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**CHEMISTRY NOTES****TABLE OF CONTENTS**

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AQUEOUS REACTIONS .....	1
THERMOCHEMISTRY .....	1
ELECTRONIC STRUCTURE OF ATOMS .....	2
PERIODIC PROPERTIES OF THE ELEMENTS .....	2
BASIC CONCEPTS OF CHEMICAL BONDING .....	3
MOLECULAR GEOMETRY BONDING THEORIES .....	3
GASES .....	4
INTERMOLECULAR FORCES .....	4
PROPERTIES OF SOLUTIONS .....	5
CHEMICAL KINETICS .....	5
CHEMICAL EQUILIBRIUM .....	6
ACID BASE EQUILIBRIA .....	6
ADDITIONAL ASPECTS OF AQUEOUS EQUILIBRIA .....	6
CHEMICAL THERMODYNAMICS .....	6
ELECTROCHEMISTRY .....	7
NUCLEAR CHEMISTRY .....	7
ORGANIC CHEMISTRY .....	8
APPENDIX .....	9
Common Ions and Radicals .....	9
Naming Inorganic Compounds .....	11

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**SOURCES**Brown, Theodore L., et al. *Chemistry: The Central Science*. Englewood Cliffs: Prentice Hall, 1994.Oxtoby, David and Norman H. Nachtrieb. *Principles of Modern Chemistry*. New York: Saunders College Publishing, 1996.

## AQUEOUS REACTIONS

SOLUBILITY RULES		
<b>Mainly Water Soluble</b>	<b>Exceptions</b>	
NO <sub>3</sub> <sup>-</sup>		
C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup>		
Cl <sup>-</sup>	Ag <sup>+</sup> , Hg <sup>+</sup> , Pb <sup>2+</sup>	
Br <sup>-</sup>	Hg <sup>2+</sup>	
I <sup>-</sup>		
SO <sub>4</sub> <sup>-</sup>	Ca <sup>2+</sup> , Sr <sup>2+</sup> , Ba <sup>2+</sup>	
<b>Mainly Water Insoluble</b>	<b>Exceptions</b>	
S <sup>2-</sup>	2A elements	NH <sub>4</sub> <sup>+</sup>
CO <sub>3</sub> <sup>2-</sup>		1A elements
PO <sub>4</sub> <sup>3-</sup>		
OH <sup>-</sup>	Ca <sup>2+</sup> , Sr <sup>2+</sup> , Ba <sup>2+</sup>	

Driving Forces for Metathesis Reactions

1. The formation of a precipitate
2. The formation of a soluble weak electrolyte or a soluble nonelectrolyte
3. The formation of a gas that escapes from solution

## THERMOCHEMISTRY

$$\Delta E = q + w$$

 $q$  = heat added $w$  = work done

$$q = C_{sp} m \Delta T$$

 $C_{sp}$  = specific heat $m$  = mass

$$C_{sp} \mathcal{M} \approx 6.0 \text{ cal/}^\circ\text{C}$$

(Dulong-Petite law)

 $\mathcal{M}$  = molar mass

**ELECTRONIC STRUCTURE OF ATOMS**

$v = \lambda f$	$v = \text{velocity}$ $\lambda = \text{wavelength}$ $f = \text{frequency}$
$E = hf = -R_H \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$	$h = \text{Planck's constant} \approx 6.63 \times 10^{-34} \text{ J}\cdot\text{s}$ $R_H = \text{Rydberg constant} \approx 2.28 \times 10^{-18} \text{ J}$ $n = \text{principal quantum number of the electron}$
$\lambda = \frac{h}{mv}$	
$Z_{\text{eff}} = Z - S$	$Z_{\text{eff}} = \text{effective nuclear charge}$ $Z = \# \text{ protons in nucleus}$ $S = \text{average } \# \text{ of shielding electrons}$

## Quantum Numbers

1. principle quantum number ( $n$ )
2. azimuthal quantum number ( $l$ ) — values from 0 through  $n - 1$

<b>value of <math>l</math></b>	0	1	2	3
<b>subshell</b>	<i>s</i>	<i>p</i>	<i>d</i>	<i>f</i>

3. the magnetic quantum number ( $m_l$ ) — values from  $-l$  through  $l$
4. the spin number ( $m_s$ ) —  $+1/2, -1/2$

**PERIODIC PROPERTIES OF THE ELEMENTS**

- metal oxide + water  $\rightarrow$  metal hydroxide
- metal oxide + acid  $\rightarrow$  salt + water
- metal + acid  $\rightarrow$  salt + hydrogen gas
- metal + nonmetal  $\rightarrow$  salt
- nonmetal oxide + water  $\rightarrow$  acid
- nonmetal oxide + base  $\rightarrow$  salt + water
  
- 1A: Alkali Metals (M)
  - soft
  - low density
  - low melting temperatures
  - extremely reactive due to low ionization energy
  - $2M (s) + 2H_2O (l) \rightarrow 2MOH (aq) + H_2 (g)$
- 2A: Alkaline Earth Metals
  - harder, denser, and have higher melting temperatures than alkali metals
  - very reactive
- 7A: Halogens (X)
  - diatomic

**BASIC CONCEPTS OF CHEMICAL BONDING**

$$F = \frac{1}{4\pi\epsilon_0} \frac{q_1 q_2}{r^2}$$

(Coulomb's law)

 $F$  = force $\epsilon_0$  = permittivity constant  $\approx 8.854 \times 10^{-12} \text{ C}^2/\text{N}\cdot\text{m}^2$  $q$  = charge $r$  = radius(lattice energy between ions may be derived by multiplying Coulomb's equation by  $r$  (since  $W = Fx$ )Formal charge: # valence electrons – (total # nonbonding electrons +  $1/2$  # bonding electrons)

## Exceptions to Octet Rule

- molecules with an odd number of electrons
- molecules in which an atom has less than an octet
  1. H — only one pair of electrons
  2. Be — only two pairs of electrons
  3. B — only three pairs of electrons
- molecules in which an atom has more than an octet

**MOLECULAR GEOMETRY BONDING THEORIES**

Total electron pairs (steric number)	Electron-pair geometry	Bonding Pairs	Nonbonding pairs	Molecular geometry
2	linear (180°)	2	0	linear
3	trigonal planar (120°)	3	0	trigonal planar
		2	1	bent/angular
4	tetrahedral (109.47°)	4	0	tetrahedral
		3	1	trigonal pyramidal
		2	2	bent/angular
5	trigonal bipyramidal (90°, 120°)	5	0	trigonal bipyramidal
		4	1	seesaw
		3	2	T-shaped
		2	3	linear
6	octahedral (90°)	6	0	octahedral
		5	1	square pyramidal
		4	2	square planar

bond order =  $1/2$  (# bonding electrons – # antibonding electrons)

## Placement of Electrons in Molecular Orbitals

1. the number of molecular orbitals formed equals the number of atomic orbitals combined
2. atomic orbitals combine most effectively with other atomic orbitals of similar energy
3. the effectiveness with which two atomic orbitals combine is proportional to their overlap with one another
4. each molecular orbital can accommodate at most two electrons, with their spins paired (Pauli exclusion principle)
5. when molecular orbitals have the same energy, one electron enters each orbital (with parallel spins) before spin pairing occurs (Hund's rule)

**GASES**

$P V = n R T$	(ideal-gas equation)	$n =$ moles of gas $R =$ gas constant
$u = \sqrt{\frac{3R T}{\mathcal{M}}}$		$u =$ root-mean-square (rms) speed $\mathcal{M} =$ molar mass
$\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$	(Graham's Law of Effusion)	$r =$ rate of effusion (derived from rms equation)
$(P + \frac{n^2 a}{V^2})(V - n b) = n R T$	(Van der Waals equation)	$a =$ molecular attraction correction constant $b =$ molecular volume correction constant

Boyle's Law: P-V

Charles's Law: V-T

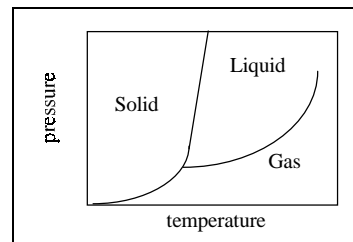
**Kinetic-Molecular Theory**

- gases consist of molecules that are in continuous, random motion
- the volume of gaseous molecules is negligible
- intermolecular forces between gas molecules are negligible
- molecular collisions are perfectly elastic
- the average kinetic energy of the molecules is proportional to the absolute temperature

**INTERMOLECULAR FORCES**

Intermolecular Forces (in order of generally increasing strength)

- London dispersion forces (induced dipole)
- dipole-dipole forces
- ion-dipole forces
- hydrogen bonding (usually H—F, H—O, or H—N bonds)



Type of solid	Form of unit particles	Forces between particles	Properties
molecular	atoms or molecules	Van der Waals	<ul style="list-style-type: none"> <li>soft</li> <li>low-moderately high melting point</li> <li>poor conduction</li> </ul>
covalent-network	atoms in a covalent-network	covalent bonds	<ul style="list-style-type: none"> <li>very hard</li> <li>high melting point</li> <li>poor conduction</li> </ul>
ionic	cations and anions	electrostatic attractions	<ul style="list-style-type: none"> <li>hard, brittle</li> <li>high melting point</li> <li>poor conduction</li> </ul>
metallic	atoms	metallic bonds	<ul style="list-style-type: none"> <li>soft-very hard</li> <li>low-very high melting point</li> <li>excellent conduction</li> <li>malleable and ductile</li> </ul>

$$\ln P = -\frac{\Delta H_{\text{vap}}}{R T} + C \quad (\text{Clausius-Clapeyron equation}) \quad \begin{array}{l} P = \text{vapor pressure} \\ R = \text{gas constant (J/mol}\cdot\text{K)} \end{array}$$

**PROPERTIES OF SOLUTIONS**

$C_g = k P_g$	(Henry's law)	$C_g$ = solubility of gas in solution phase (M) $k$ = Henry's law constant $P_g$ = partial pressure of gas over solution
$P_A = X_A P_A^\circ$	(Raoult's law)	$X_A$ = mole fraction of solvent $P_A^\circ$ = vapor pressure of pure solvent
$\Delta T_b = K_b m$		$\Delta T_b$ = boiling point elevation $K_b$ = molal boiling-point-elevation constant $m$ = molality
$\Delta T_f = K_f m$		$\Delta T_f$ = freezing point depression $K_f$ = molal freezing-point-depression constant
$p = M R T$		$p$ = osmotic pressure $M$ = molarity (derived from ideal-gas equation)

- molarity ( $M$ ) :  $\frac{\text{moles solute}}{\text{L solution}}$
- molality ( $m$ ) :  $\frac{\text{moles solute}}{\text{kg solvent}}$
- normality ( $N$ ) :  $\frac{\text{equivalents solute}}{\text{L solution}}$  (1 equivalent = moles substance that transfers 1 mol  $e^-$ )

## Colligative Properties

- lowering of vapor pressure
- increase in boiling point
- decrease in freezing point
- osmotic pressure

**CHEMICAL KINETICS**

Reaction rates affected by:

1. reactant concentrations
2. reacting temperature
3. presence of a catalyst
4. surface area of reactants or catalysts

Order	Rate Expression	Concentration–Time Relationship
0	rate = $k$	$[A]_t - [A]_0 = -k t$
1	rate = $k [A]$	$\ln \frac{[A]_t}{[A]_0} = -k t$
2	rate = $k [A]^2$	$\frac{1}{[A]_t} - \frac{1}{[A]_0} = k t$

$\text{rate} = -\frac{\Delta[A]}{a\Delta t} = -\frac{\Delta[B]}{b\Delta t} = \frac{\Delta[P]}{p\Delta t} = \frac{\Delta[Q]}{q\Delta t}$	$aA + bB \rightarrow pP + qQ$
$k = A e^{-E_a/RT}$ (Arrhenius equation)	$k$ = rate constant $A$ = frequency factor $E_a$ = activation energy $R$ = gas constant (J/mol·K)
$\ln \frac{k_2}{k_1} = -\frac{E_a}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)$	(derived from Arrhenius equation)

**CHEMICAL EQUILIBRIUM**

$$K_c = \frac{[P]^p [Q]^q}{[A]^a [B]^b}$$

$K_c$  = equilibrium constant  
 $aA + bB \rightarrow pP + qQ$

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

$K_p$  = equilibrium pressure constant  
 $R$  = gas constant (L·atm/mol·K)  
 $n$  = moles of gas

**ACID-BASE EQUILIBRIA**

	acid	base
<b>Arrhenius</b>	increase $[H^+]$	increase $[OH^-]$
<b>Brønsted-Lowry</b>	proton donor	proton acceptor
<b>Lewis</b>	electron-pair acceptor	electron-pair donor

**ADDITIONAL ASPECTS OF AQUEOUS EQUILIBRIA**

$$pH = pK_a + \log \frac{[X^-]}{[HX]} \quad (\text{Henderson-Hasselbalch equation})$$

**CHEMICAL THERMODYNAMICS**

$$S = k \ln W$$

$S$  = entropy  
 $k$  = Boltzmann's constant =  $\frac{R}{N_A} \approx 1.38 \times 10^{-23} \text{ J/K}$   
 $R$  = gas constant (J/mol·K)  
 $N_A$  = Avogadro's number  $\approx 6.022 \times 10^{23}$   
 $W$  = # of possible molecular arrangements

$$\Delta G = \Delta H - T \Delta S = \Delta G^\circ + R T \ln Q$$

$G$  = Gibbs free energy  
 $H$  = enthalpy  
 $R$  = gas constant (J/mol·K)  
 $Q$  = reaction quotient

## Laws of Thermodynamics

1. energy is conserved in any process
2. in any spontaneous process, the entropy of the universe increases
3. the entropy of a pure, crystalline solid at 0 K is zero.

**ELECTROCHEMISTRY**

$$\Delta G = -n F E$$

$n$  = moles of electrons transferred  
 $F$  = Faraday's constant  $\approx 96500 \text{ C/mol e}^-$   
 $E$  = electromotive force

$$E = E^\circ - \frac{R T}{n F} \ln Q$$

(Nernst equation)

$R$  = gas constant (J/mol·K)  
 (derived from Gibbs free energy equation)

	<b>cathode (site of reduction)</b>	<b>anode (site of oxidation)</b>
<b>voltaic cell</b>	+	-
<b>electrolytic cell</b>	-	+

**NUCLEAR CHEMISTRY**

$$E = m c^2$$

$c$  = speed of light  $\approx 3.00 \times 10^8 \text{ m/s}$

General Situations for Radioactive Decay

- nuclei above the belt of stability (high neutron:proton ratio)  $\rightarrow$   **$\beta$** -emission
  - nuclei below belt of stability (low neutron:proton ratio)  $\rightarrow$   ${}^0_1\text{e}$ -emission or  ${}^0_{-1}\text{e}$ -capture
  - nuclei with atomic numbers  $\geq 84 \rightarrow$   **$\alpha$** -emission
  - $\gamma$** -radiation almost always accompanies other radioactive emission
- Nuclei with 2, 8, 20, 28, 50, or 82 protons or 2, 8, 20, 28, 50, 82, or 126 neutrons are generally more stable than nuclei that do not contain these "magic numbers" of nucleons
  - nuclei with even numbers of both protons and neutrons are generally more stable than those with odd numbers of nucleons



## ORGANIC CHEMISTRY

$\text{R}-\text{O}-\text{H}$	$\text{R}-\text{O}-\text{R}'$	$\begin{array}{c} \text{O} \\    \\ \text{R}-\text{C}-\text{H} \end{array}$	$\begin{array}{c} \text{O} \\    \\ \text{R}-\text{C}-\text{R}' \end{array}$
alcohol (-ol)	ether (-ether)	aldehyde (-al)	keytone (-one)
$\begin{array}{c} \text{O} \\    \\ \text{R}-\text{C}-\text{O}-\text{H} \end{array}$	$\begin{array}{c} \text{O} \\    \\ \text{R}-\text{C}-\text{O}-\text{R}' \end{array}$	$\begin{array}{c} \text{R}' \text{ (or H)} \\    \\ \text{R}-\text{N}-\text{R}'' \text{ (or H)} \end{array}$	$\begin{array}{c} \text{O} \\    \\ \text{R}-\text{C}-\text{N} \end{array} \begin{array}{l} / \\ \backslash \end{array}$
carboxylic acid (-oic acid)	ester (-ate)	amine (-amine)	amide (-amide)

Precedence for numbering along carbon chains:

1. multiple bonds
2. functional groups
3. quantities of R groups

**General Form (at most one multiple bond)**

<b>Alkanes</b>	$\text{C}_n\text{H}_{2n+2}$
<b>Alkenes</b>	$\text{C}_n\text{H}_{2n}$
<b>Alkynes</b>	$\text{C}_n\text{H}_{2n-2}$

number of C-atoms	structural prefix
1	methyl-
2	ethyl-
3	propyl-
4	butyl-
5	pentyl-
6	hexyl-
7	heptyl-
8	octyl-
9	nonyl-
10	decyl-

## APPENDIX

## COMMON IONS AND RADICALS

<b>Cations</b>				
<b>+1</b>	<b>+2</b>	<b>+3</b>	<b>+4</b>	<b>+5</b>
Ag	Ba	Al	C	As (arsenic)
Au (aurous)	Be	As (arsenous)	Pb (plumbic)	P (phosphoric)
Cs	C	Au (auric)	Pt	Sb (stibic)
Cu (cuprous)	Ca	B	Si	
H	Cd	Bi	Sn (stannic)	
Hg (mercurous)	Cr (chromous)	Co (cobaltic)		
K	Co (cobaltous)	Cr (chromic)		
Li	Cu (cupric)	Fe (ferric)		
NH <sub>4</sub> (ammonium)	Fe (ferrous)	N		
Na	Hg (mercuric)	Ni (nickelic)		
	Mg	P (phosphorous)		
	Mn (manganous)	Sb (stibous)		
	Ni (nickelous)			
	Pb (plumbous)			
	Ra			
	Sn (stannous)			
	Sr			
	Zn			
<i>alkali metals</i>	<i>alkaline earth metals</i>	<i>rare earth metals (lanthanides)</i>		

## Anions

-1		-2		-3		-4		Common Acids	
AlO <sub>2</sub>	(aluminate)	B <sub>4</sub> O <sub>7</sub>	(tetraborate)	As	(arsenide)	Cl <sub>4</sub>	(tetrachloide)	HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	(acetic)
Br	(bromide)	CO <sub>3</sub>	(carbonate)	AsO <sub>3</sub>	(arsenite)	Fe(CN) <sub>6</sub>	(ferrocyanide)	HCl	(hydrochloric)
BrO <sub>3</sub>	(bromate)	C <sub>2</sub> O <sub>4</sub>	(oxalate)	AsO <sub>4</sub>	(arsenate)	SiO <sub>4</sub>	(silicate)	HNO <sub>3</sub>	(nitric)
CN	(cyanide)	C <sub>4</sub> H <sub>4</sub> O <sub>5</sub>	(tartrate)	BO <sub>3</sub>	(borate)			H <sub>3</sub> PO <sub>4</sub>	(phosphoric)
CNO	(cyanate)	CrO <sub>4</sub>	(chromate)	Fe(CN) <sub>6</sub>	(ferricyanide)			H <sub>2</sub> SO <sub>4</sub>	(sulfuric)
CNS	(thiocyanate)	Cr <sub>2</sub> O <sub>7</sub>	(dichromate)	N	(nitride)				
HCO <sub>3</sub>	(bicarbonate; hydrogen carbonate)	MoO <sub>4</sub>	(molybdate)	P	(phosphide)				
C <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	(acetate)	MnO <sub>3</sub>	(manganate)	PO <sub>3</sub>	(phosphite)				
Cl	(chloride)	O	(oxide)	PO <sub>4</sub>	(phosphate)				
ClO	(hypochlorite)	O <sub>2</sub>	(peroxide)	SbO <sub>4</sub>	(antimonate)				
ClO <sub>2</sub>	(chlorite)	HPO <sub>4</sub>	(hydrogen phosphate)						
ClO <sub>3</sub>	(chlorate)	S	(sulfide)						
ClO <sub>4</sub>	(perchlorate)	SO <sub>3</sub>	(sulfite)						
F	(fluoride)	SO <sub>4</sub>	(sulfate)						
H	(hydride)	S <sub>2</sub> O <sub>3</sub>	(thiosulfate)						
I	(iodide)	SiO <sub>3</sub>	(silicite)						
IO <sub>3</sub>	(iodate)	SnO <sub>3</sub>	(stannate)						
IO <sub>4</sub>	(periodate)								
MnO <sub>4</sub>	(permanganate)								
N <sub>3</sub>	(azide)								
NO <sub>2</sub>	(nitrite)								
NO <sub>3</sub>	(nitrate)								
OH	(hydroxide)								
H <sub>2</sub> PO <sub>4</sub>	(dihydrogen phosphate)								
HSO <sub>3</sub>	(bisulfite; hydrogen sulfite)								
HSO <sub>4</sub>	(bisulfate; hydrogen sulfate)								
S <sub>2</sub> O <sub>8</sub>	(persulfate)								
SiO <sub>4</sub>	(silicate)								
	<i>halogens</i>		<i>chalcogens</i>						

**NAMING INORGANIC COMPOUNDS****Ionic Compounds**

- cations
  - cation named and listed first; anion named and written last
  - elements that can form more than one positive ion
    - standard method:  
labeled with the charge of the ion indicated by a Roman numeral in parentheses following the named metal
    - older method:  
the suffixes *-ous* and *-ic* may be used with the Latin root of the element to distinguish between the higher and lower charged ions respectively
- anions
  - monatomic/simple anions: *-ide*
  - oxyanions (in order of :increasing amounts of O)
    - (*hypo-...-ite*)
    - *-ite*
    - *-ate* (common/representative oxyanion)
    - (*per-...-ate*)
- acids

<b>anion</b>	<b>acid</b>
<i>-ide</i>	<i>hydro-...-ic</i>
( <i>hypo-...</i> ) <i>-ite</i>	( <i>hypo-...</i> ) <i>-ous</i>
( <i>per-...</i> ) <i>-ate</i>	( <i>per-...</i> ) <i>-ic</i>

**Molecular Compounds**

- element with the more positive nature named and listed first; element with the more negative nature named with an *-ide* suffix
- if a pair of elements can form different molecular compounds, prefixes are used to indicate the numbers of atoms of each present element

<b>prefix</b>	<b>meaning</b>
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10