

AP QUICK REVIEW

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

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

$$\text{Average Atomic Mass} = \sum \left(\frac{\% \text{ abundance}}{100} \right) (\text{mass of isotope})$$

$$= \left(\frac{\% \text{ abund. \# 1}}{100} \right) (\text{mass of isotope 1}) + \left(\frac{\% \text{ abund. \# 2}}{100} \right) (\text{mass of isotope 2}) \dots$$

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 Ca^{2+} , Sr^{2+} , Ba^{2+} , Hg_2^{2+} , Pb^{2+} , which are insoluble.

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

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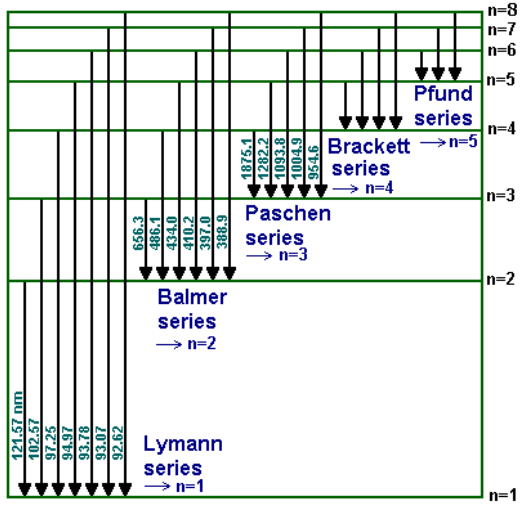
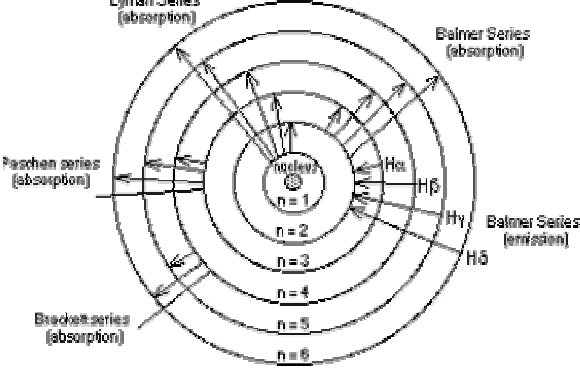
OXIDIZERS		REDUCERS	
MnO ₄ ⁻ in acid	Mn ²⁺	Halide ions	Halogens
MnO ₂ in acid	Mn ²⁺	Metal element	Metal ion
MnO ₄ ⁻ in neutral or basic solution	MnO ₂	Sulfite	Sulfate
Cr ₂ O ₇ ²⁻ in acid	Cr ³⁺	Nitrite	Nitrite
HNO ₃ concentrated	NO ₂	Halogen element in dilute basic solution	Hypohalite ion an halide ion
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 Kelvins)</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>Boyle's Law:</p>	<p>PV = K</p>
<p>Charles Law:</p>	<p>V/T = K</p>
<p>Root Mean Square: (R = 8.3145 J/ K * mol, M = mass of a mole of the gas in kg)</p>	$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$
<p>Rate of Effusion:</p>	$\frac{\text{Rate of Effusion for Gas 1}}{\text{Rate of Effusion for Gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$

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<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>Atomic Theory:</p>	
<p>(n = integer, h = Planck's constant = $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$)</p>	<p>$\Delta E = n h \nu$</p>
<p>Energy per photon: ($c = 2.99979 \times 10^8 \text{ m/s}$)$\nu =$</p>	<p>Energy_{photon} = $\frac{hc}{\lambda}$</p>
<p>de Broglie's equation:</p>	<p>$\lambda = \frac{h}{m v}$</p>
<p>Beer's Law: where A = absorbance, a = wavelength-dependent absorptivity coefficient, b = path length, and c = analyte concentration.</p>	<p>$A = abc$</p>
 <p>Energy level diagram for hydrogen showing Lyman, Balmer, Paschen, Brackett, and Pfund series with wavelength values in nm.</p>	 <p>Bohr model diagram of a hydrogen atom showing energy levels n=1 to n=6 and transitions for Lyman, Balmer, Paschen, Brackett, and Pfund series.</p> $\Delta E = -2.178 \times 10^{-18} \text{ J} \left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)$
<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, ... determines the energy of the e⁻.</p>	<p>Angular Momentum Quantum Number (ℓ) has integral values {0 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, Schrodinger- quantum theory(Heisenberg's- uncertainty principle and deBroglie-dual nature) Millikan- charge of electron, Chadwick- neutron, Mosley - atomic number, Pauli's-exclusion principle, Mendeleev- periodic table</p>	

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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

1A	2A		3A	4A	5A	6A	7A	0

← Ionic size (cations) decreases
Ionic size (anions) decreases →

(a)

(b)

-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 $\downarrow Z_{eff}$,
 $\uparrow Z$ same n $\uparrow Z_{eff}$,

\uparrow shielding $\downarrow Z_{eff}$ (shielding is only a factor when comparing e^- from different n 's)

Ionization energy Exceptions














1) $Xs^2 > Xp^1$ e.g. ${}^4\text{Be} > {}^5\text{B}$
 WHY? - s shields p $\therefore p < Z_{eff}$ 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 it is 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.

2) $Xp^3 > Xp^4$ 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_{eff} \rightarrow Takes more energy to remove.

Electron Affinity Exceptions

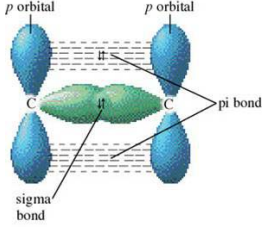
WHY? - $Xs^2 < 0$: diamagnetic atom with all paired electrons. Electron-electron repulsion.
 $Xp^3 < 0$: atom with 3 unpaired p -orbital electrons each occupying its own subshell. The fourth electron will be the second electron in an orbital and will experience electron-electron repulsion.
 $Xp^6 < 0$: 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 ⁻

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<p>Metallic bond: delocalized \bar{e} Ionic: transfer of \bar{e} Nonpolar Covalent: equal sharing of \bar{e} Polar Covalent: unequal sharing of \bar{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>
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Hybrids explains why bonds in molecules with different atomic orbitals behave as identical bonds ie. CH₄

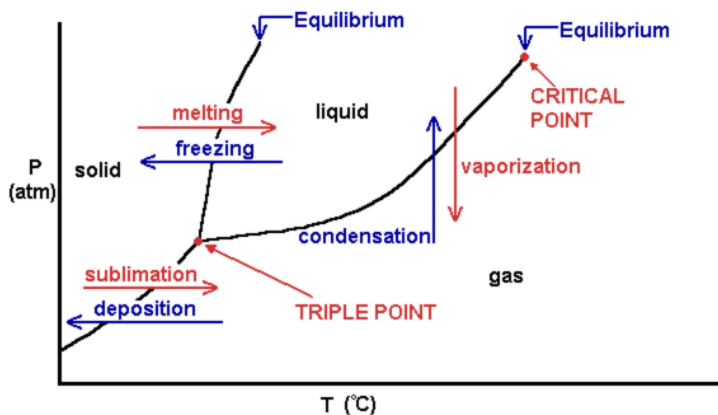
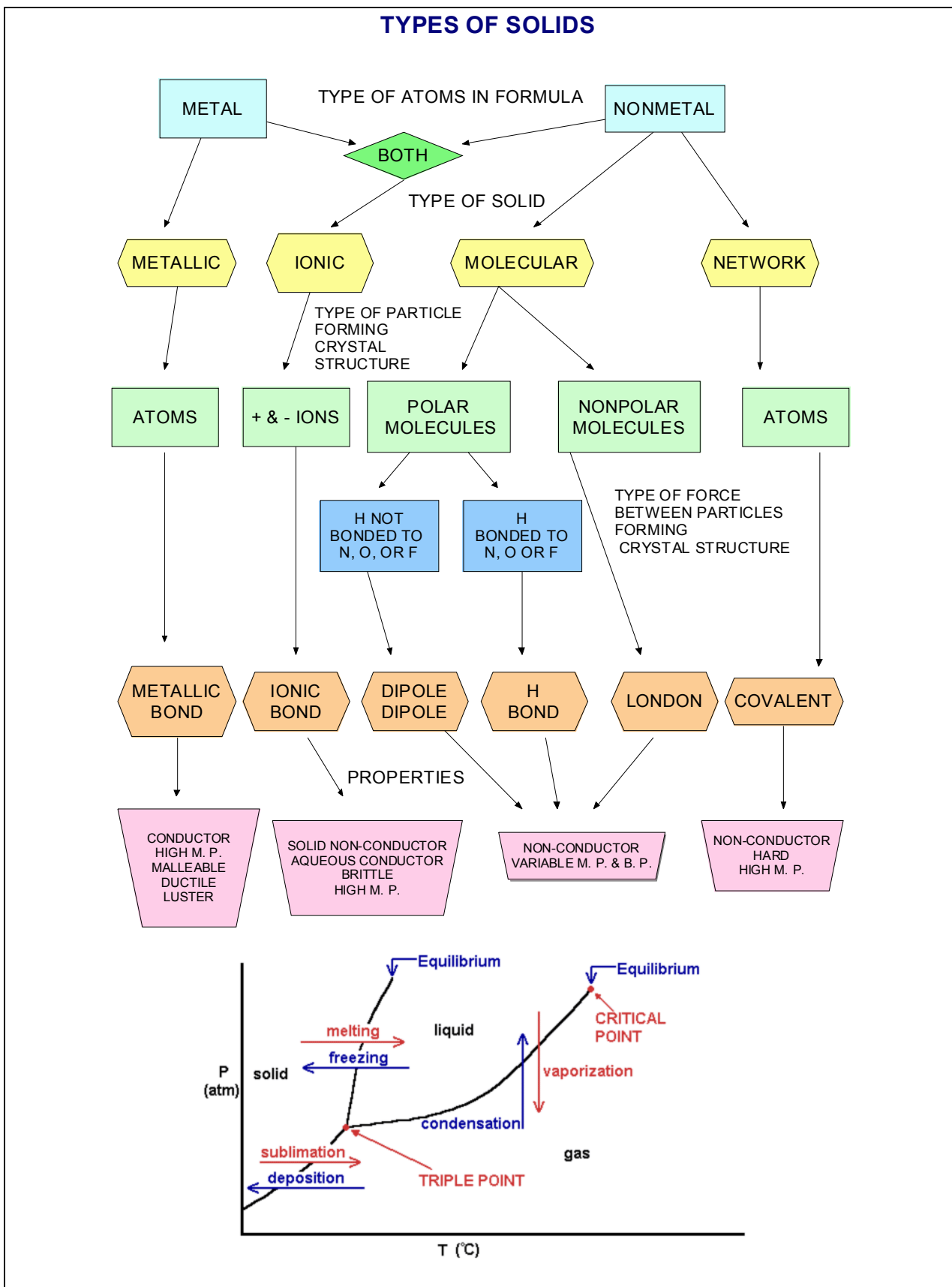
Resonances explains why bonds in molecules with a combination of single, double or triple bonds behave identical ie. C₆H₆

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: not on AP test	# valence \bar{e} - lone pairs + 1/2 shared \bar{e}

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



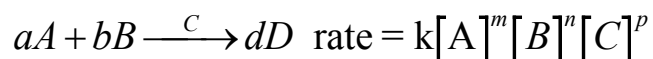
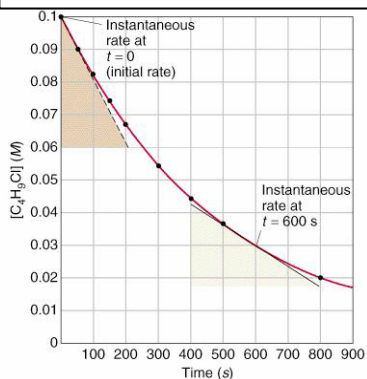
Solutions:

Molality:	$\text{molality} = \frac{\text{moles of solute}}{\text{kg solvent}}$
P _{soln} : vapor pressure of the solution X _{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

Instantaneous rate = slope of line tangent to the curve

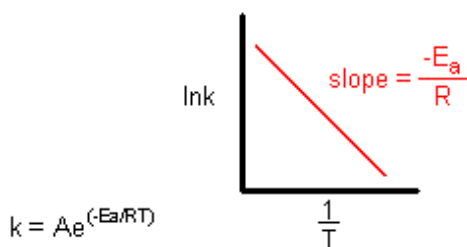


$m + n + p = \text{order of reaction}$
 $k = \text{rate constant } C \text{ is catalyst}$

$$\text{Rate} = \frac{-\Delta[A]}{\Delta t}$$

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

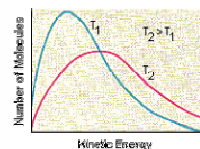
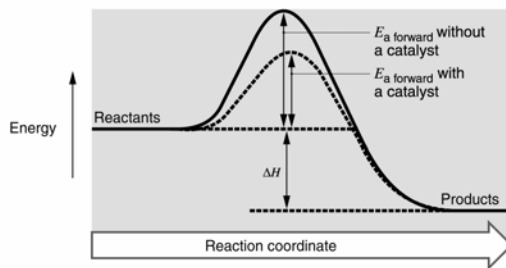
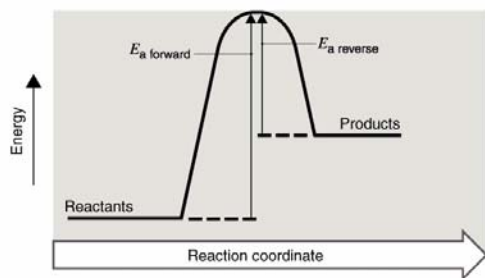


Derived Arrhenius Equation

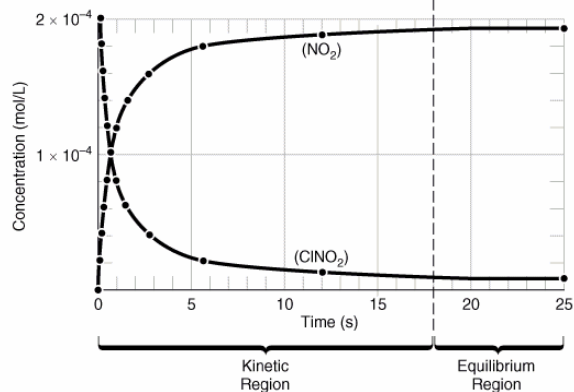
$$\ln k = \frac{-E_a}{R} \left(\frac{1}{T} \right) + \ln A$$

$$\ln k = \frac{-E_a}{R} \left(\frac{1}{T} \right) + \ln(A) \quad y = mx + b \text{ [graphical method]}$$

$$\ln \left(\frac{k_2}{k_1} \right) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right) \text{ [plug in equation]}$$

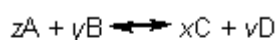


Chemical Equilibrium:



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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

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

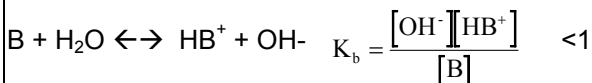
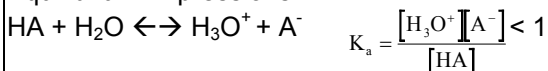
Changing the Coefficients Will Change K:

- 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:



CB's (anions of weak acids) - react like
nonhydroxide bases in water:
NaF: $\text{F}^- + \text{H}_2\text{O} \leftrightarrow \text{HF} + \text{OH}^-$
CA's (cations of weak bases) - react like acids in
water:
NH₄Cl: $\text{NH}_4^+ + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{NH}_3$

$$K_w = [\text{OH}^-][\text{H}^+] = 10^{-14} @25^\circ\text{C}$$

$$K_w = K_a \cdot K_b$$

$$\text{p}K_w = 14$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) = \text{p}K_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right)$$

% Dissociation

$$\% \text{ dissociation} = \frac{\text{amount dissociated}}{\text{initial concentration}} \times 100$$

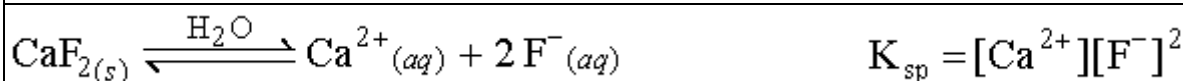
$$\underbrace{\text{HCl}}_{\text{acid}} + \underbrace{\text{NH}_3}_{\text{base}} \rightleftharpoons \underbrace{\text{NH}_4^+}_{\text{CA}} + \underbrace{\text{Cl}^-}_{\text{CB}}$$

amphoteric

$$\left\{ \begin{array}{l} \underbrace{\text{H}_2\text{PO}_4^-}_{\text{acid}} + \underbrace{\text{H}_2\text{O}}_{\text{base}} \rightleftharpoons \underbrace{\text{H}_3\text{O}^+}_{\text{CA}} + \underbrace{\text{HPO}_4^{2-}}_{\text{CB}} \\ \underbrace{\text{H}_2\text{PO}_4^-}_{\text{base}} + \underbrace{\text{H}_2\text{O}}_{\text{acid}} \rightleftharpoons \underbrace{\text{H}_3\text{PO}_4}_{\text{CA}} + \underbrace{\text{OH}^-}_{\text{CB}} \end{array} \right.$$

pH = pKa at 1/2 equivalence point

Solubility Equilibrium:

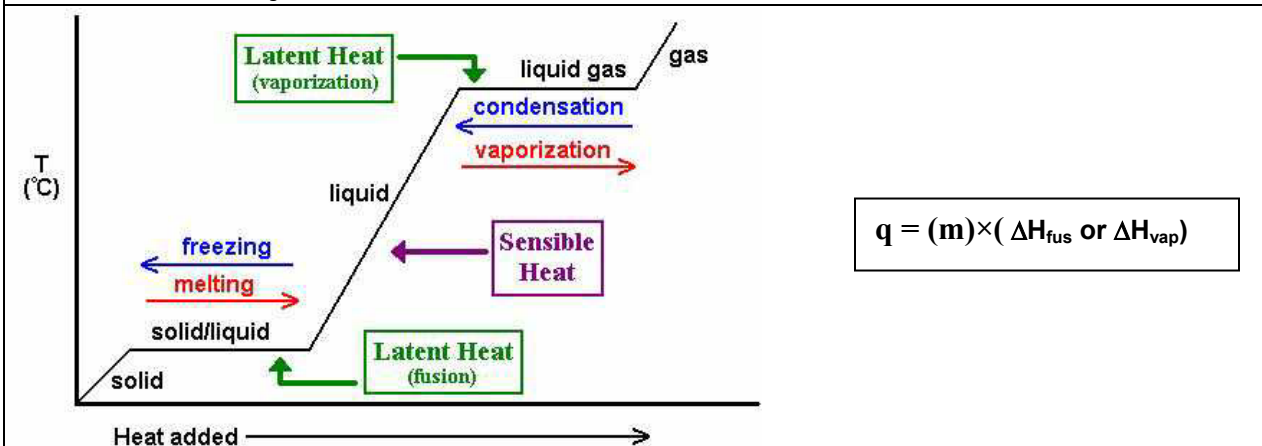


$Q_{sp} < K_{sp}$ - This means there are not enough ions in the solution. In order to return to equilibrium, more of the solid salt must dissociate into its ions.

$Q_{sp} = K_{sp}$ - 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 m = mass of solution or total mass</p>
<p>$\Delta H = q$ coffee cup calorimetry</p>	<p>$q = mc\Delta T$</p>

ΔH_f° for element = 0	$\Delta H^\circ = \sum n\Delta H_{f(\text{products})}^\circ - \sum n\Delta H_{f(\text{reactants})}^\circ$
	$\Delta S^\circ = \sum n\Delta S^\circ_{(\text{products})} - \sum n\Delta S^\circ_{(\text{reactants})}$
	$\Delta G^\circ = \sum n\Delta G_{f(\text{products})}^\circ - \sum n\Delta G_{f(\text{reactants})}^\circ$
$\Delta S_{\text{surroundings}} = \frac{-\Delta H}{\Delta T}$	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ <u>$-\Delta G^\circ = \text{rxn spontaneous}$</u>
$\Delta G = 0$ at equilibrium	$\Delta G = \Delta G^\circ + RT \ln Q$ $\Delta G^\circ = -RT \ln K$

Dependence Of Spontaneity On Temperature

ΔS	ΔH	<u>RESULTS</u>
POSITIVE,	NEGATIVE	SPONTANEOUS ALL TEMPERATURES
POSITIVE,	POSITIVE	SPONTANEOUS HIGH TEMPERATURES
NEGATIVE,	NEGATIVE	SPONTANEOUS LOW TEMPERATURES
NEGATIVE,	POSITIVE	NOT SPONTANEOUS

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

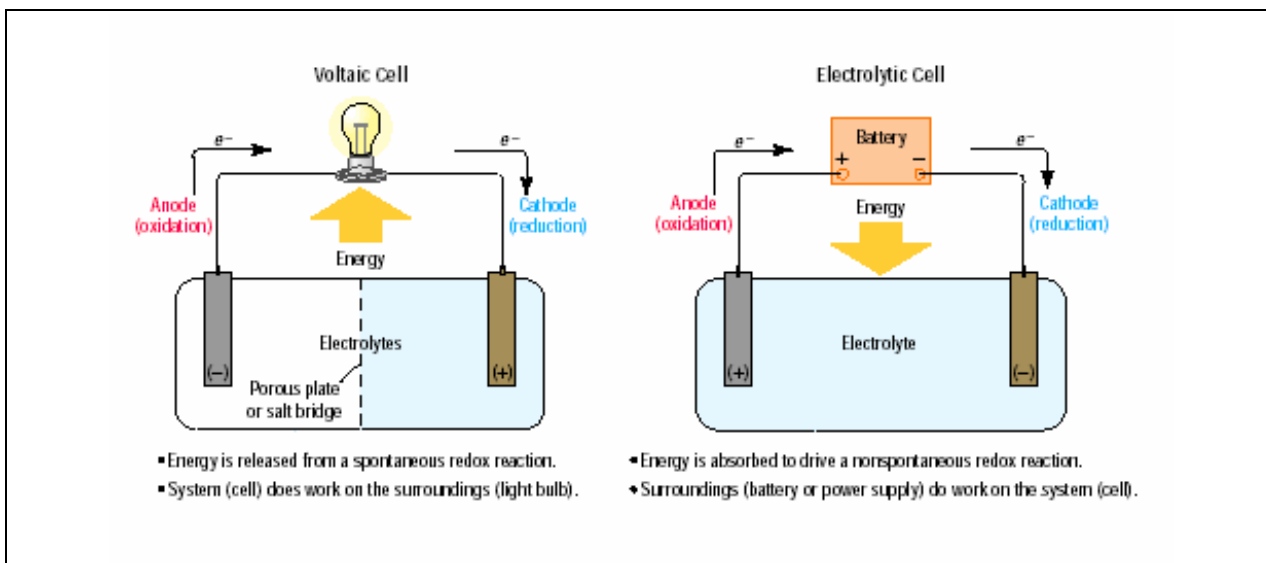
$$1 \text{ V} = \frac{1 \text{ J}}{1 \text{ C}} = \frac{(\text{joule})}{(\text{coulomb})}$$

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>Oxidation written before Reduction</p> <table border="1"> <tr> <td style="text-align: center;">Oxidation</td> <td style="text-align: center;">Reduction</td> </tr> <tr> <td style="text-align: center;">$\text{Zn} \text{Zn}^{2+}_{(1.0M)}$</td> <td style="text-align: center;">$\text{Cu}^{2+}_{(1.0M)} \text{Cu}$</td> </tr> </table> <p>Reactants listed before products</p> <p>Change in state Junction between half-cells (salt bridge) Concentration of aqueous solution</p>	Oxidation	Reduction	$\text{Zn} \text{Zn}^{2+}_{(1.0M)}$	$\text{Cu}^{2+}_{(1.0M)} \text{Cu}$	<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>
Oxidation	Reduction				
$\text{Zn} \text{Zn}^{2+}_{(1.0M)}$	$\text{Cu}^{2+}_{(1.0M)} \text{Cu}$				
Standard conditions 1 atm 25°C	$w_{\max} = -qE$ $\Delta G^{\circ} = -nFE^{\circ}$ $n = \text{number of moles of electrons transferred}$				
Nernst Equation F = faraday's constant $n = \text{number of electrons}$	$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln Q = E_{\text{cell}}^{\circ} - \frac{0.0592}{n} \log Q \quad \text{at } 25^{\circ}\text{C}$				
At equilibrium $E_{\text{cell}} = 0$	$\log K = \frac{nE^{\circ}}{0.0592}$				
<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

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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.