important oxidizers and reducers
2. "added acid" or "acidified"
3. an oxidizer reacts with a reducer of the same element to produce the element at intermediate oxidation state

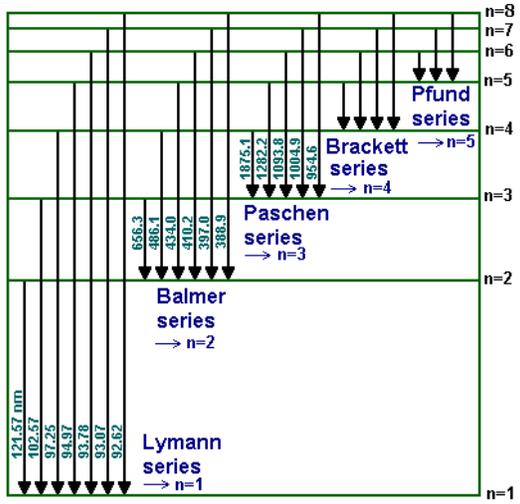
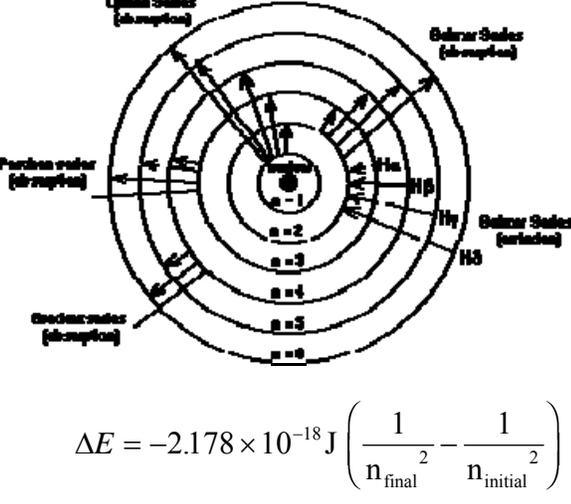
AP Quick Review

OXIDIZERS		REDUCERS	
MnO ₄ ⁻ in acid	Mn ²⁺	Halide ions	Halogens
MnO ₂ in acid	Mn ²⁺	Metal element	Metal ion
MnO ₄ ⁻ in neutral or basic solution	MnO ₂	Sulfite	Sulfate
Cr ₂ O ₇ ²⁻ in acid	Cr ³⁺	Nitrite	Nitrite
HNO ₃ concentrated	NO ₂	Halogen element in dilute basic solution	Hypohalite ion an halide ion
HNO ₃ dilute	NO	Halogen element in concentrated basic solution	Halite ion
H ₂ SO ₄ , hot, concentrated	SO ₂	Metal ous ion	Metallic ion
Metal ic ions	Metal ous ions	H ₂ O ₂	O ₂
Halogens diatomic	Halide ions	C ₂ O ₄ ²⁻	CO ₂
Na ₂ O ₂	NaOH		
HClO ₄	Cl ⁻		
H ₂ O ₂	H ₂ O		

Combustion - complete combustion of hydrocarbons results in CO₂ and H₂O. (results in forming the oxide of the elements of the compound)

Gas Laws:

<p>Kinetic molecular Theory: particles in constant random motion particles do not have volume particles do not attract T ↑with ↑KE = ½m×v²</p>	<p>Ideal: particles 0 volume no attraction, deviate from ideal at low temperature and high pressure Van der Waal's Equation $\left[P_{\text{obs}} + a \left(\frac{n}{V} \right)^2 \right] (V - nb) = nRT$ a corrects for attraction & b corrects for volume molecules that deviate from ideal gas behavior: large volume and large attraction between molecules ideal gases: molecules: have zero volume and do not attract or repel each other</p>
<p>The Ideal Gas Law: (P = pressure in atm, V = volume in liters, n = moles, R = 0.08201 L*atm/mol*K, T = temperature in Kelvin's)</p>	<p>PV = nRT</p>
<p>Combine Gas Law</p>	$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$
<p>Mole fraction; $\chi = \frac{\text{moles A}}{\text{total moles}}$</p>	<p>$P_A = P_{\text{total}} \chi_A$ $P_{\text{total}} = P_A + P_B + P_C + \dots$</p>
<p>Rate of Effusion:</p>	$\frac{\text{Rate of Effusion for Gas 1}}{\text{Rate of Effusion for Gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$
<p>Boyle's Law:</p>	<p>PV = K</p>
<p>Charles Law:</p>	<p>V/T = K</p>

<p>Root Mean Square: ($R = 8.3145 \text{ J/K} \cdot \text{mol}$, $M = \text{mass of a mole of the gas in kg}$)</p>	$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$
<p>Atomic Theory:</p>	
<p>($n = \text{integer}$, $h = \text{Planck's constant} = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$)</p>	$\Delta E = n h \nu$
<p>Energy per photon: ($c = 2.99979 \times 10^8 \text{ m/s}$)$\nu =$</p>	$\text{Energy}_{\text{photon}} = \frac{hc}{\lambda}$
<p>de Broglie's equation:</p>	$\lambda = \frac{h}{m v}$
<p>Beer's Law: where $A = \text{absorbance}$, $a = \text{wavelength-dependent absorptivity coefficient}$, $b = \text{path length}$, and $c = \text{analyte concentration}$.</p>	$A = abc$
 <p>The diagram shows energy levels for $n=1$ to $n=8$. Transitions are labeled as follows: - Lyman series (absorption): $n=1 \rightarrow n=2, 3, 4, 5, 6, 7, 8$ - Balmer series (absorption): $n=2 \rightarrow n=3, 4, 5, 6, 7, 8$ - Paschen series (absorption): $n=3 \rightarrow n=4, 5, 6, 7, 8$ - Brackett series (absorption): $n=4 \rightarrow n=5, 6, 7, 8$ - Pfund series (absorption): $n=5 \rightarrow n=6, 7, 8$ Wavelengths for Lyman series are listed: 121.57 nm, 102.57, 97.25, 94.97, 93.78, 93.07, 92.62 nm.</p>	 <p>The Bohr model shows concentric circles for $n=1$ to $n=8$. Labels include Lyman Series (absorption), Balmer Series (absorption), Paschen series (absorption), and Balmer Series (emission). The energy level formula is given as: $\Delta E = -2.178 \times 10^{-18} \text{ J} \left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)$</p>
<p>Principle Quantum Number (n) has integral values 1, 2, 3, The principle quantum number is related to size of the electron cloud, $n = 1, 2, 3, \dots$ determines the energy of the e⁻.</p>	<p>Angular Momentum Quantum Number (ℓ) has integral values $\{0 \text{ to } n-1\}$. This quantum number gives us the shape of the probability pattern and each shape is referred to as a sublevel $\ell = \{0 \text{ to } (n-1)\}$ the type of orbital (subshell) 0 = s, 1 = p, 2 = d, 3 = f</p>
<p>Magnetic Quantum Number (m_ℓ) has integral values $\{-\ell \text{ to } +\ell\}$. The Magnetic quantum number refers to the orientation of each sublevel. m or $m_\ell = \{-\ell \text{ to } +\ell\}$ which orientation of the orbital (x, y, z... for p orbitals)</p>	<p>Magnetic Spin Quantum Number (m_s) differentiates the two electrons that can exist in an orbital and has the value of $+1/2$ and $-1/2$. s or $m_s = +1/2$ or $-1/2$ the "spin" of the electron Diamagnetic = paired e⁻ Paramagnetic = unpaired e⁻</p>
<p>Dalton- Atomic Theory, Thomson- cathode ray experiment / electron and proton, Rutherford's- gold leaf/ nucleus, Bohr- energy levels, Schrödinger- quantum theory(Heisenberg's- uncertainty principle and deBroglie-dual nature) Milliken- charge of electron, Chadwick- neutron, Mosley - atomic number, Pauli's- exclusion principle, Mendeleev- periodic table</p>	

AP Quick Review

4

Shielding is constant →

Atomic radius decreases →

Ionization energy increases →

Electronegativity increases →

Nuclear charge increases →

Nuclear charge increases ↓

Shielding increases ↓

Atomic radius increases ↓

Ionic size increases ↓

Ionization energy decreases ↓

Electronegativity decreases ↓

Ionic size (cations) decreases →

Ionic size (anions) decreases →

(a) Radial probability vs. Distance from the nucleus. Shows a single peak with 'Penetration' indicated near the nucleus and 'Most probable distance from the nucleus' at the peak.

(b) Radial probability vs. Distance from the nucleus. Shows three curves: 3s (highest peak, furthest from nucleus), 3p (middle peak), and 3d (lowest peak, closest to nucleus). 'Penetration' is indicated for each curve.

-Stating the trend is not an explanation:

- ease of removing e^- : $f > d > p > s$ because of penetration -s e^- spends a greater percent of time closer to nucleus than p and p more than d
- $\uparrow n$ or radius of atom \downarrow nuclear attraction for valence e^-
- $\uparrow Z$ same n \uparrow nuclear attraction for valence e^-
- \uparrow shielding \downarrow nuclear attraction for valence e^-

(shielding is only a factor when comparing e^- from different n's)

Atomic Number	Element	First Ionization Energy (kJ/mol)
3	Li	~520
4	Be	~900
5	B	~800
6	C	~1080
7	N	~1400
8	O	~1310
9	F	~1680
10	Ne	~2080

Ionization energy Exceptions

- $Xs^2 > Xp^1$ e.g. ${}^4\text{Be} > {}^5\text{B}$
 WHY? - (s shields p \therefore p $<$ further from nucleus than s) (s e^- spends more time closer to nucleus than e^- in p)
 Therefore, it requires less energy to remove the first electron in a p orbital than to remove one from a filled s orbital. Or :The energy of an electron in an Xp orbital is greater than the energy of an electron in its respective Xs orbital. Therefore, it requires less energy to remove the first electron in a p orbital than it is to remove one from a filled s orbital.(further from nucleus)
- $Xp^3 > Xp^4$ e.g. ${}^7\text{N} > {}^8\text{O}$
 WHY? - After the separate degenerate orbitals have been filled with single electrons, the fourth electron must be paired. The electron-electron repulsion makes it easier to remove the outermost, paired electron.
 Subsequent ionization energies increase greatly once an ion has reached the state like that of a noble gas (what not why!!). Why is --. Fewer core electrons, closer to nucleus, greater $Z_{\text{eff}} \rightarrow$ Takes more energy to remove.

Atomic Number	Element	Electron Affinity (kJ/mol)
1	H	~78
2	He	< 0
3	Li	~60
4	Be	< 0
5	B	~48
6	C	~122
7	N	< 0
8	O	~141
9	F	~328

Electron Affinity Exceptions

WHY? - $Xs^2 < 0$: diamagnetic atom with all paired electrons. Electron-electron repulsion.

$Xp^3 < 0$: atom with 3 unpaired p-orbital electrons each occupying its own subshell. The fourth electron will be the second electron in an orbital and will experience electron-electron repulsion.

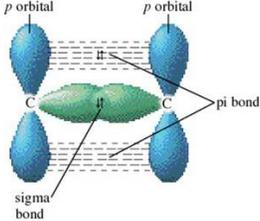
$Xp^6 < 0$: next electron in energy level further from nucleus

Bonding: know the difference between structural pair geometry and molecular geometry

Hybridization	# of σ Bonds	# of Non-Bonding Pairs	Molecular Shape	Bond Angles	Example
sp	2	0	 Linear	180°	BeH ₂ , CO ₂
sp ²	3	0	 Trigonal planar	120°	SO ₃ , BF ₃
sp ²	2	1	 Angular	<120°	SO ₂ , O ₃
sp ³	4	0	 Tetrahedral	109.5°	CH ₄ , CF ₄ , SO ₄ ²⁻
sp ³	3	1	 Trigonal pyramidal	<109.5°	NH ₃ , PF ₃ , AsCl ₃
sp ³	2	2	 Angular	<109.5°	H ₂ O, H ₂ S, SF ₂
sp ³ d	5	0	 Trigonal bipyramidal	120°, 90°	PF ₅ , PCl ₅ , AsF ₅
sp ³ d	4	1	 Sawhorse (irregular tetrahedron)	<120°, <90°	SF ₄
sp ³ d	3	2	 T-shaped	<90°	ClF ₃
sp ³ d	2	3	 Linear	180°	XeF ₂ , I ₃ ⁻ , IF ₂
sp ³ d ²	6	0	 Octahedron	90°	SF ₆ , PF ₆ ⁻ , SiF ₆ ²⁻
sp ³ d ²	5	1	 Square pyramidal	<90°	IF ₅ , BrF ₅
sp ³ d ²	4	2	 Square planar	90°	XeF ₄ , IF ₄ ⁻

AP Quick Review

6

<p>Metallic bond: delocalized e^- Ionic: transfer of e^- Nonpolar Covalent: equal sharing of e^- Polar Covalent: unequal sharing of e^- σ overlap s/s, s/p, p/p - π overlap p/p single σ, double σ-π, triple σ-π-π single < double < triple: strength single > double > triple: length</p>	 <div style="display: flex; justify-content: center; align-items: center; margin-top: 10px;"> $\text{:}\ddot{\text{O}}=\ddot{\text{S}}-\ddot{\text{O}}\text{:} \leftrightarrow \text{:}\ddot{\text{O}}-\ddot{\text{S}}=\ddot{\text{O}}\text{:}$ </div> <div style="border: 1px solid black; padding: 5px; margin-top: 10px; width: fit-content; margin-left: auto; margin-right: auto;"> Resonance occurs when you have a <i>combination</i> of a multiple and single </div>
---	---

Hybrids explains why bonds in molecules with different atomic orbitals behave as identical bonds ie. CH_4

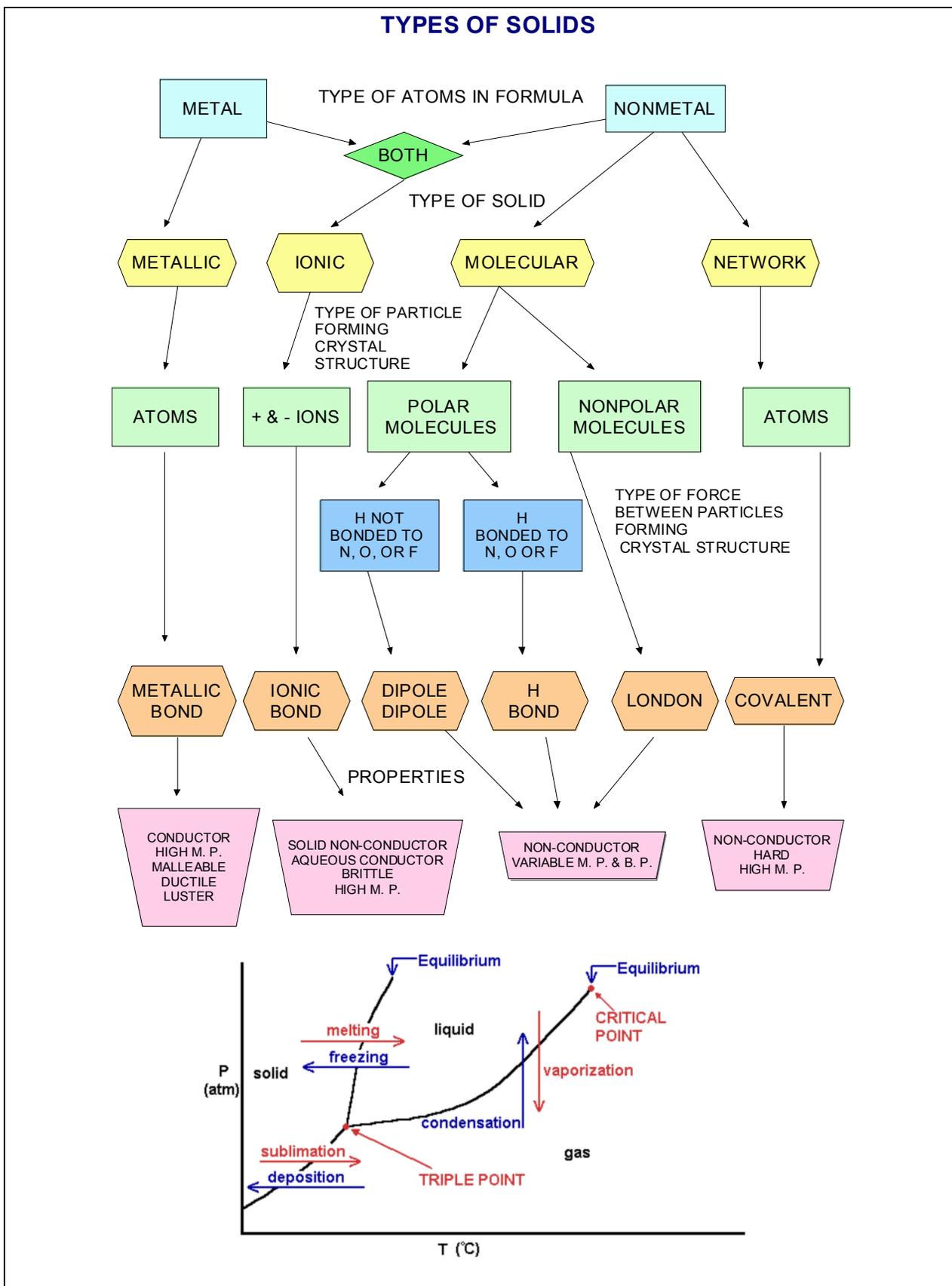
Resonances explains why bonds in molecules with a combination of single, double or triple bonds behave identical ie. C_6H_6

0 → 0.4 NONpolar covalent 0.41 → 1.67 POLAR covalent above 1.67—IONIC

Calculating lattice energy (k is proportionality constant, Q is charge of ions, r = shortest distance between centers of the cations and anions)	$\text{Lattice energy} = k \left(\frac{Q_1 Q_2}{r} \right)$ strength of ionic bond
	$\Delta H_{\text{rxn}} = \sum \text{bond energy of reactants} - \sum \text{bond energy of products}$ <div style="display: flex; justify-content: space-around; font-size: small;"> bond broken bonds formed </div>
Bond Order:	# of shared pairs linking X & Y ÷ # of X-Y links
Formal Charge = number of valence e^- - (# of lone e^-) - $\frac{1}{2}$ (# of bonding e^- 's)	

Intermolecular Forces:

TYPE OF INTERACTION	CHARACTERISTICS OF INTERACTION	TYPE OF SUBSTANCE	MAGNITUDE RANGE KJ/MOLE
DIPOLE-DIPOLE	Positive end of a permanent dipole aligns itself with negative end of a permanent dipole on another molecule. {dipole moment}	TWO POLAR MOLECULAE	5-25
HYDROGEN BOND	Special type of dipole-dipole. Is strong due to the lack of shielding around a hydrogen nucleus (when hydrogen is bonded to a highly electronegative atom.) {dipole moment}	TWO POLAR MOLECULES WITH HYDROGEN BONDED TO A N, F, OR O ATOM	5-25
INDUCED DIPOLE-INDUCED DIPOLE (LONDON DISPERSION)	Attractive forces between molecules resulting from momentary mutual distortion (induced dipole) of electron clouds. (temporary dipoles) Increase in strength with increase molecular weight (electron cloud size) called polarizability. {induced dipole,- induced dipole}	TWO NONPOLAR MOLECULES	0.05 -40

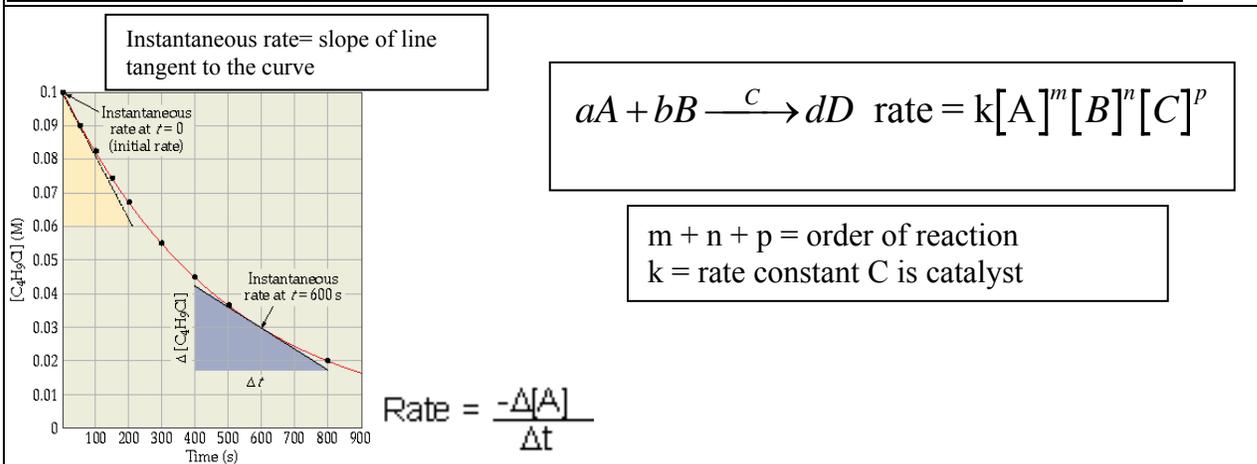


Solutions:

Molality:	$\text{molality} = \frac{\text{moles of solute}}{\text{kg solvent}}$
P_{soln} : vapor pressure of the solution χ_{solvent} : mole fraction of solvent P_{solvent} : vapor pressure of pure solvent	Raoult's Law $P_{\text{soln}} = \chi_{\text{solvent}} P_{\text{solvent}}$
Mass Percent:	$\text{Mass percent} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
Osmotic pressure formula for electrolytes π : osmotic pressure M: molarity R: 0.0821 L*atm/K*mol	$\pi = iMRT$
Henry's Law P: partial pressure of gaseous solute C: concentration of dissolved gas k: constant	$P = kC$
Boiling point elevation m: molality K _b : constant	$\Delta T_b = iK_b \cdot m$
Freezing Point Depression m: molality K _f : constant	$\Delta T_f = iK_f \cdot m$
van't Hoff factor	$i = \frac{\text{moles of particles in solution}}{\text{moles of solute dissolved}}$

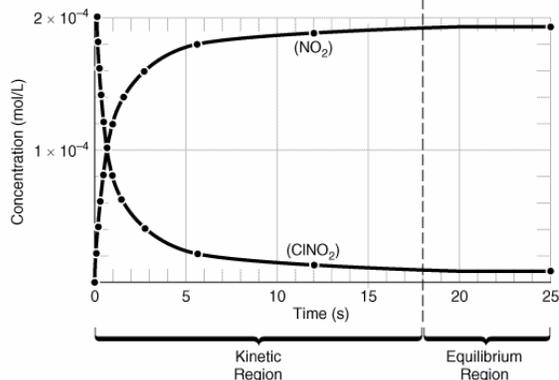
Chemical Kinetics:

FACTOR	AFFECT ON RATE	COLLISION THEORY
CONCENTRATION OF REACTANTS	increasing concentration increases rate (only reactants in the rate determining step affect the rate)	more molecules more collisions
TEMPERATURE	increasing temperature increases rate (approximately every 10°C increase in temperature doubles rate)	molecules have more kinetic energy ~ collide more often and harder ~ collisions more frequent and a greater % of the collisions are effective
CATALYST:	increase the rate of a reaction but not consumed during the reaction	lowers the activation energy therefore a greater % of collisions are effective
SURFACE AREA OF REACTANTS	increase surface area increases rate	more collisions



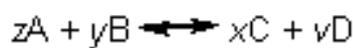
General Rate Law rate and [], Integrated rate law time and []					
Order in [A]	Rate Law	Integrated Form $y = mx + b$	Straight Line Plot	Half-Life $t_{1/2}$	
zeroth order (n=0)	$rate = k [A]^0 = k$	$[A]_t = -k t + [A]_0$	$[A]_t$ vs t (slope = -k)	$t_{1/2} = \frac{[A]_0}{2k}$	
first order (n=1)	$rate = k [A]^1$	$\ln[A]_t = -k t + \ln[A]_0$	$\ln[A]_t$ vs t (slope = -k)	$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$	
second order (n=2)	$rate = k [A]^2$	$\frac{1}{[A]_t} = k t + \frac{1}{[A]_0}$	$\frac{1}{[A]_t}$ vs t (slope = k)	$t_{1/2} = \frac{1}{k[A]_0}$	
Arrhenius Equation the whole $e^{(stuff)}$ equation represents the fraction of collisions with sufficient energy to produce a rxn					
Derived Arrhenius Equation $\ln k = \frac{-E_a}{R} \left(\frac{1}{T}\right) + \ln A$			$\ln k = \frac{-E_a}{R} \left(\frac{1}{T}\right) + \ln(A)$ $y = mx + b$ [graphical method] $\ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right)$ [plug in equation]		
Endothermic		Exothermic		Copyright 1999 by John Wiley and Sons, Inc. All rights reserved.	

Chemical Equilibrium:



To determine direction of reaction calculate Q: **If**
 $Q > K$ reaction \leftarrow **If** $Q < K$ reaction \rightarrow

Calculating Equilibrium Constant:
(leave out solids and liquids)



$$K = \frac{[C]^x [D]^v}{[A]^z [B]^y}$$

Δn = coefficients of gaseous products -
coefficients gaseous reactants

$$K_p = K_c(RT)^{\Delta n}$$

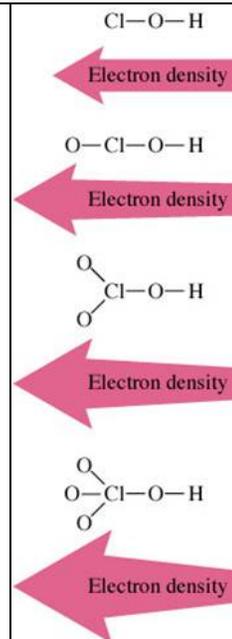
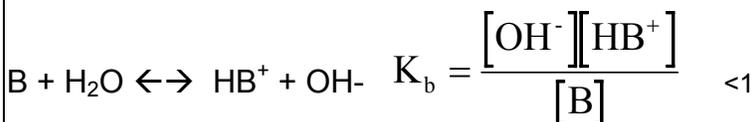
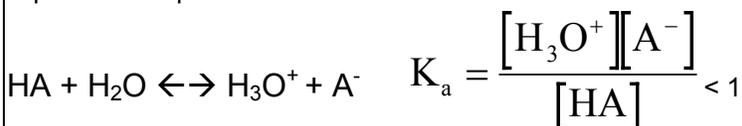
Changing the Coefficients Will Change K:

- Multiplying coefficients by a #, $K_{\text{new}} = K_{\text{old}}^{\#}$
- Reversing the equation, $K_{\text{new}} = 1/K_{\text{old}}$
- Adding several equations to get a net equation
 $K_{\text{new}} = (K_{\text{old}})(K_{\text{old}})\dots$

Acids and Bases:

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ \text{pOH} &= -\log [\text{OH}^-] \\ K_w &= 1 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-] \\ 14 &= \text{pH} + \text{pOH} \\ [\text{H}_3\text{O}^+] &= 10^{-(\text{pH})} \\ [\text{OH}^-] &= 10^{-(\text{pOH})} \end{aligned}$$

Equilibrium Expressions:



pH of Salts:

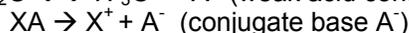
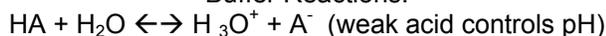
1. Strong acid + strong base = neutral salt
2. Strong acid + weak base = acidic salt
3. Weak acid + strong base = basic salt
4. Weak acid + weak base = ? (must look at K values to decide)

<p>CB's (anions of weak acids) - react like nonhydroxide bases in water: NaF strong base weak acid salt anion is a CB: $F^- + H_2O \leftrightarrow HF + OH^-$ CA's (cations of weak bases) - react like acids in water: NH₄Cl weak base strong acid salt cation is a CB: $NH_4^+ + H_2O \leftrightarrow H_3O^+ + NH_3$</p>	$K_w = [OH^-][H^+] = 10^{-14} @25^\circ C$ $K_w = K_a \cdot K_b$ $pK_w = 14$
$\underbrace{HCl}_{\text{acid}} + \underbrace{NH_3}_{\text{base}} \rightleftharpoons \underbrace{NH_4^+}_{\text{CA}} + \underbrace{Cl^-}_{\text{CB}}$	<p>amphoteric</p> $\left\{ \begin{array}{l} \underbrace{H_2PO_4^-}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons \underbrace{H_3O^+}_{\text{CA}} + \underbrace{HPO_4^{2-}}_{\text{CB}} \\ \underbrace{H_2PO_4^-}_{\text{base}} + \underbrace{H_2O}_{\text{acid}} \rightleftharpoons \underbrace{H_3PO_4}_{\text{CA}} + \underbrace{OH^-}_{\text{CB}} \end{array} \right.$
<p>% Dissociation</p>	$\% \text{ dissociation} = \frac{\text{amount dissociated}}{\text{initial concentration}} \times 100$

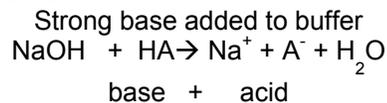
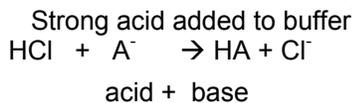
Summary of Buffer Reactions:

Weak Acid and Conjugate Base Buffer:

Buffer Reactions:

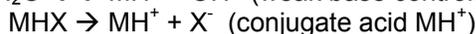
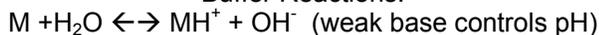


Neutralization Reactions:

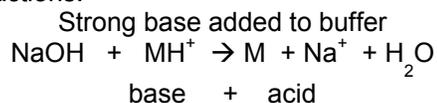
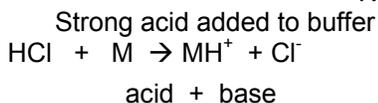


Weak Base and Conjugate Acid Buffer

Buffer Reactions:



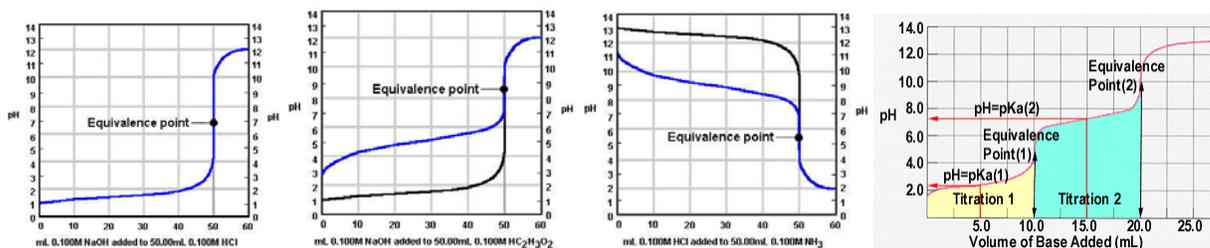
Neutralization Reactions:



✎ RICE TABLE AFTER STRONG ACID OR BASE IS ADDED TO A BUFFER: CHECK ✎

- When a strong acid is added to buffer subtract M of strong acid from the base and add to the acid
- When strong base is added subtract the M of strong base from the acid and add to the base:

TITRATION

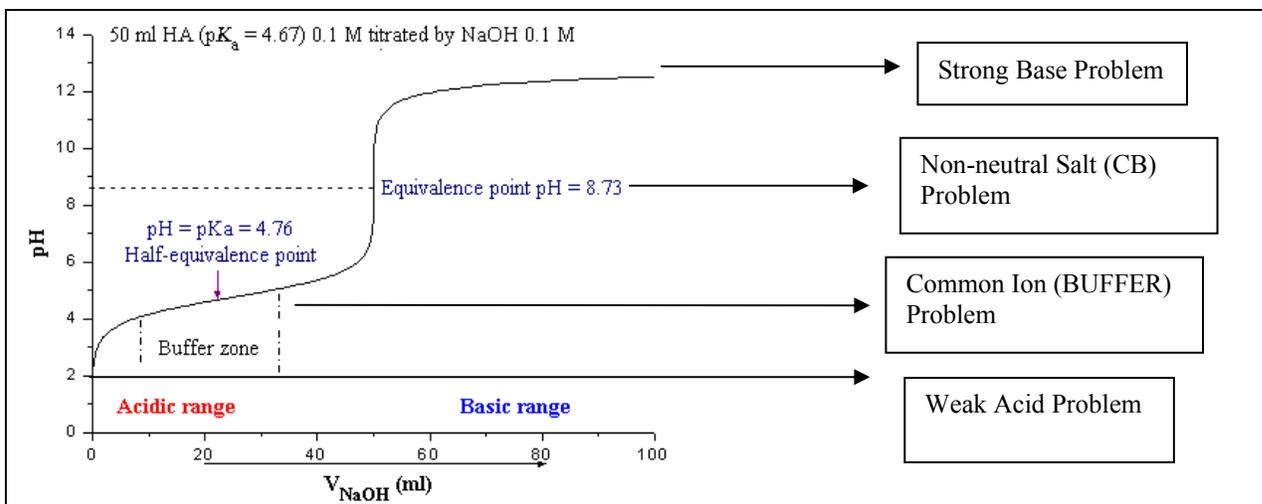


pH = pKa at 1/2 equivalence point

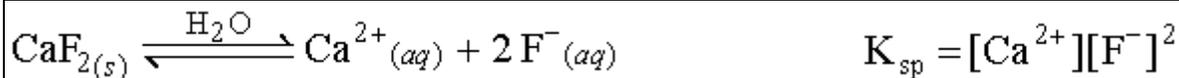
Titration: If an ACID & BASE are reacting write a neutralization reaction:

(both strong) do Stoichiometry to find molarity of substance of interest; **AT** equivalence pt 7/ (one of them weak)

SCAM TABLE then RICE TABLE, **AT** equivalence pt a **non neutral** salt will be CA or CB problem $K_w = K_a \cdot K_b$



Solubility Equilibrium:

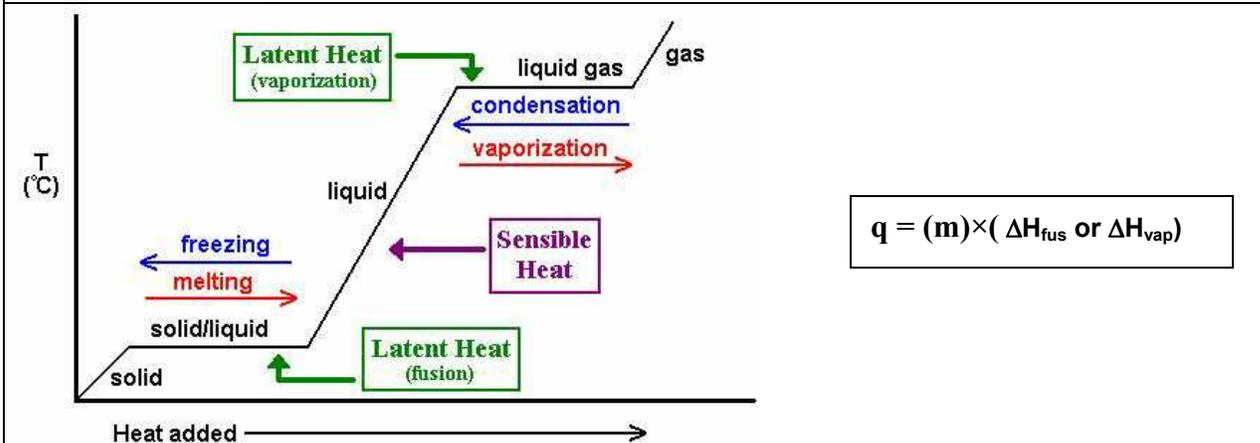


$Q_{sp} < K_{sp}$ - This means there are not enough ions in the solution. In order to return to equilibrium, more of the solid salt must dissociate into its ions.

$Q_{sp} = K_{sp}$ - This means that the system is at equilibrium.

$Q_{sp} > K_{sp}$ - This means that there are too many ions in the solution. In order to return to equilibrium, the excess ions will precipitate to form more solid.

Thermochemistry:



standard conditions 1 atm and 25°C

<p>C = heat capacity</p> $C_p = \frac{q}{\Delta t} \text{ or } C_p = \frac{q}{\Delta t \times m}$	<p>C_p also called specific heat</p> <p>m = mass of solution or total mass</p>
<p>$\Delta H = q$ coffee cup calorimetry</p>	<p>$q = mc\Delta T$</p>
<p>ΔH_f° for element = 0 → Unit KJ/mole</p>	$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{products}) - \sum n\Delta H_f^\circ(\text{reactants})$
<p>ΔS° is <u>not</u> 0 → Units in J//mole K</p>	$\Delta S^\circ = \sum n\Delta S^\circ(\text{products}) - \sum n\Delta S^\circ(\text{reactants})$
<p>ΔG_f° for element = 0 → Unit KJ/mole</p>	$\Delta G^\circ = \sum n\Delta G_f^\circ(\text{products}) - \sum n\Delta G_f^\circ(\text{reactants})$

$\Delta S_{\text{surroundings}} = \frac{-\Delta H}{\Delta T}$	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ <u>$-\Delta G^\circ = \text{rxn spontaneous}$</u>
$\Delta G = 0$ at equilibrium	$\Delta G = \Delta G^\circ + RT \ln Q$ $\Delta G^\circ = -RT \ln K$ (can use ΔG or ΔG°)

TABLE 20.1 Criteria for Spontaneous Change: $\Delta G = \Delta H - T\Delta S$

Case	ΔH	ΔS	ΔG	Result	Example
1	-	+	-	spontaneous at all temp	$2 \text{ N}_2\text{O(g)} \longrightarrow 2 \text{ N}_2\text{(g)} + \text{O}_2\text{(g)}$
2	-	-	{ - + }	{ spontaneous at low temp nonspontaneous at high temp }	$\text{H}_2\text{O(l)} \longrightarrow \text{H}_2\text{O(s)}$
3	+	+	{ + - }	{ nonspontaneous at low temp spontaneous at high temp }	$2 \text{ NH}_3\text{(g)} \longrightarrow \text{N}_2\text{(g)} + 3 \text{ H}_2\text{(g)}$
4	+	-	+	nonspontaneous at all temp	$3 \text{ O}_2\text{(g)} \longrightarrow 2 \text{ O}_3\text{(g)}$

REACTION STATE	EQUILIBRIUM CONSTANT "K"	GIBBS FREE ENERGY "ΔG"	CELL POTENTIAL "ΔE"
SPONTANEOUS	greater than one	negative	positive
EQUILIBRIUM	zero	Zero	Zero
NON-SPONTANEOUS	Less than one	positive	negative

Electrochemistry:

LEO goes GER
lose electrons oxidize gain electrons reduction



the reactant reduced is the oxidizing agent and visa versa

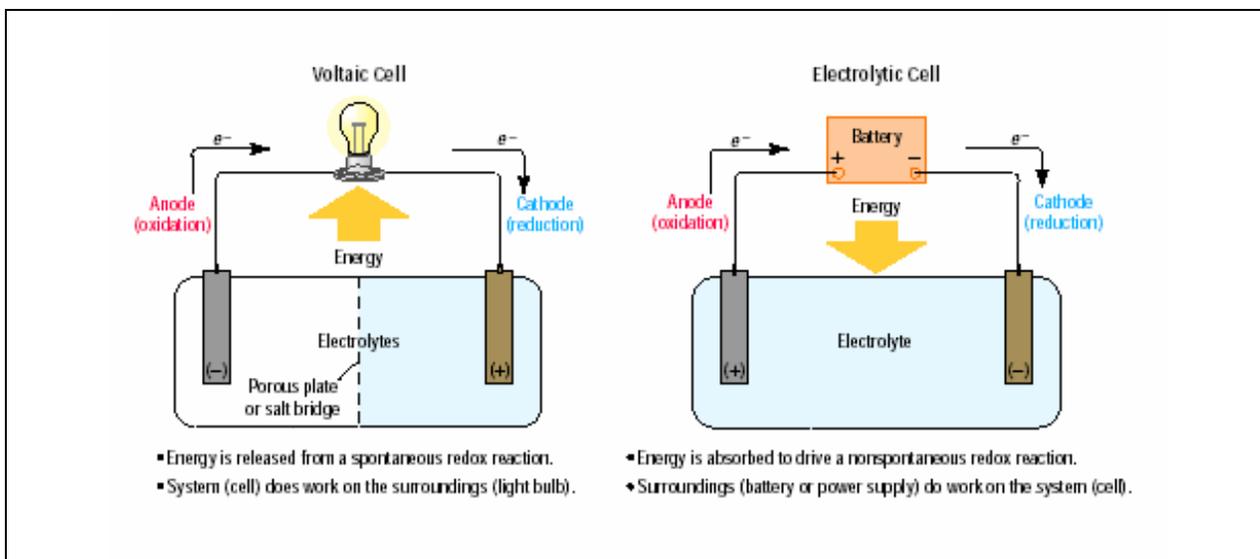
$$1 \text{ V} = \frac{1 \text{ J}}{1 \text{ C}} = \frac{(\text{joule})}{(\text{coulomb})}$$

E_{cell} is: + spon rxn, - nonspon rxn, 0 rxn at eq

$$E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}}$$

Largest E°_{red} is reduced \therefore best oxidizing agent

Anode	Cathode
<ul style="list-style-type: none"> • Oxidation • Anions to it from SB • - in voltaic cells • + in electrolytic cells • Electrons flow away from it • Cell notation $\text{Cl}^- / \text{Cl}_2$ 	<ul style="list-style-type: none"> • Reduction • Cations flow to it from SB • + in voltaic cells • - in electrolytic cells • Electrons flow towards it • Cell notation $\text{Cu}^{2+} \rightarrow \text{Cu}_{(s)}$
Electrochemical Cell: Voltaic cell	
<p style="text-align: center;"> $\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^-$ $\text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s)$ </p> <p style="text-align: center;"> → Movement of cations ← Movement of anions </p>	
<p>Standard conditions 1 atm 25°C</p>	$w_{\text{max}} = -qE \quad \Delta G^{\circ} = -n f E^{\circ}$ <p>n = number of moles of electrons transferred</p>
<p>Nernst Equation F = faraday's constant n = number of electrons</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 10px auto;"> <p>At equilibrium $E_{\text{cell}} = 0$</p> </div>	$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln Q = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log Q \quad \text{at } 25^{\circ}\text{C}$ $\log K = \frac{nE^{\circ}}{0.0592}$
<p style="text-align: center;">Electrolysis</p> <p>Electrolysis involves forcing a current through a cell to produce a chemical change for which the cell potential is negative ~ electrical work causes an otherwise non-spontaneous chemical reaction to occur.</p>	<p>If more than one reactant is available for reduction, the one with the <i>highest</i> E°_{red} potential is most likely to be reduced. Metal ions or water can be reduced. When electrolysis occurs in aqueous solutions, if the metal has a reduction potential smaller than -0.8, then only water is reduced because water has the larger E°_{red}. Negative ions will be oxidized. If there are mixtures of metal ions, the metals will be reduced in order of largest E°_{red} to smallest E°_{red}.</p>



Faraday's Constant - The charge on a mole of electrons: $F = 96,500 \text{ C}$
coulomb = (amp) (sec)

ORGANIC—carbon containing compounds

- **–ane–** $\text{C}_n\text{H}_{2n+2}$; has all single bonds and is therefore saturated with hydrogens
- **–ene–** C_nH_{2n} ; has double bonds and is therefore unsaturated [dienes have 2 double bonds, trienes have 3, etc.; both are polyunsaturated]
- **–yne–** $\text{C}_n\text{H}_{2n-2}$; have a triple bond and are also unsaturated
- **ISOMERS**—same structural formula yet different arrangement of atoms—can change physical properties;
- **Alcohols have –OH—hydroxyl; *not a base*!!!!!!!**
 - large IMF's due to H-bonding; increases BP, lowers VP, etc.
 - Increases solubility
- **–COOH—carboxylic acid group; weak organic acids**

Alkane		Aldehyde	
Alkene		Ketone	
Alkyne		Carboxylic Acid	
Alcohol		Ester	
Ether		Amine	

AP Quick Review

16

Flame Test Colors: Cu green, Li & Sr & red, Na yellow, Ba yellow green, K violet pink
Colors of stuff: $\text{Cr}_2\text{O}_7^{2-}$ orange, CrO_4^{2-} yellow, MnO_4^- purple, Cr^{3+} Cu^{2+} , Ni^{2+} , green or blue, PbI_2 yellow, HF etches glass, NH_3 fertilizer, NaN_3 air bags, Zn added to Fe prevent rusting,
Gases: O_2 is colorless and sustains combustion, H_2 is colorless and explodes!, Cl_2 is green, Br_2 is orange-brown; NO_2 is brown; Iodine sublimates into a purple vapor

PROCESS	CHANGE IN A	CHANGE IN Z	CHANGE IN NEUTRON/PROTON RATIO
ALPHA EMISSION	-4	-2	INCREASE
BETA EMISSION	0	1	DECREASE
GAMMA	0	0	X
POSITRON EMISSION	0	-1	INCREASE
ELECTRON CAPTURE	0	-1	INCREASE

ALPHA	BETA	GAMMA	POSITRON	ELECTRON CAPTURE
helium nucleus	fast moving electron	high energy electromagnetic radiation	antiparticle of an electron	inner orbital electron
2 protons & 2 neutrons	electron's mass Very small	no mass	electron's mass	electron's mass
+2 charge	-1 charge	no charge	+1 charge	-1 charge
$\alpha, {}^4_2\text{He}$	$\beta, {}^0_{-1}\text{e}$	${}^0_0\gamma$	${}^0_1\text{e}$	${}^0_{-1}\text{e}$
Do not travel far Not penetrating	More penetrating than alpha	Very dangerous		Gamma rays are produced
Common with heavy nuclides	Nuclides with too many neutrons	Accompanies other forms of radiation	Nuclides with too few neutrons	Nuclides with too few neutrons
A = -4 Z = -2	A = 0 Z = +1	None	A = 0 Z = -1	A = 0 Z = -1

Test Taking Tips

Multiple Choice:

- Do not spend too much time on difficult questions.
- Go through entire test and answer all the easy questions first.
- When you come to a question that you can answer but it will take considerable time, mark the question and come back to it after you have gone through the entire test.
- You are penalized for wrong answers so do not randomly guess.
- You can mark in your test booklet so elimination works well. If you can eliminate several choices you may want to guess.

Free Response:

- Questions require you to apply and explain chemical concepts and solve multiple step problems.
- You do not have to answer in essay form and may save time using one of the following methods: bullet format, chart format or outline format.
- Write your answers in the space provided and number your answer clearly.
- There is a slight penalty for incorrect sig figs.
- Stating a Principle, Law, Theory or stating the name of the Principle law or theory is not an explanation or justification for an answer. Stating a trend is also not an explanation. State and apply the Principle, Law or Theory to the specific situation in the questions to explain your answer and use the reason for trends as explanations for trends not just stating the trend.

DESCRIPTION OF THE AP CHEMISTRY EXAM

Multiple Choice:

- 90 minutes
- 75 questions
- May not use a calculator
- 50% of total score
- Periodic Table and Reduction Potential Table are available
- No equation sheet
- This portion of the test covers more material than any individual student is expected to know.

Free Response:

- 95 minutes
- 50 % of total score
- Tables containing commonly used chemistry equations are provided with each examination for students to use when taking the free-response section.
Two parts: Part A: problems and Part B: concepts
- Part A:
55 Minutes for part A
May use calculator
Part A; students will answer three problems—one problem involving chemical equilibrium and two other problems, one of which may involve quantitative analysis of data in a laboratory-based problem.
- Part B:
40 minutes for part B
May not use calculators
Part B; three questions will be scored; Question four involves predicting equations.

Question 4 will now contain 3 required balanced net ionic equations and a question related to that equation. A Periodic Table & Standard Reduction Potentials are provided.

For example:

(i) A magnesium strip is added to a solution of silver nitrate.

Answer: $\text{Mg} + 2 \text{Ag}^+ \rightarrow \text{Mg}^{2+} + 2\text{Ag}$

(ii) Which species is oxidized? Explain

Answer: Magnesium because it has lost 2 electrons

Two other free response questions will be answered. (Questions 5 and 6) requiring applying and explaining concepts, one of which may be based on laboratory in the case that no laboratory-based problem appears in Part A